

# LESSON 4

## CHEMICAL BONDING

### 1. ELECTRONIC THEORY OF VALENCY

The theory of valency explains chemical combination in terms of electrons. The theory was developed independently by W. Kossel and G.N. Lewis (1916) and extended by Irving Langmuir (in 1919).

The chemical behaviour of an atom is determined to a large extent by the number and arrangement of electrons in the outer orbitals of the atom. Only these electrons are involved in chemical combination and so these are called the valence electrons.

#### COMPLETED ELECTRON OCTET OR DUPLET

Group 0 of the periodic table contains the noble gases. With the exception of helium which has a  $1s^2$  electron arrangement others have  $ns^2 np^6$  configuration in the outer orbitals.

He  $1s^2$

Ne  $1s^2 2s^2 2p^6$

Ar  $1s^2 2s^2 2p^6 3s^2 3p^6$

Kr  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

Xe  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6$

Since the atoms of the noble gases were not known to form chemical bonds, it was argued that the presence of 8 electrons (an electron octet) in the valence shell makes the atom stable. Therefore ***all other atoms must undergo bonding by gaining or losing or sharing electrons so as to acquire the electronic configuration of the nearest inert gas.*** The presence of 8 electrons gives the name **octet rule** to this concept. In the case of the first few elements such as hydrogen, lithium and beryllium the atoms combine in such a way as to attain the stable structure of helium with 2 electrons (**duplet**) in its only one valence shell. There are, however, many exceptions to the octet rule. Also compounds of noble gases, especially xenon, have been synthesized. The various types of chemical bonds are discussed below.

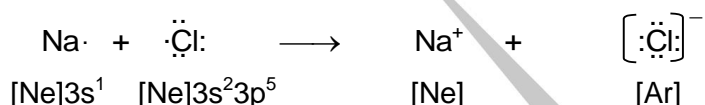
### 2. IONIC BOND (OR ELECTROVALENT LINKAGE)

An ionic bond is formed **by the complete transfer of electron(s) from one atom to another**. Atoms of metals generally lose electrons and those of non-metals gain electrons.

### (i) Formation of sodium chloride, NaCl

A sodium atom ( $Z = 11$ ) transfers its valence electron to a chlorine atom ( $Z = 17$ ). The sodium atom by losing an electron acquires the electronic configuration of neon ( $1s^2 2s^2 2p^6$ ) and becomes sodium ion  $\text{Na}^+$  carrying a unit positive charge. The chlorine atom by gaining an electron acquires the stable configuration of argon ( $1s^2 2s^2 2p^6 3s^2 3p^6$ ) and becomes a chloride ion,  $\text{Cl}^-$ , with a unit negative charge.

The transfer of electron results in the formation of the ionic bond.



- Here we have used **Lewis dot symbols** in which the symbol of an element is surrounded by dots (or crosses) to represent electrons in the outermost (valence) shell. Formulae of compounds using Lewis symbols are called **Lewis formulae**.
- When atoms form a bond by electron transfer, the **number of electrons lost and gained must be equal**, because the resulting ionic compound is neutral.
- The number of electrons lost or gained by an atom in the formation of an ionic bond is its valence. Thus Na and Cl have a valence of 1.
- Loss of electron** is called **oxidation**; thus Na is oxidised to  $\text{Na}^+$ . The **gain of electron** is **reduction**; thus Cl is reduced to  $\text{Cl}^-$ . Formation of an ionic bond from elements is an oxidation - reduction or **redox reaction**. Generally the metal is oxidised and the non-metal is reduced.
- $\text{Na}^+$  and Ne are **isoelectronic**: since they contain the **same number of electrons**. Similarly  $\text{Cl}^-$  and Ar are isoelectronic.
- Because  $\text{Na}^+$  and  $\text{Cl}^-$  carry opposite charges, electrostatic forces of attraction hold them together. Sodium chloride may be represented as  $\text{Na}^+ \text{Cl}^-$ .

### (ii) Formation of magnesium oxide, MgO



## 2.1 CONDITIONS FOR THE FORMATION OF IONIC BOND

- The difference between the electronegativity of two combining atoms must be greater than two.
- Low ionization energy of the metal:** Low ionization energy means that the metal atom requires only a small amount of energy to release its valence electron. For example, sodium, which has a low I.E. readily gives up its loosely held electron and forms  $\text{Na}^+$  ion. Metals of s-block have low ionization energies and so readily form the corresponding cations.
- Ionization energy of an element with a single electron in its valence shell is less than that with two electrons. In going across a period of the periodic table from left to right, I.E. increases and the formation of the cation is less likely. On going down a group, the outermost electron gets further away from the nucleus, and hence is more easily removed i.e., I.E. decreases; the formation of the cation becomes more likely.
- High electron affinity of the non-metal:** An atom with a high electron affinity releases a lot of energy when it takes up an electron and forms an anion. For example, chlorine which has a high

electron affinity, readily takes up an electron from the Na atom and forms  $\text{Cl}^-$  ion. Non-metals of groups VI A and VII A have high electron affinity and can form ionic bonds.

- ☛ In going across a period from left to right, electron affinity (energy released) increases and so the formation of the negative ion becomes more likely. On going down a group, electron affinity decreases and so the formation of anion becomes less likely.

**(iv) High lattice energy of the crystal:** In the formation of sodium chloride crystal, the  $\text{Na}^+$  ion attracts the  $\text{Cl}^-$  ion to form an ion-pair  $\text{Na}^+\text{Cl}^-$ . Since the electrostatic force of attraction is present in all directions, this ion-pair will attract other ion-pairs and build up into a crystal lattice. A crystal lattice is three dimensional basic pattern of points, in which each point corresponds to a unit of the crystal, say an ion (atom or molecule). As the lattice builds up, energy is released. The energy released when sufficient number of cations and anions come together to form 1 mole of the compound is called the lattice energy of the compound.

Therefore, an ionic compound is formed when the energy released in (iii) and (iv) exceeds the energy absorbed in (ii).

## 2.2 GENERAL CHARACTERISTICS OF IONIC COMPOUNDS

- (i) Generally ionic compounds are hard solids. As single ions of a metal are not associated in the solid with single ions of a non-metal, separate units of ionic compounds do not exist. It is, therefore, wrong to talk of a molecule of an ionic compound. The formula only indicates the ratio of number of ions and the crystal consists of a very large number of oppositely charged ions. Thus in NaCl crystal each  $\text{Na}^+$  ion is surrounded by  $6\text{Cl}^-$  ions and vice versa (in an octahedral arrangement). The attraction between  $\text{Na}^+$  and  $\text{Cl}^-$  ions is quite large.
- (ii) As a good deal of thermal energy is required to overcome the large electrostatic forces of attraction in an ionic crystal, ionic compounds have high melting and boiling points.
- (iii) Ionic compounds are commonly soluble in water and other polar solvents (which separate the ions). They are practically insoluble in organic solvents such as benzene, carbon tetrachloride, etc., as there is no attraction between ions and the molecules of the non-polar liquids.
- (iv) Ionic compounds are electrolytes. In the presence of an ionizing solvent such as water, the electrostatic forces between the ions are so greatly reduced that the ions get separated. (This is due to the electrostatic attraction between the ions and the polar molecules of the solvent.) The free ions in solution conduct electricity and on passing a current, the ionic compound undergoes chemical decomposition (called electrolysis). When an ionic compound is melted, the crystal lattice structure is broken and free ions are produced. It is the free movement of ions, which makes an ionic compound a conductor and to undergo electrolysis in the molten condition.
- (v) When an ionic compound dissolves in water, the ions get solvated (in this case hydrated). The energy released is called solvation energy. This energy counters wholly or in part the high lattice energy of the ionic compound. Insoluble ionic compounds (eg., sulphates, phosphates and fluorides of Ca, Sr and Ba) have very high lattice energies and the solvation energy of the constituent ions is insufficient to counteract the high lattice energies and make them soluble.
- (vi) The chemical properties of an ionic compound are the properties of its constituent ions. Thus all chlorides give the characteristic reactions of the chloride ion (reactions with conc.  $\text{H}_2\text{SO}_4$ ,  $\text{AgNO}_3$

solution, etc). All acids, which contain  $\text{H}^+$  ions give the same reactions (change blue litmus to red, effervesce with a carbonate, etc).

- (vii) Reactions between solutions of ionic compounds are almost instantaneous, because they are reactions between ions (and do not involve the breaking up of bonds as in covalent compounds, v.). For example, when silver nitrate solution is added to sodium chloride solution, silver chloride is immediately precipitated. The reaction may be represented thus:



### 3. COVALENT BOND

A covalent bond is formed by **the sharing of a pair of electrons between two atoms, each atom contributing one electron to the shared pair**. The shared pair of electrons should have opposite spins and they are localized between the two nuclei concerned. A covalent bond is usually represented by a short line (i.e., a dash) between the two atoms. Note that the covalent bond consists of a pair of electrons shared between two atoms, and occupying a combination of two stable orbitals, one of each atom; the shared electrons of each covalent bond are counted for each of the two atoms connected by the covalent bond.

The difference between the electronegativities of the combining atoms is less than two.

#### (1) Formation of the hydrogen molecule

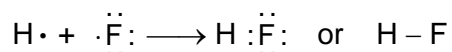
Each hydrogen atom requires 1 electron to become isoelectronic with helium, the nearest inert gas. The hydrogen atoms share their electrons thus:



Once the covalent bond is formed, the two bonding electrons are attracted by the two nuclei (instead of one) and the bonded state is more stable than the non-bonded state. The resultant attraction is responsible for the strength of the covalent bond.

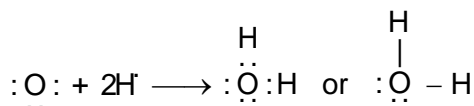
#### (2) Formation of hydrogen fluoride

The hydrogen atom has in its orbital 1 electron. It can achieve the helium configuration by forming a single covalent bond with another atom. Fluorine has 7 electrons in its outer, i.e., L shell. F can acquire the Neon configuration by forming a single covalent bond using its unpaired electron. This may be represented as follows.



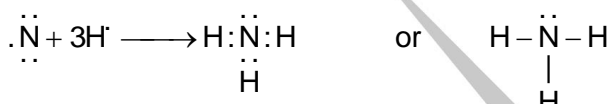
The single covalent bond holds the H and F atoms firmly together. Similarly we can explain the formation of HCl, HBr and HI.

### (3) Formation of water, H<sub>2</sub>O



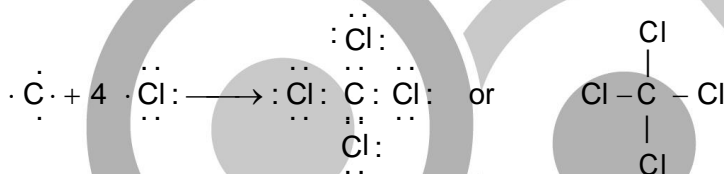
Similarly we can explain the formation of H<sub>2</sub>S (hydrogen sulphide), H<sub>2</sub>Se (hydrogen selenide) and H<sub>2</sub>Te (hydrogen telluride).

### (4) Formation of ammonia, NH<sub>3</sub>

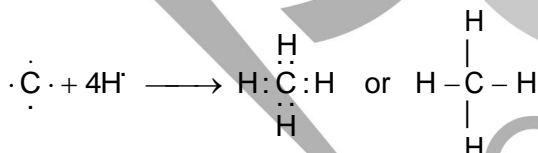


The structures of phosphine (PH<sub>3</sub>), arsine (AsH<sub>3</sub>) and stibine (SbH<sub>3</sub>) are similar to that of ammonia.

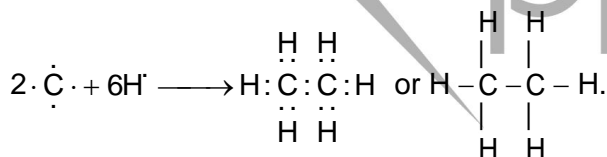
### (5) Formation of carbon tetrachloride, CCl<sub>4</sub>



### (6) Formation of methane, CH<sub>4</sub>



### (7) Formation of ethane, C<sub>2</sub>H<sub>6</sub>

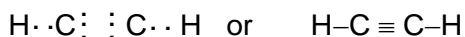


(i) The number of electrons needed by an atom to acquire its octet (C-4, N-3, O-2, Cl-1) is equal to the number of covalent bonds commonly formed.

(ii) When two pairs of electrons are shared between two atoms, there is a double bond as in ethylene, C<sub>2</sub>H<sub>4</sub>.



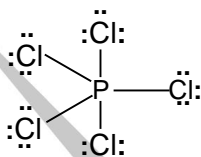
(iii) When three pairs of electrons are shared between two atoms, there is a triple bond as in acetylene, C<sub>2</sub>H<sub>2</sub>.



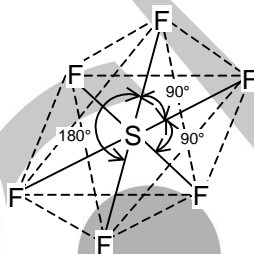
(iv) Generally all atoms involved in covalent bonding have completed octets (except hydrogen,

which has a duplet of electrons). Sometimes an atom forms more than 4 covalent bonds.

An example is phosphorus pentachloride  $\text{PCl}_5$ . In this molecule phosphorus atom is surrounded by 5 chlorine atoms, with each of which it forms a covalent bond (with some ionic character, about which we shall learn later). In this compound the phosphorus atom seems to use 5 of the nine orbitals of the M shell (rather than only 4 of the most stable orbitals). It seems likely that of the nine or more orbitals in the M, N and O shells, four are especially stable, but one or more others may be occasionally utilized.



Another example is  $\text{SF}_6$



### 3.1 GENERAL CHARACTERISTICS OF COVALENT COMPOUNDS

- (i) In a purely covalent compound the electrons in the bond are shared equally between the atoms linked by the bond; the resultant particles formed are not electrically charged. So, separate molecules of the covalent compounds exist. Covalent compounds may therefore be expected to be gases or low boiling liquids or soft, low melting solids at ordinary temperature. In the solid state they may be amorphous or present as molecular crystals, the molecules being held together by what are called weak van der Waals' forces of attraction.
- (ii) Since the molecules are held together by weak van der Waals' forces, covalent compounds (except those consisting of giant molecules) have low melting and boiling points; very little thermal energy is needed to overcome these weak intermolecular forces.
- (iii) They are non-electrolytes, i.e., they do not contain ions. Even in giant molecules such as diamond there are no free electrons. So they are very poor conductors of electricity.
- (iv) They are generally soluble in organic (non-polar) solvents such as benzene or carbon tetrachloride but are insoluble in water or other ionizing solvents. (The solubility of covalent compounds is very much dependent on the size of the molecules; giant molecules are practically insoluble in nearly all solvents.)
- (v) Reactions between covalent compounds are slow and often incomplete and reversible. This is so because the reaction involves breaking and making of bonds i.e., energy considerations are involved for reactants, activated complexes and products.
- (vi) A covalent bond is a space-directed bond and it may exhibit isomerism.



### 3.2 POLAR COVALENT BONDS – ELECTRONEGATIVITY

The shared pair of electrons may be shared equally between two atoms; then the covalent bond is said to be non-polar. Equal sharing occurs between identical atoms, as in  $\text{H} - \text{H}$  or  $\text{Cl} - \text{Cl}$  (i.e., in homonuclear molecules) or between identical atoms with identical neighbours as in  $\text{H}_3\text{C}-\text{CH}_3$ . When the two bonded atoms are dissimilar (i.e., in heteronuclear molecules) the sharing is unequal. For example a chlorine atom has a greater electron attracting power than a hydrogen atom; so in  $\text{H} - \text{Cl}$ , the shared pair of electrons are drawn more towards chlorine and away from hydrogen. The result is separation of charges within the molecule, the chlorine end acquiring a slight negative charge and the

hydrogen end a slight but equal positive charge:  $\text{H}^{\delta+} - \text{Cl}^{\delta-}$ . Such covalent bonds are said to be polar (i.e., bonds formed by sharing a pair of electrons between two atoms but displaced towards the nucleus of one of the bonded atoms).

**The net tendency of a bonded atom in a covalent molecule to attract the shared pair of electrons towards itself is known as electronegativity.** (This word does not mean the actual content of the electric charge, but just the tendency to acquire it in a molecule). Thus F is highly electronegative, but  $\text{F}^-$ , which has already an extra electron, is not.

**Table 3.2a: Table of Electronegativities (Pauling)**

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			As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0			Sb 1.9	Te 2.1	I 2.5

To assess the tendency of an atom of a given element to attract electrons towards itself in a covalent bond, relative electronegativity values are used. Table 3.2a. gives the relative electronegativity values of atoms calculated by Pauling (adopting arbitrarily the value of 4 units for the electronegativity of fluorine).

- ☛ (i) Electronegativity values increase across a period and decrease down a group.
- (ii) Smaller atoms have greater electronegativity than larger ones and so they attract electrons more towards them than larger ones. Alkali metals have low electronegativities and halogens high electronegativities.
- (iii) Atoms with nearly filled shells of electrons (e.g., halogens) have greater electronegativity than those with sparsely occupied shells.
- (iv) Elements with low electronegativity values such as Cs (0.8) and Rb (0.8) tend to form positive ions, i.e., these are metals. Elements with high electronegativity values such as F(4.0) and O(3.5) tend to form negative ions, i.e., these are non-metals.
- (v) Electronegativity value may be used to make rough predictions of the type of bonding to be found in a compound. The larger the difference between electronegativity values of two combining atoms, the more polar the covalent bond. If the difference is greater than 2, the greater the chance for ionic

bonding (i.e., the chance of covalent bond assuming 100% ionic character). From this point of view ionic bond may be considered to be an extreme case of a polar bond (with total separation of charges).

If the difference between the electronegativities of the combining atoms is zero or small, the bond is essentially non-polar.

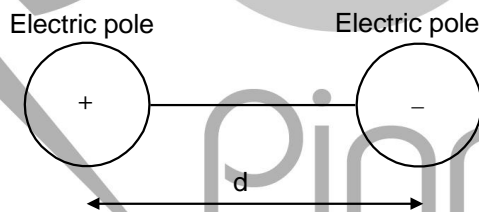
Let  $X_A$  and  $X_B$  represent the electronegativities of two atoms A and B. If  $X_B - X_A = 1.7$ , the covalent bond A – B is said to have 50% ionic character. On this basis, the % ionic character in some typical bonds is calculated (Table 3.2b). These calculations are very qualitative.

**Table 3.2b: % Ionic Character of Bonds**

C – H	N – H	O – H	F – H
4%	19%	39%	60%
C – F	C – Cl	C – Br	C – I
43%	11%	3%	0%

### 3.3 DIPOLE MOMENTS

A dipole consists of a positive and an equal negative charge separated by a distance within a molecule. The degree of polarity of a bond is given by the dipole moment ( $\mu$ ), which is the product of either charge ( $e$ ) and the distance ( $d$ ) between them.  $\mu = d \times e$ . 'e' is of the order of magnitude of the electronic charge, i.e., about  $10^{-10}$  esu and  $d$  is the distance between the atomic centres, i.e., about  $10^{-8}$  cm. Hence dipole moments may be expected to have values around  $10^{-10} \times 10^{-8} = 10^{-18}$  esu-cm. It is however, general practice to express dipole moments in **Debye units (D)**,  $1 \text{ D} = 10^{-18}$  esu-cm.



If the charge is in SI units (Coulombs) and  $d$  in metre,  $\mu$  will be coulomb-metre ( $\text{C} \cdot \text{m}$ ) units.

$$1\text{D} = 3.336 \times 10^{-30} \text{ C} \cdot \text{m}.$$

Any covalent bond which has a certain degree of polarity will have a corresponding dipole moment, though it does not follow that compounds containing such bonds will have dipole moments, for the **polarity of the molecule as a whole is the vector sum of the individual bond moments**. For example,  $\text{CO}_2$  has zero dipole moment, although the  $\text{C} = \text{O}$  bond is a polar bond. This shows that  $\text{CO}_2$  is a linear molecule,  $\text{O} = \text{C} = \text{O}$ , so that the dipole moments of the two  $\text{C} = \text{O}$  bonds cancel out. The  $\text{C} \rightarrow \text{Cl}$  bond has a definite polarity and a definite dipole moment but carbon tetrachloride has zero dipole moment because it is a tetrahedral molecule, and the resultant of the  $4\text{C} - \text{Cl}$  bond moments is zero. On the contrary  $\text{CH}_3\text{Cl}$ ,  $\text{CH}_2\text{Cl}_2$  and  $\text{CHCl}_3$  have definite dipole moments.

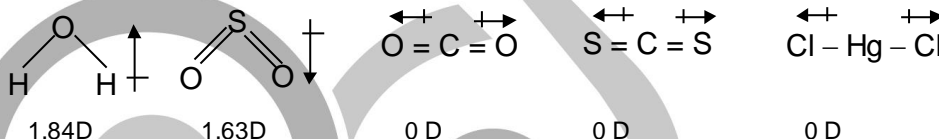


### 3.4 APPLICATIONS OF DIPOLE MOMENT MEASUREMENTS

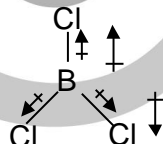
Dipole moment is a measure of the electrical dissymmetry (polarity) in the molecule and so its measurement provides valuable information concerning the shapes of molecules. Conversely, when the symmetry of the molecules is known, dipole moment could be estimated fairly.

#### (A) Inorganic substances:

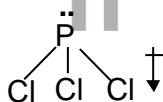
- (i) **Monatomic molecules** such as He, Ne, etc., have zero dipole moment because they are symmetrical.
- (ii) **Diatomic molecules** such as  $H_2$ ,  $Cl_2$  and  $N_2$  have no dipole moment; so these molecules are symmetrical.
- (iii) **Triatomic molecules:** Some of these molecules possess zero dipole moment; so they have a symmetrical linear structure, e.g.,  $CO_2$ ,  $CS_2$ ,  $HgCl_2$ . Others like water and sulphur dioxide have definite dipole moments. They are said to have angular or bent structure or V-shaped structure.



- (iv) **Tetratomic molecules:** Some molecules like  $BCl_3$  have zero dipole moment. They are said to possess a flat and symmetrical (triangular) structure; other examples are  $BF_3$ ,  $BBr_3$ ,  $CO_3^{2-}$  and  $NO_3^-$ .



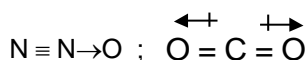
$PCl_3$ ,  $AsCl_3$ ,  $NH_3$ ,  $PH_3$ ,  $AsH_3$ ,  $H_3O^+$  have appreciable dipole moments. They possess trigonal pyramidal structures.



#### Illustration 1

**Question:** Both  $CO_2$  and  $N_2O$  are linear but dipole moment of  $CO_2$  is zero but for  $N_2O$  it is non-zero, why?

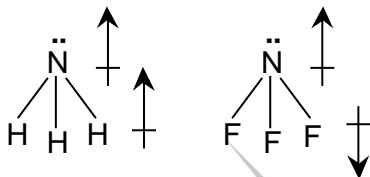
**Solution:** The answer lies in the structure of these molecules.  $CO_2$  is a symmetrical molecule while  $N_2O$  is unsymmetrical. Thus for  $N_2O$ , dipoles do not cancel each other, leaving the molecule with a resultant dipole, while the bond moment of  $CO_2$  cancel each other, so  $CO_2$  has no net dipole moment.



**Illustration 2**

**Question:** Compare the dipole moment of  $\text{NH}_3$  and  $\text{NF}_3$ .

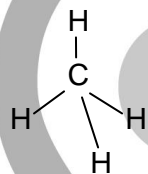
**Solution:** Let's draw the structure of both the compounds and then analyse their dipole directions.



The structure of both  $\text{NH}_3$  and  $\text{NF}_3$  are pyramidal with three bond pairs and one lone pair. In  $\text{NH}_3$ , as N is more electronegative than hydrogen, so the resultant bond dipole is towards N, which means that both the lone pair and bond pair dipoles are acting in the same direction and are summed up. In case of  $\text{NF}_3$ , the bond dipole (of N-F bonds) is acting towards fluorine, (as fluorine is more electronegative than N) so in  $\text{NF}_3$  the lone pair and bond pair dipoles are acting in opposition, resulting in a decreased dipole moment. Thus,  $\text{NH}_3$  has higher dipole moment than  $\text{NF}_3$ .

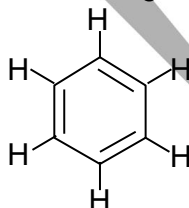
**(B) Organic substances**

(i) **Methane and  $\text{CCl}_4$**  have zero dipole moment. So they possess symmetrical tetrahedral structures with C atom at the centre of the tetrahedron.



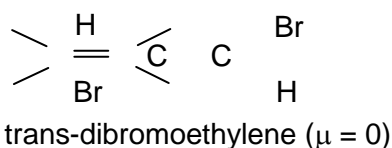
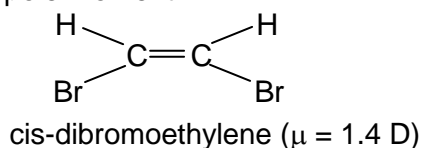
Methane

(ii) **Benzene** has zero dipole moment. All the 6 C and 6 H atoms are assumed to be in the same plane (symmetrical hexagonal structure).



(iii) **Measurement of dipole moments will enable** us to detect cis-and trans-isomers of organic compounds (you will learn about cis-trans or geometrical isomerism later in the organic chemistry).

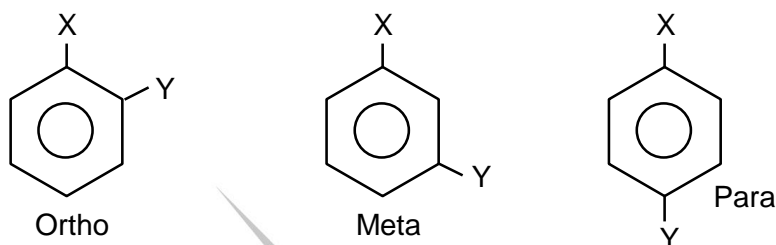
The trans- isomer, which is symmetrical, has zero dipole moment while the cis-isomer has a definite dipole moment.



**(iv) Dipole moment in aromatic ring system**

The dipole moments of the aromatic compounds present a very good illustration of dipole moment. We know that when substituted benzene is treated with reagent different products (namely ortho, meta and para products) are formed. The dipole moments of these products are different since the orientation of the groups is different. Let us take an example to clarify it. Let us take three isomers,

o-nitrophenol, m-nitrophenol and p-nitrophenol. We have also have three other isomers, o-aminophenol, m-aminophenol and p-aminophenol. We want to arrange these isomers in the order of their dipole moments.



In those cases where  $X = Y$ , the para isomer becomes symmetrical and have zero dipole moment. In order to find their dipole moment, we need to know about the nature of the groups linked to the benzene ring. In nitro phenols, one group (OH) is electron pushing and the other ( $\text{NO}_2$ ) is electron withdrawing while in aminophenols, both the groups (OH and  $\text{NH}_2$ ) attached are electron pushing. So, depending on the nature of the groups attached, the isomers have different dipole moment. Then calculation of dipole moment follows as:

**Case (i):** When X and Y both are electron pushing or electron withdrawing.

Let the bond dipole of C–X bond is represented by  $\mu_1$  and that of C–Y bond by  $\mu_2$ . Now let us assume that the electron pushing groups have +ve bond moment and the electron withdrawing groups have –ve bond moment. The net dipole moment is the resultant of two bond dipoles at different orientations.

$$\mu_{\text{ortho}} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos 60^\circ} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cdot \frac{1}{2}}$$

$$\therefore \mu_o = \sqrt{\mu_1^2 + \mu_2^2 + \mu_1\mu_2}$$

$$\mu_{\text{meta}} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos 120^\circ}$$

$$\therefore \mu_m = \sqrt{\mu_1^2 + \mu_2^2 - \mu_1\mu_2}$$

$$\mu_{\text{para}} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos 180^\circ} = \sqrt{\mu_1^2 + \mu_2^2 - 2\mu_1\mu_2}$$

$$\therefore \mu_p = \mu_1 - \mu_2$$

From the above expressions of  $\mu_o$ ,  $\mu_m$  and  $\mu_p$ , it is clear that when both X and Y are of the same nature i.e., both are electron withdrawing or both are electron pushing the para product has the least dipole moment and ortho product has the highest dipole moment. When  $X = Y$ ,  $\mu_1 = \mu_2$ , thus  $\mu_p$  would be zero.

**Case (ii):** When X is electron pushing and Y is electron withdrawing or vice versa.

Let the bond moment of C–X dipole is  $\mu_1$  and that of C–Y dipole is  $\mu_2$ .

$$\mu_o = \sqrt{\mu_1^2 + (-\mu_2)^2 + 2\mu_1(-\mu_2)\cos 60^\circ}$$

$$= \sqrt{\mu_1^2 + \mu_2^2 - \mu_1\mu_2}$$

$$= \sqrt{(\mu_1 + \mu_2)^2 - 3\mu_1\mu_2}$$

$$\mu_{\text{meta}} = \sqrt{\mu_1^2 + (-\mu_2)^2 + 2\mu_1(-\mu_2)\cos 120^\circ}$$

$$\begin{aligned}
 &= \sqrt{\mu_1^2 + \mu_2^2 + \mu_1\mu_2} \\
 &= \sqrt{(\mu_1 + \mu_2)^2 - \mu_1\mu_2} \\
 \mu_{\text{para}} &= \sqrt{\mu_1^2 + (-\mu_2)^2 + 2\mu_1(-\mu_2)\cos 180^\circ} \\
 &= \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2} \\
 &= \mu_1 + \mu_2
 \end{aligned}$$

Looking at the expressions of  $\mu_o$ ,  $\mu_m$  and  $\mu_p$ , it is clear that the para isomer has the highest dipole moment and ortho has the least.

### 3.5 DIPOLE MOMENT AND PERCENTAGE IONIC CHARACTER

The measured dipole moment of a substance may be used to calculate the percentage ionic character of a covalent bond in simple molecules.

1 unit charge = Magnitude of electronic charge =  $4.8 \times 10^{-10}$  e.s.u.

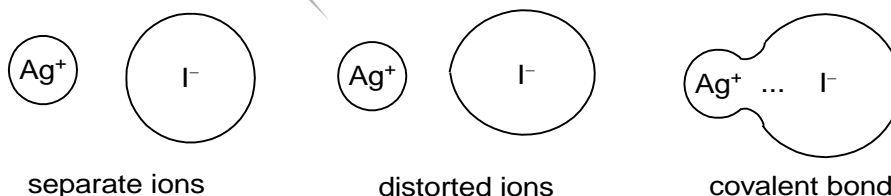
1 D =  $1 \times 10^{-18}$  e.s.u-cm.

$$\therefore \% \text{ ionic character} = \frac{\text{Observed dipole moment}}{\text{Theoretical dipole moment}} \times 100$$

Theoretical dipole moment is confined to when we assume that the bond is 100 % ionic and it is broken into ions while observed dipole moment is with respect to fractional charges on the atoms of the bond.

### 3.6 TRANSITION FROM IONIC TO COVALENT BOND – FAJANS' RULE

Just as a covalent bond may have partial ionic character, an ionic bond may also show a certain degree of covalent character. When two oppositely charged ions approach each other closely, the cation would attract the electrons in the outer shell of the anion and simultaneously repel its nucleus. This produces distortion or polarization of the anion, which is accompanied by some sharing of electrons between the ions, i.e., the bond acquires a certain covalent character. The formation of a covalent bond between two ions may be illustrated with reference to formation of AgI.



### FACTORS INFLUENCING ION – DEFORMATION OR INCREASING COVALENT CHARACTER

#### (i) Large charge on the ions:

The greater the charge on the cation, the more strongly will it attract the electrons of the anion. For example,  $\text{Al}^{3+}$  can distort  $\text{Cl}^-$  ion more than  $\text{Na}^+$  ion. So aluminium chloride is a covalent compound whereas  $\text{NaCl}$ ,  $\text{AlF}_3$ ,  $\text{AgF}$  are ionic.

**(ii) Small cation and large anion:**

For a small cation, the electrostatic force with which its nucleus will attract the anion will be large. Moreover a large anion cannot hold the electrons in its outermost shell, especially when they are attracted by a neighbouring cation. Hence there will be increased covalence with a small cation and a large anion, as in  $\text{AgI}$ .

**(iii) Cation with a non-inert gas type of electronic configuration:**

A cation with a 18 electron outermost shell such as  $\text{Ag}^+$  ( $[\text{Kr}] 4d^{10}$ ) polarizes anions more strongly than a cation with a 8 electron arrangement as in  $\text{K}^+$ . The 'd' electrons in  $\text{Ag}^+$  do not screen the nuclear charge as effectively as the 's' and 'p' electron shell in  $\text{K}^+$ . Thus  $\text{AgI}$  is more covalent than  $\text{KI}$ , although  $\text{Ag}^+$  and  $\text{K}^+$  ions are nearly of the same size. Cuprous and mercurous salts are covalent.

The above statements regarding the factors, which influence covalent character, are called Fajans' rules. It can thus be seen easily that there is nothing like a purely ionic compound or a purely covalent compound.

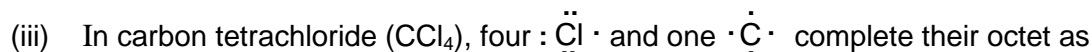
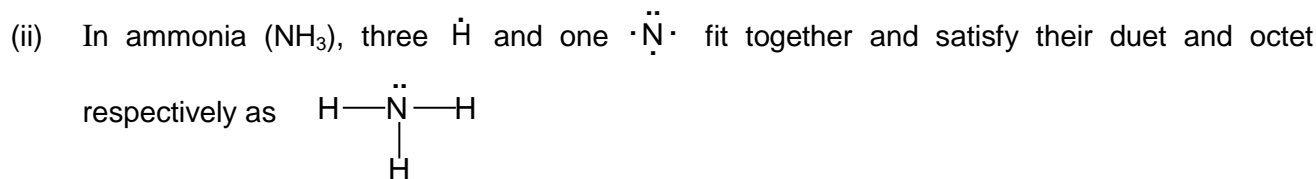
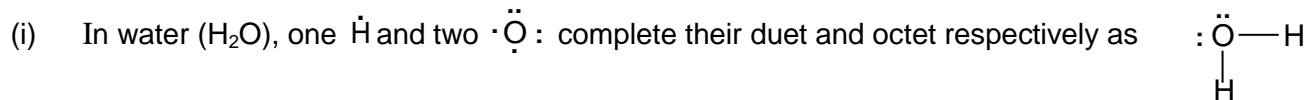
## 4. LEWIS STRUCTURES OF MOLECULES

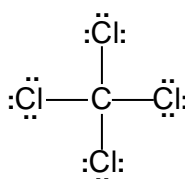
The formula of a molecule shows the number of atoms of each element but does not show the bonding arrangement of the atoms. To represent the bonding pattern in a molecule, the electron dot symbols of the elements are arranged such that the shared pairs and unshared pairs (called lone pairs) are shown and the octet rule (or duet for hydrogen) is satisfied. For example, a molecule of fluorine is shown as  $:\ddot{\text{F}}:\ddot{\text{F}}:$  or  $:\ddot{\text{F}}-\ddot{\text{F}}:$  and a molecule of hydrogen fluoride is shown as



Arrangement of dot symbols used to represent molecules are called **Lewis structures**. Lewis structures do not convey any information regarding the shape of the molecule. Usually, the shared pairs of electrons are represented by lines between atoms and any unshared pairs are shown as dot pairs.

Lewis structures are written by fitting the element dot symbols together to show shared electron pairs and to satisfy the octet rule. For example,





For the given molecules, we have adopted hit & trial method to fit the dot symbols together and satisfy the octet rule. But remember that hydrogen form one bond, oxygen forms two bonds, nitrogen three bonds and carbon forms four bonds. For simpler molecules, the hit & trial method works perfectly but for slightly complicated polyatomic species, this may give us more than one possible structure. Thus, a systematic approach is needed to design the Lewis structures of such polyatomic species. But before proceeding further, let us understand the limitation of this approach.

#### 4.1 LIMITATIONS OF LEWIS THEORY OF DRAWING STRUCTURE

This method would be applicable to only those molecules/species, which follow octet rule except hydrogen.

There are three kinds of molecules/species, which do not follow octet rule.

- Molecules, which have contraction of octet. Such molecules are electron deficient. For example,  $\text{BH}_3$ ,  $\text{BF}_3$ ,  $\text{BCl}_3$ ,  $\text{AlCl}_3$ ,  $\text{GaCl}_3$  etc.
- Molecules, which have expansion of octet. Such species have more than eight electrons in their outermost shell. This is possible in those molecules, which have vacant d-orbitals, thus they can expand their octet. For example,  $\text{PCl}_5$ ,  $\text{SF}_6$  etc.
- Molecules containing odd number of electrons (in total) cannot satisfy octet rule. Such species are called odd electron species and are paramagnetic in nature due to presence of unpaired electron. For example,  $\text{NO}$ ,  $\text{NO}_2$  and  $\text{ClO}_2$ .

Pinnacle



## 4.2 METHOD OF DRAWING LEWIS STRUCTURES

To draw the Lewis structures of polyatomic species, follow the given sequence.

- (i) First calculate  $n_1$ .

$n_1$  = Sum of valence electron of all the atoms of the species  $\pm$  net charge on the species.

For a negatively charged species, electrons are added while for positively charged species, the electrons are subtracted. For an uninegatively charged species, add 1 to the sum of valence electrons and for a dinegatively charged species, add 2 and so on.

- (ii) Then calculate  $n_2$ .

$n_2 = (8 \times \text{number of atoms other than H}) + (2 \times \text{number of H atoms})$

- (iii) Subtract  $n_1$  from  $n_2$ , which gives  $n_3$ .

$n_3 = n_2 - n_1$  = number of electrons shared between atoms = number of bonding electrons.

$\frac{n_3}{2} = \frac{n_2 - n_1}{2}$  = number of shared (bonding) electron pairs = number of bonds.

- (iv) Subtracting  $n_3$  from  $n_1$  gives  $n_4$ .

$n_4 = n_1 - n_3$  = number of unshared electrons or non-bonding electrons.

$\frac{n_4}{2} = \frac{n_1 - n_3}{2}$  = number of unshared electron pairs = number of lone pairs.

- (v) Identify the central atom. Generally, the central atom is the one, which is least electronegative of all the atoms, when the other atoms do not contain hydrogen. When the other atoms are hydrogen only, then the central atom would be the more electronegative atom. However some exceptions are possible, for example  $\text{Cl}_2\text{O}$ .

- (vi) Now around the central atom, place the other atoms and distribute the required number of bonds (as calculated in step (iii)) & required number of lone pairs (as calculated in step (iv)), keeping in mind that every atom gets an octet of electrons except hydrogen.

- (vii) Then calculate the formal charge on each atom of the species.

Formal charge on an atom = number of valence electrons of the atom – number of bonds formed by that atom – number of unshared electrons ( $2 \times$  lone pairs) of that atom.

- (viii) When two adjacent atoms get opposite formal charges, then charges can be removed by replacing the covalent bond between the atoms by a dative (co-ordinate) bond. This bond will have the arrowhead pointing towards the atom with negative formal charge. It is not mandatory to show the dative bonds unless required to do so.

- (ix) The given Lewis structure should account for the factual aspects of the molecule like resonance (delocalization), bond length,  $p\pi$ - $d\pi$  back bonding etc.

Sometimes, there are more than one acceptable Lewis structure for a given species. In such cases, we select the most plausible Lewis structure by using formal charges and the following guidelines:

- For neutral molecules, a Lewis structure in which there are no formal charges is preferable to one in which formal charges are present.
- Lewis structures with large formal charges (+2, +3 and/ or -2, -3 and so on) are less plausible than those with small formal charges.
- Among Lewis structures having similar distributions of formal charges, the most plausible structure is the one in which negative formal charges are placed on the more electronegative atoms.

**Illustration 3**

**Question:** Determine Lewis structure of  $\text{NO}_3^-$  ion.

**Solution:**

(i)  $n_1 = 5 + (6 \times 3) + 1 = 24$

(ii)  $n_2 = (4 \times 8) = 32$

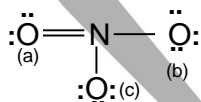
(iii)  $n_3 = n_2 - n_1 = 32 - 24 = 8$

$$\therefore \text{Number of bonds} = \frac{8}{2} = 4$$

(iv)  $n_4 = n_1 - n_3 = 24 - 8 = 16$

$$\therefore \text{Number of lone pairs} = \frac{16}{2} = 8$$

- (v) Nitrogen is the central atom (as it is less electronegative than O). Arranging three O atoms around it and distributing 4 bonds and 8 lone pairs as



- (vi) Calculating formal change on each atom.

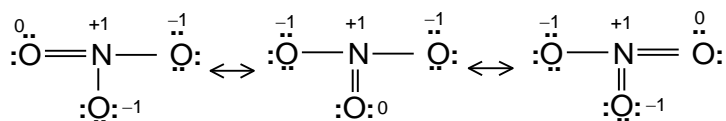
$$\text{Formal charge on N} = 5 - 4 - 0 = +1$$

$$\text{Formal charge on O (a)} = 6 - 2 - 4 = 0$$

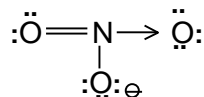
$$\text{Formal charge on O (b)} = 6 - 1 - 6 = -1$$

$$\text{Formal charge on O (c)} = 6 - 1 - 6 = -1$$

Thus, the structures can now be shown as



Final structure of  $\text{NO}_3^-$  is therefore shown as



which even accounts for resonance in  $\text{NO}_3^-$  ion.

**Illustration 4**

**Question:** Determine Lewis structure of  $\text{CN}^-$  ion.

**Solution:**

(i)  $n_1 = 4 + 5 + 1 = 10$

(ii)  $n_2 = (2 \times 8) = 16$

(iii)  $n_3 = n_2 - n_1 = 16 - 10 = 6$

$$\therefore \text{Number of bonds} = \frac{6}{2} = 3$$

(iv)  $n_4 = n_1 - n_3 = 10 - 6 = 4$

$$\therefore \text{Number of lone pairs} = \frac{4}{2} = 2$$

- (v) Carbon is the central atom (C is less electronegative than N) and arrange N, number of bonds and number of lone pairs around it as



(vi) Formal charge on C =  $4 - 3 - 2 = -1$

Formal charge on N =  $5 - 3 - 2 = 0$

Thus, final Lewis structure of  $\text{CN}^-$  would be


**Illustration 5**

**Question:** Draw Lewis structure for  $\text{NH}_4^+$  ion.

**Solution:**

(i)  $n_1 = 5 + (4 \times 1) - 1 = 8$

(ii)  $n_2 = (8 \times 1) + (2 \times 4) = 16$

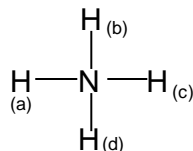
(iii)  $n_3 = n_2 - n_1 = 16 - 8 = 8$

$$\therefore \text{Number of bonds} = \frac{8}{2} = 4$$

(iv)  $n_4 = n_1 - n_3 = 8 - 8 = 0$

$$\therefore \text{Number of lone pairs} = 0$$

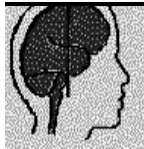
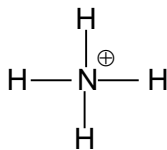
- (v) Nitrogen being the central atom, distributing other atoms (H) around it, and 4 bonds with the 4 H atoms, the structure looks like



(vi) Formal charge on N =  $5 - 4 - 0 = +1$

Formal charge on H(a)/H(b)/H(c)/H(d) =  $1 - 1 - 0 = 0$

Thus, final Lewis structure of  $\text{NH}_4^+$  would be



Draw Lewis structures for the following molecules/species.

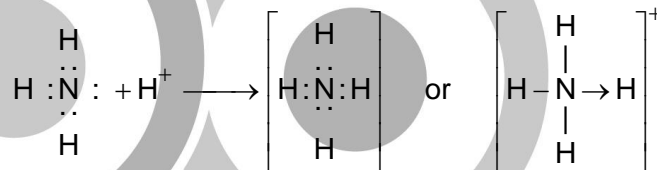
- (a) CO (b) CO<sub>2</sub> (c) NO<sub>2</sub><sup>-</sup> (d) CO<sub>3</sub><sup>2-</sup> (e) ClO<sub>3</sub><sup>-</sup>  
(f) SO<sub>4</sub><sup>2-</sup> (g) O<sub>3</sub> (h) NCO<sup>-</sup> (isocyanate ion) (i) CNO<sup>-</sup> (cyanate ion)

## 5. CO-ORDINATE COVALENT BOND OR DATIVE BOND

We have seen that in the formation of a covalent bond between two atoms, each atom contributes one electron to the shared pair. Sometimes **both the electrons of the shared pair may come from one of the atoms. The covalent bond thus formed is a co-ordinate bond or dative bond.**

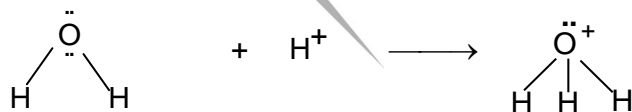
### (i) Formation of ammonium ion

The ammonia molecule has a lone pair of electrons i.e., an unshared pair. The hydrogen ion H<sup>+</sup>, has an empty s orbital. The lone pair comes to be shared between the nitrogen and hydrogen atoms:

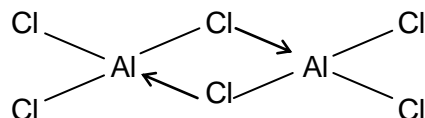


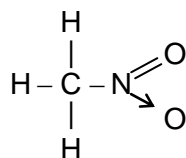
Nitrogen atom is called the donor and H<sup>+</sup>, the acceptor. The arrow-head in N → H shows that N-atom is electron donor and H-atom is electron acceptor. NH<sub>3</sub> is a neutral molecule. H<sup>+</sup> carries a unit positive charge; so NH<sub>4</sub><sup>+</sup> ion carries a unit positive charge. Once the NH<sub>4</sub><sup>+</sup> ion is formed, all the N-H bonds become identical.

### (ii) Hydronium ion, H<sub>3</sub>O<sup>+</sup>



### (iii) Aluminium chloride, Al<sub>2</sub>Cl<sub>6</sub>



(iv) Nitromethane,  $\text{CH}_3\text{NO}_2$ 


## GENERAL CHARACTERISTICS OF COORDINATE COVALENT COMPOUNDS

As is to be expected the properties of coordinate covalent compounds are mostly similar to the properties of covalent compounds.

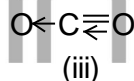
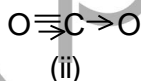
- (i) The nuclei in coordinate covalent compounds (such as in  $\text{NH}_4^+$ ) are held firmly by shared electrons and so do not form ions in water.
- (ii) Their covalent nature makes them sparingly soluble in water and more soluble in organic solvents.
- (iii) The coordinate bond is also rigid and directional, just like covalent bonds.

## 6. RESONANCE

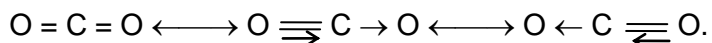
Carbon dioxide may be represented by Lewis dot formula as



The bond length of  $\text{C}=\text{O}$  is 1.22 Å, but the actual measured value is 1.15 Å. Further  $\text{CO}_2$  is quite stable and does not show the characteristic reactions of the carbonyl group, as shown by aldehydes and ketones. Without shifting, the relative positions of atoms of  $\text{CO}_2$  can be represented by two more Lewis formulae:



In (ii) and (iii), the two bonds between C and O are different, one being a triple bond and the other a single bond. Both the C–O bonds in  $\text{CO}_2$  are identical. It is now obvious that none of these structures actually represents  $\text{CO}_2$ . To explain this difficulty the concept of resonance was introduced, according to which  $\text{CO}_2$  cannot be accurately depicted by any Lewis formula. The actual structure of  $\text{CO}_2$  is a resonance hybrid of the three structures:



These different structures are called the **canonical or contributing structures**. **The actual structure of  $\text{CO}_2$  is different from the canonical structures and although it is closely related to them, the actual structure cannot be represented on paper** using the accepted symbols. **All the molecules of  $\text{CO}_2$  have the same structure.** Usually, a double-headed arrow  $\longleftrightarrow$  is used between the canonical structures.

## 6.1 CONDITIONS FOR RESONANCE

Resonance can occur when the canonical structures

- (i) have the constituent atoms in the same relative positions;
- (ii) have nearly the same energy;
- (iii) have the same number of unpaired electrons (to allow for continuous change from one type of bond to another);
- (iv) differ in the distribution of electrons around the constituent atoms;
- (v) (molecules or ions) are planar.

## 6.2 RESONANCE ENERGY

The resonance hybrid is a more stable structure than any of the contributing structures. This means that resonance hybrid has less energy than any of the contributing structures. **The difference in energy between the actual observed energy of the resonance hybrid and the most stable of the contributing structures is called resonance energy.**

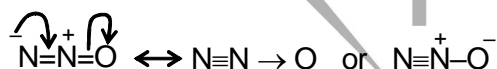
For  $\text{CO}_2$ , structure (i) has less energy than structure (ii).

## 6.3 OTHER EXAMPLES OF RESONANCE

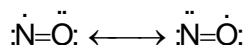
### (i) Sulphur dioxide $\text{SO}_2$



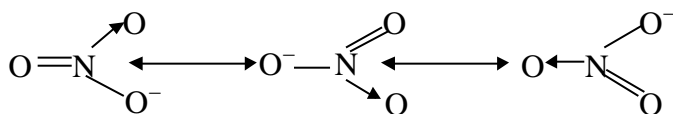
### (ii) Nitrous oxide (dinitrogen oxide), $\text{N}_2\text{O}$



### (iii) Nitric oxide, $\text{NO}$

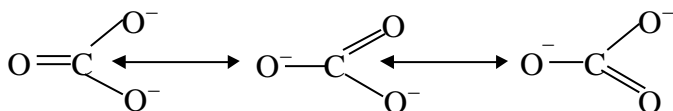


### (iv) Nitrate ion, $\text{NO}_3^-$ (planar, triangular)

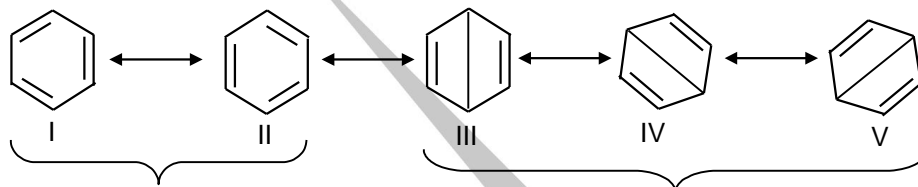




(v) **Carbonate ion,  $\text{CO}_3^{2-}$**  (planar, triangular)



(vi) **Benzene,  $\text{C}_6\text{H}_6$** . It is a resonance hybrid of the following structures (hexagonal, planar).



**Kekulé Structures**

**Dewar Structures**

C–C bond length is 1.54 Å; C=C bond length is 1.34 Å. In benzene, all the C–C bonds are identical in length, 1.39 Å, i.e., intermediate between those of single and double bonds. Note that shortening of bond length and therefore increased stability is an indication of the existence of resonance [Decrease in dipole moment also indicates resonance]. Resonance energy of benzene is –152 kJ/mol.

## 7. MOLECULAR GEOMETRY AND VSEPR THEORY

Molecular geometry is the three-dimensional arrangement of atoms in a molecule. A molecule's geometry affects its physical and chemical properties such as melting point, boiling point and the types of reactions it undergoes. In general, bond length and bond angles are determined by experiments. However, there is a simple procedure to predict the overall geometry of a molecule or ion with considerable accuracy, if we know the number of electrons surrounding a central atom in its Lewis structure. The basis of this approach is the assumption that electron pairs in the valence shell (outermost electron-occupied shell of an atom) of an atom repel one another. In a polyatomic species, the repulsion between electrons in different bonding pairs causes them to remain as far as possible. Thus, the geometry assumed by the species ultimately minimizes the repulsion. This approach is called valence-shell electron-pair repulsion (VSEPR) theory because it accounts for the geometric arrangements of electron pairs around a central atom in terms of the electrostatic repulsion between electron pairs.

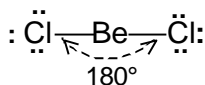
Molecules in this theory are divided into two categories, depending on whether the central atom has lone pair of electrons or not.

7.1 Molecules in which the central atom has no lone pairs.

For simplicity, we will consider molecules that contain only two types of atoms, A and B, of which A is the central atom. These molecules have the general formula  $\text{AB}_x$ , where  $x$  is an integer 2, 3 ..... (if  $x = 1$ , the molecule will be diatomic, which is linear by definition).

**(a) Molecules having general formula  $AB_2$**

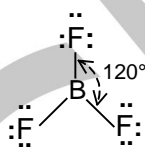
$BeCl_2$  is representing the general formula  $AB_2$ . The Lewis structure of beryllium chloride in the gaseous state is



Since the bonding pairs repel each other, they must be at opposite ends of a straight line in order for them to be as far apart as possible. Thus,  $ClBeCl$  bond angle is predicted to be  $180^\circ$  and the molecule is **linear**.

**(b) Molecules having general formula  $AB_3$**

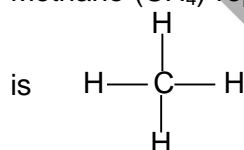
The general formula  $AB_3$  is represented by the molecule  $BF_3$ .  $BF_3$  has three bonding pairs, which points to the corners of an equilateral triangle with boron at the center of the triangle.



This geometry of  $BF_3$  is referred as **trigonal planar** with  $FBF$  bond angle to be  $120^\circ$ . In this structure, all four atoms lie in the same plane,

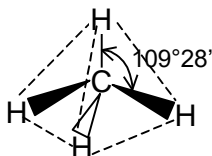
**(c) Molecules having general formula  $AB_4$**

Methane ( $CH_4$ ) represents the best example of this class of molecules. The Lewis structure of  $CH_4$



The four bonding pairs in  $CH_4$  can be arranged to form a tetrahedron, so as to minimize the repulsion between them. A tetrahedron has four faces, all of which are equilateral triangles.

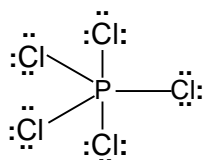
In a **tetrahedral** molecule, the central atom (carbon) is located at the center of the tetrahedron and the other four atoms (H) are at the corners. The  $HCH$  bond angles are all  $109^\circ 28'$ .



**(d) Molecules with general formula  $AB_5$**

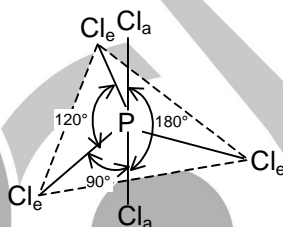
The general formula  $AB_5$  is represented by the molecule  $PCl_5$ . The Lewis structure of  $PCl_5$  (in gas

phase) is



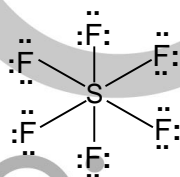
The only way to minimize the repulsive forces among the five bonding pairs is to arrange the P–Cl bonds in the form of a **trigonal bipyramid**. Joining two tetrahedrons along a common triangular base can generate a trigonal bipyramid.

The central atom (P) is at the center of the common triangular with the surrounding atoms positioned at the five corners of the trigonal bipyramid. The atoms that are above and below the triangular plane are said to occupy axial positions and those, which are in the triangular plane, are said to occupy equatorial positions. The angle between any two equatorial bonds is  $120^\circ$ , that between an axial bond and an equatorial bond is  $90^\circ$  and that between two axial bonds is  $180^\circ$ .

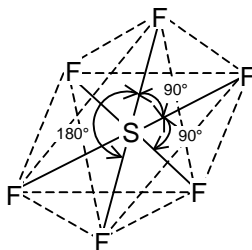


#### (e) Molecules having general formula $AB_6$

The molecule  $SF_6$  exhibits the general formula  $AB_6$ . The Lewis structure of  $SF_6$  is



The most stable arrangement of the six S–F bonding pairs is in the shape of an **octahedron**. An octahedron has eight faces and can be generated by joining two square pyramids on a common base. The central atom (S) is at the center of the square base and the surrounding atoms (F) are at the six corners. All bond angles are  $90^\circ$  except the one made by the bonds between the central atom and the pairs of atoms that are diametrically opposite to each other, which is  $180^\circ$ . Since, all the bonds are equivalent in an octahedral molecule, the terms axial and equatorial are not used here.



#### 7.2 MOLECULES IN WHICH THE CENTRAL ATOM HAS ONE OR MORE LONE PAIRS

In such molecules, there are three types of repulsive interactions—between bonding pairs, between lone pairs and between a bonding pair and a lone pair. In general, according to VSEPR theory, the repulsive forces decrease in the following order: lone pair–lone pair repulsion > lone pair – bond pair

repulsion > bond pair – bond pair repulsion.

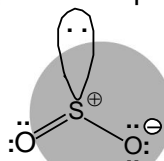
Bond pair electrons are held by the attractive forces exerted by the nuclei of the two bonded atoms. These electrons have less “spatial distribution” than lone pairs i.e., they take up less space than lone pair electrons, which are associated with only one nuclei (or one atom). Because lone-pair electrons in a molecule occupy more space, they experience greater repulsion from neighbouring lone pairs and bonding pairs.

To keep track of total number of bonding pairs and lone pairs, we designate molecules with lone pairs as  $AB_xE_y$ , where A is the central atom, B is the surrounding atoms and E is a lone pair on A. Both  $x$  and  $y$  are integers,  $x = 2, 3, \dots$  and  $y = 1, 2, \dots$ . Thus,  $x$  and  $y$  denote the number of surrounding atoms and number of lone pairs on the central atom, respectively.

**(a) Molecules with general formula  $AB_2E$**

Example of this type is  $SO_2$ . The Lewis structure of  $SO_2$  is  $:\ddot{O}=\overset{+}{S}-\ddot{O}^-:$

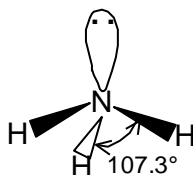
VSEPR theory treats double bond and triple bonds as though they were single bonds. Thus,  $SO_2$  molecule can be viewed as having three electron pairs on the central atoms, of which, two are bonding pairs and one is a lone pair. The overall arrangement of three electron pairs is trigonal planar. But since one of the electron pair is a lone pair, the  $SO_2$  molecule looks like



This shape is referred as **bent or angular**. The shape is determined only by the bonding pairs and not by lone pairs. Since lone pair repels the bonding pairs more strongly, the SO bonds are pushed together slightly and the OSO angle is less than  $120^\circ$ .

**(b) Molecules having general formula  $AB_3E$**

The general formula  $AB_3E$  is exhibited by the molecule  $NH_3$ . Ammonia has overall four electron pairs, of which three are bonding pairs and one is lone pair. The overall arrangement of four electron pairs is tetrahedral but since one of the electron pairs is a lone pair, so the shape of  $NH_3$  is **trigonal pyramidal**. Because the lone pairs repels the bonding pairs more strongly, the three N–H bonds are pushed closer together. Thus the HNH bond angle is smaller than the ideal tetrahedral angle of  $109^\circ 28'$ .

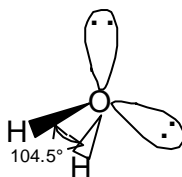


**(c) Molecules with general formula  $AB_2E_2$**

Example of such a molecule is  $H_2O$ . A water molecule has 2 bonding pairs and two lone pairs  $H-\ddot{O}-H$ .

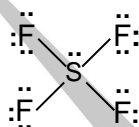
The overall arrangement of the four electron pairs in water is tetrahedral. However, unlike  $NH_3$ ,  $H_2O$  has 2 lone pairs on the central O atom. These lone pairs tend to be as far from each other as possible. Consequently, the two OH bonding pairs are pushed toward each other and  $H_2O$  shows

even greater deviation from tetrahedral angle than in  $\text{NH}_3$ . The shape of  $\text{H}_2\text{O}$  is referred as **bent or angular** with HOH bond angle of  $104.5^\circ$ .

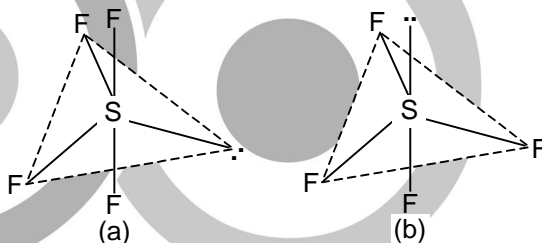


**(d) Molecules having general formula  $\text{AB}_4\text{E}$**

Example to this class of molecule is  $\text{SF}_4$ . The Lewis structure of  $\text{SF}_4$  is



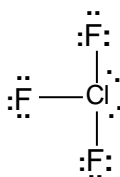
The S atom in  $\text{SF}_4$  has 5 electron pairs, which can be arranged as trigonal bipyramidal. In  $\text{SF}_4$ , since one of the electron pair is a lone pair, so the molecule can have any one of the following geometries:



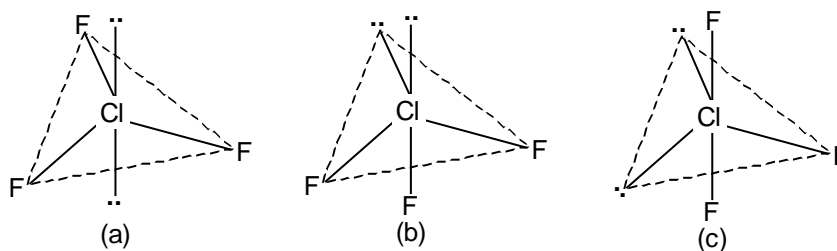
In (a), the lone pair occupies an equatorial position and in (b), it occupies an axial position. Repulsion between the electrons pairs in bonds only  $90^\circ$  apart are greater than repulsion between electron pairs in bonds  $120^\circ$  apart. Each axial bond has three electron pairs  $90^\circ$  away while each equatorial bond has only two electron pairs  $90^\circ$  away. Thus axial bonds (electron pairs) experience greater repulsion than the equatorial bonds. Thus, atoms at the equatorial positions are closer to the central atom than atoms at the axial positions i.e. equatorial bond lengths are smaller than axial bond lengths. Thus, when the central atom also has lone pairs along with the bonding pairs, it will occupy a position where the repulsions are less, so lone pairs in trigonal bipyramidal are more comfortable at equatorial positions. Thus, (a) is the appropriate structure of  $\text{SF}_4$ . It is referred as **see-saw shaped or irregular tetrahedron**.

**(e) Molecules with general formula  $\text{AB}_3\text{E}_2$**

Example of this type is  $\text{ClF}_3$ . The Lewis structure of  $\text{ClF}_3$  is



The Cl atom in  $\text{ClF}_3$  has 5 electron pairs, of which 2 are lone pairs and 3 are bonding pairs. The molecule can have any of the following three geometries:



In structure (a), there are 6 lone pair–bond pair repulsions at  $90^\circ$  and one lone pair – lone pair repulsion at  $180^\circ$ . In structure (b), 1 lone pair – lone pair repulsion is at  $90^\circ$  and there are 3 lone pair–bond pair repulsions at  $90^\circ$ , 2 at  $120^\circ$  and 1 at  $180^\circ$ . While in structure (c), there are 4 lone pair–bond pair repulsions at  $90^\circ$ , 2 at  $120^\circ$  and one lone pair–lone pair repulsion at  $120^\circ$ . The structure (b) is out rightly ruled out since the lone pair–lone pair repulsion is of highest magnitude. Among structures (a) and (c), each structure has 4 lone pair–bond pair repulsions at  $90^\circ$ . Apart from these repulsions, (a) has 1 lone pair–lone pair repulsion at  $180^\circ$  and 2 lone pair–bond pair repulsions at  $90^\circ$  while (c) has 1 lone pair–lone pair repulsion at  $120^\circ$  & 2 lone pair–bond pair repulsions also at  $120^\circ$ . So, the structure (c) has overall lesser repulsions than (a). Thus, (c) is the appropriate structure of  $\text{ClF}_3$ . It is called **T-shaped** structure.

### 7.3 PREDICTING GEOMETRY OF SPECIES USING VSEPR THEORY

With the help of VSEPR theory, we can predict the geometry of various species in a systematic way. The scheme makes use of the following steps:

- (i) Identify the central atom and count the number of valence electrons on the central atom.
- (ii) Add to this, the number of other atoms (which form single bonds only). Here, oxygen atoms are not added as they form two bonds.
- (iii) If the species is an anion, add negative charges and if it is a cation, subtract positive charges.
- (iv) This gives us a number, which we refer as N.
- (v) Divide N by 2 and we get the sum of bonding and non-bonding electron pairs.  

$$\frac{N}{2} = \text{Number of other atoms} + \text{number of lone pairs.}$$
- (vi) Compare the result  $\left(\frac{N}{2} \text{ value}\right)$  with the value given in table, corresponding to the given number of lone pairs.



N/2 value	No. of lone pairs	Shape of the species	Example
2	0	Linear	HgCl <sub>2</sub> , BeCl <sub>2</sub>
3	0	Triangular planar	BF <sub>3</sub> , AlCl <sub>3</sub> , BH <sub>3</sub> , NO <sub>3</sub> <sup>-</sup> , SO <sub>3</sub>
	1	Angular or bent	SnCl <sub>2</sub> , SO <sub>2</sub> , NO <sub>2</sub> <sup>-</sup>
4	0	Tetrahedral	CCl <sub>4</sub> , BeF <sub>4</sub> <sup>2-</sup> , BF <sub>4</sub> <sup>-</sup> , PCI <sub>4</sub> <sup>+</sup> , ClO <sub>4</sub> <sup>-</sup>
	1	Trigonal pyramidal	NH <sub>3</sub> , PCl <sub>3</sub> , PF <sub>3</sub> , ClO <sub>3</sub> <sup>-</sup>
	2	Angular or bent	H <sub>2</sub> O, H <sub>2</sub> S, OF <sub>2</sub> , ClO <sub>2</sub> <sup>-</sup>
	3	Linear	ClO <sup>-</sup>
5	0	Trigonal bipyramidal	PCl <sub>5</sub> , PF <sub>5</sub>
	1	See-saw or irregular tetrahedron	SF <sub>4</sub> , IF <sub>4</sub> <sup>+</sup>
	2	T-shaped	ClF <sub>3</sub> , BrF <sub>3</sub>
	3	Linear	I <sub>3</sub> <sup>-</sup> , Br <sub>3</sub> <sup>-</sup> , XeF <sub>2</sub>
6	0	Octahedral or square bipyramidal	SF <sub>6</sub> , PCI <sub>6</sub> <sup>-</sup>
	1	Square pyramidal	BrF <sub>5</sub> , IF <sub>5</sub>
	2	Square planar	ICI <sub>4</sub> <sup>-</sup> , XeF <sub>4</sub> , IF <sub>4</sub>
7	0	Pentagonal bipyramidal	IF <sub>7</sub>

**Note:** XeF<sub>6</sub> does not have octahedral structure. It's structure is **capped octahedron**.

Let us see the usefulness of the VSEPR theory to predict the geometry of few molecules/ ions.

**(i) BeCl<sub>2</sub> molecule:**

The central atom is Be and it has two other Cl atoms.

$$\therefore \frac{N}{2} = \frac{2+2}{2} = 2$$

Since, the number of other atoms are 2, so the number of lone pairs are zero. Thus, shape of BeCl<sub>2</sub> is **linear**.

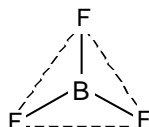


**(ii) BF<sub>3</sub> molecule:**

In BF<sub>3</sub>, central atom is boron and it has three other atoms.

$$\therefore \frac{N}{2} = \frac{3+3}{2} = 3$$

Since, the number of other atoms are three, so the number of lone pairs are zero. Therefore, shape is **triangular planar**.

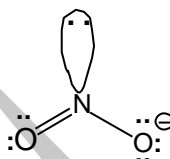


**(iii)  $\text{NO}_2^-$  ion:**

The central atom in  $\text{NO}_2^-$  is N and it has two other atoms.

$$\therefore \frac{N}{2} = \frac{5+1}{2} = 3$$

Since, the number of other atoms are 2, so the number of lone pairs would be 1. Thus, shape of  $\text{NO}_2^-$  ion is **angular or bent**.

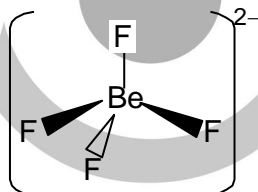


**(iv)  $\text{BeF}_4^{2-}$  ion:**

In  $\text{BeF}_4^{2-}$ , the central atom is Be and it has four other F atoms.

$$\therefore \frac{N}{2} = \frac{2+4+2}{2} = 4$$

The number of lone pairs are zero, as the number of other atoms are 4. Therefore, shape of  $\text{BeF}_4^{2-}$  is **tetrahedral**.

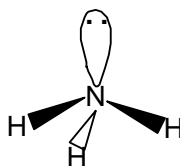


**(v)  $\text{NH}_3$  molecule:**

In  $\text{NH}_3$ , the central atom is N and it has 3 other H atoms.

$$\therefore \frac{N}{2} = \frac{5+3}{2} = 4$$

Since, the number of other atoms is 3, so the number of lone pairs would be 1. Therefore, shape of  $\text{NH}_3$  molecule is **trigonal pyramidal**.

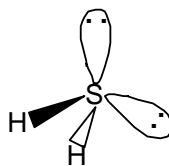


**(vi)  $\text{H}_2\text{S}$  molecule:**

The central atom is S and there are 2 other H atoms in  $\text{H}_2\text{S}$  molecule.

$$\therefore \frac{N}{2} = \frac{6+2}{2} = 4$$

Since, the number of other atoms is 2, so the number of lone pairs would be 2. Thus, the shape of  $\text{H}_2\text{S}$  is **angular or bent**.

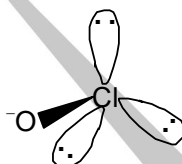


**(vii)  $\text{ClO}^-$  ion:**

The central atom is Cl and it has one other atom.

$$\therefore \frac{N}{2} = \frac{7+1}{2} = 4$$

The number of lone pairs would be 3 as the other atom is only one. Thus, the shape of  $\text{ClO}^-$  is **linear**.

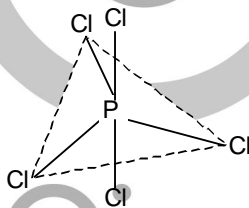


**(viii)  $\text{PCl}_5$  molecule:**

The central atom in  $\text{PCl}_5$  is P and it has 5 other Cl atoms.

$$\therefore \frac{N}{2} = \frac{5+5}{2} = 5$$

The number of lone pairs would be zero, as the number of other atoms is 5. Thus, the shape of  $\text{PCl}_5$  is **trigonal bipyramidal**.

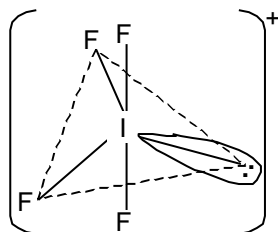


**(ix)  $\text{IF}_4^+$  ion:**

The central atom in  $\text{IF}_4^+$  ion is I and it has 4 other F atoms.

$$\therefore \frac{N}{2} = \frac{7+4-1}{2} = 5$$

Since, the number of other atoms is 4, so the number of lone pairs would be 1. Thus, the shape of  $\text{IF}_4^+$  ion is **see-saw or irregular tetrahedron**.

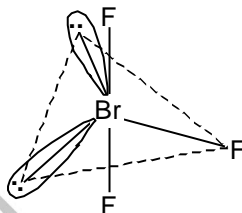


**(x)  $\text{BrF}_3$  molecule:**

In  $\text{BrF}_3$  molecule, the central atom is Br and it has 3 other F atoms.

$$\therefore \frac{N}{2} = \frac{7+3}{2} = 5$$

Since, the number of other atoms is 3, so the number of lone pairs is 2. Therefore, the shape of  $\text{BrF}_3$  is **T-shaped**.

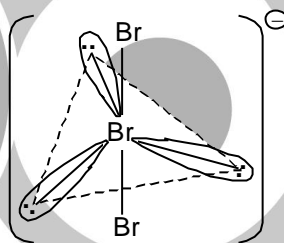


**(xi)  $\text{Br}_3^-$  ion:**

The central atom is a Br atom and it has 2 other Br atoms as surrounding atoms.

$$\therefore \frac{N}{2} = \frac{7+2+1}{2} = 5$$

Since, the number of other atoms is 2, so the number of lone pairs is 3. Thus, the shape of  $\text{Br}_3^-$  ion is **linear**.

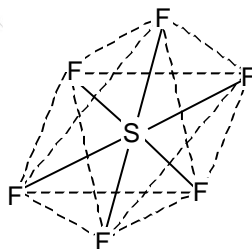


**(xii)  $\text{SF}_6$  molecule:**

The central atom in  $\text{SF}_6$  is S and it has 6 other F atoms as surrounding atoms.

$$\therefore \frac{N}{2} = \frac{6+6}{2} = 6$$

Since, the number of other atoms is 6, so the number of lone pairs are zero. Thus, the shape of  $\text{SF}_6$  molecule is **octahedron or square bipyramidal**.

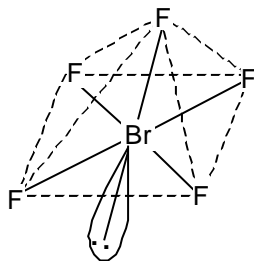


**(xiii)  $\text{BrF}_5$  molecule:**

In  $\text{BrF}_5$ , the central atom is Br and 5 F atoms are the surrounding atoms.

$$\therefore \frac{N}{2} = \frac{7+5}{2} = 6$$

Since, the number of other atoms is 5, so the number of lone pairs would be 1. Therefore, the shape of  $\text{BrF}_5$  is **square pyramidal**.

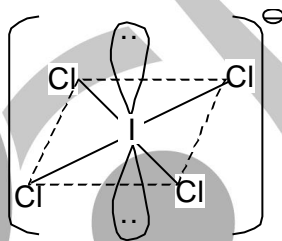


**(xiv)  $\text{ICl}_4^-$  ion:**

In  $\text{ICl}_4^-$  ion, the central atom is I and 4 Cl atoms are the surrounding atoms.

$$\therefore \frac{N}{2} = \frac{7 + 4 + 1}{2} = 6$$

The number of lone pairs would be 2 as the number of other atoms is 4. So, the shape of  $\text{ICl}_4^-$  ion is **square planar**.

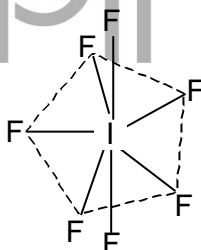


**(xv)  $\text{IF}_7$  molecule:**

In  $\text{IF}_7$  molecule, the central atom is I and 7 F atoms are the surrounding atoms.

$$\therefore \frac{N}{2} = \frac{7 + 7}{2} = 7$$

Since, the number of other atoms is 7, so the number of lone pairs would be zero. Thus, the shape of  $\text{IF}_7$  molecule is **pentagonal bipyramidal**.



Predict the shape and structure of following species.

- (a)  $\text{BF}_4^-$  (b)  $\text{SO}_3$  (c)  $\text{CO}_3^{2-}$  (d)  $\text{NO}_3^-$  (e)  $\text{PCl}_4^+$  (f)  $\text{PO}_4^{3-}$  (g)  $\text{PCl}_3$  (h)  $\text{ClO}_3^-$  (i)  $\text{OF}_2$   
(j)  $\text{POCl}_3$  (k)  $\text{SOCl}_2$  (l)  $\text{SF}_4$  (m)  $\text{ClF}_3$  (n)  $\text{XeF}_2$  (o)  $\text{XeF}_4$  (p)  $\text{XeF}_6$  (q)  $\text{PCl}_6^-$  (r)  $\text{IF}_5$   
(s)  $\text{PCl}_2\text{Br}_2\text{F}$

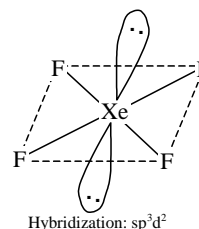
**Illustration 6**

**Question:** Write the geometry of  $\text{XeF}_4$  and  $\text{OSF}_4$  using VSEPR theory and clearly indicate the position of lone pair of electrons and hybridization of the central atom.

**Solution:**

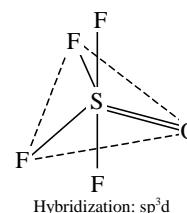
$$\text{XeF}_4: \frac{N}{2} = \frac{8+4}{2} = 6$$

There are two lone pairs. Structure is octahedral and shape is square planar.



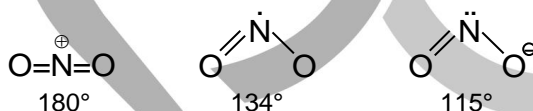
$$\text{OSF}_4: \frac{N}{2} = \frac{6+4}{2} = 5$$

Structure is irregular trigonal bipyramidal with less electronegative element occupying equatorial position. There is no lone pair.



In addition to all this, VSEPR theory can also be used to determine the geometry of other covalently bonded molecules and their bond angles. In order to predict these, following generalizations would be helpful.

1. Lone pair causes greater repulsions than a lone electron. For example,



2. Repulsions exerted by bond pairs decrease as the electronegativity of the bonded atom increases. For example,

$$\text{OH}_2 (104.5^\circ) > \text{OF}_2 (103.1^\circ)$$

$$\text{NH}_3 (107.2^\circ) > \text{NF}_3 (102.3^\circ)$$

$$\text{PI}_3 (102^\circ) > \text{PBr}_3 (101.0^\circ) > \text{PCl}_3 (100.3^\circ)$$

$$\text{AsI}_3 (101^\circ) > \text{AsBr}_3 (100.5^\circ) > \text{AsCl}_3 (98.4^\circ)$$

3. Repulsion between bonded electron pairs in filled shells is greater than those between electron pairs in incomplete shells.

$$\text{OH}_2 (104.5^\circ) \gg \text{SH}_2 (92^\circ) > \text{SeH}_2 (91^\circ) > \text{TeH}_2 (89.5^\circ)$$

$$\text{NH}_3 (107.2^\circ) \gg \text{PH}_3 (93.8^\circ) > \text{AsH}_3 (91.8^\circ) > \text{SbH}_3 (91.3^\circ)$$

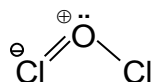
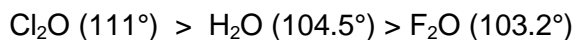
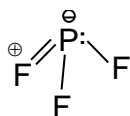
$$\text{N}(\text{CH}_3)_3, 109^\circ > \text{P}(\text{CH}_3)_3, 102.5^\circ > \text{As}(\text{CH}_3)_3, 96^\circ$$

4. When an atom with a filled valence shell & one or more lone pairs is bonded to an atom with an incomplete valence shell, or a valence shell that can become complete by electron shifts, there is a tendency for the lone pairs to be partially transferred from the filled to the unfilled shell.

$$\text{PH}_3 (93.3^\circ) < \text{PF}_3 (97.7^\circ)$$

$$\text{AsH}_3 (91.8^\circ) < \text{AsF}_3 (102^\circ)$$





5. The size of a bonding electron pair decreases with increasing electronegativity of the ligand. Also, the two electron pairs of a double bond (or the three electron pairs of a triple bond) take up more room than does the one electron pair of a single bond.

Using above facts one can rationalize size the trends which are given below:

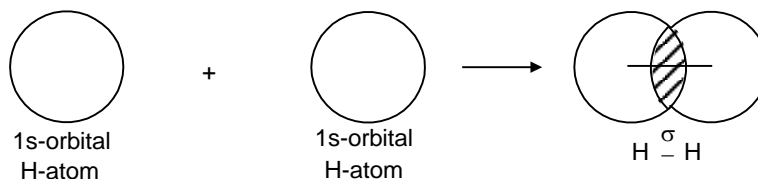
Molecule	Angles		
	XCX	XCO	XCC
F <sub>2</sub> CO	108°	126°	—
Cl <sub>2</sub> CO	111°	124°	—
(NH <sub>2</sub> ) <sub>2</sub> CO	118°	121°	—
F <sub>2</sub> SO	93°	107°	—
H <sub>2</sub> C=CF <sub>2</sub>	110°	—	125°
OPF <sub>3</sub>	103°	—	—
OPCl <sub>3</sub>	104°	—	—
H <sub>2</sub> C=CCl <sub>2</sub>	114°	—	123°

## 8. SIGMA (σ) AND PI (π) BONDS

Valence bond theory explains that a covalent bond is formed by the overlapping of the electron clouds of the atomic orbitals of the constituent atoms. The greater the overlap, the stronger the bond.

### 1. Formation of hydrogen molecule:

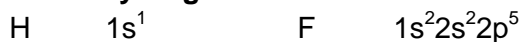
Hydrogen (1s<sup>1</sup>) has only one electron in its 1s-orbital. When two hydrogen atoms come together, overlap of their s-orbitals takes place (s-s overlap), energy is released (bond energy) and a covalent bond called the σ bond is formed.



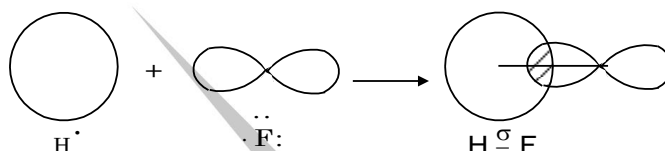
The electrons shared between the two atoms are to a large extent located in the region of space between the two nuclei. So the region of overlap is the region of high electron density. The electron density (or electron cloud) is distributed symmetrically about the bond axis, i.e., the line joining the nuclei. **Such a**

**bond formed by the axial overlapping of the orbitals is called a sigma ( $\sigma$ ) bond.**

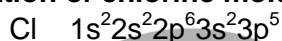
**2. Formation of hydrogen fluoride molecule:**



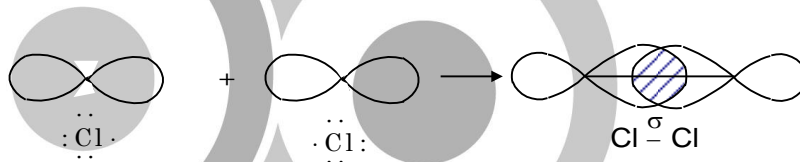
The 1s-orbital of the hydrogen atom and one of the 2p-orbitals of the fluorine atom contain only one electron each. The 1s-orbital of the hydrogen atom and the partly filled p-orbital of the fluorine atom overlap axially and form a  $\sigma$  bond (s-p overlap).



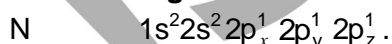
**3. Formation of chlorine molecule:**



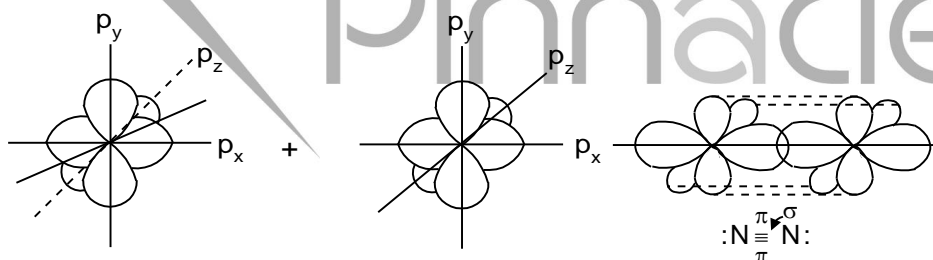
One of the 3p-orbitals of the chlorine atom contains only one electron. The half-filled p-orbital of a chlorine atom overlaps axially with the half-filled p-orbital of the other chlorine atom and forms a  $\sigma$  bond (p-p overlap).



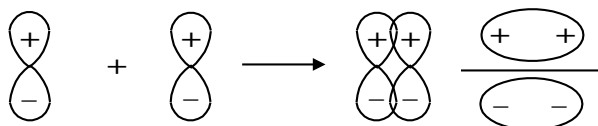
**4. Formation of nitrogen molecule:**



There are three unpaired electrons in the 2p-orbitals of a nitrogen atom. When two nitrogen atoms combine the three 2p-orbitals of one atom mutually overlap with those of the other atom and form three bonds.



Suppose the orbitals approach along the  $x$ -axis the  $p_x$ -orbitals overlap axially and form a  $\sigma$  bond (p-p overlap). The  $p_y$  and  $p_z$ -orbitals of the N atoms cannot overlap axially and so make a lateral (side to side) overlap forming two Pi ( $\pi$ ) bonds.



**Generally in any multiple bond between two atoms one bond is a  $\sigma$  bond and the others  $\pi$  bonds.** A double bond will consist of a  $\sigma$  bond and a  $\pi$  bond and a triple bond will consist of a  $\sigma$  bond and

two  $\pi$  bonds. In a  $\pi$  bond formed between two p-orbitals, the upper lobe overlaps the upper lobe and the lower lobe overlaps the lower lobe. Together they constitute  $\pi$  bond. **The  $\pi$  electron cloud will lie above and below the plane of the bond.**

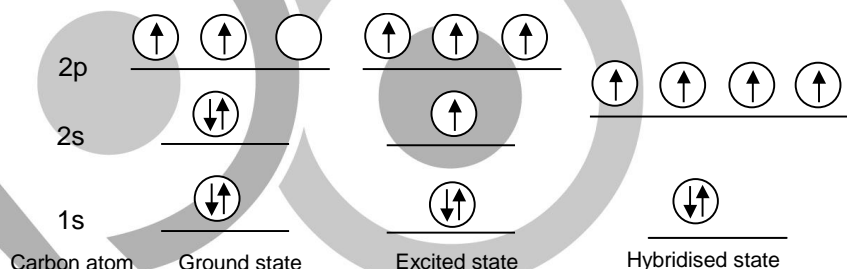
## 9. HYBRIDISATION

### 9.1 $sp^3$ HYBRIDISATION

Carbon atom has the electronic configuration  $1s^2 2s^2 2p_x^1 2p_y^1$ . It has two half-filled orbitals.

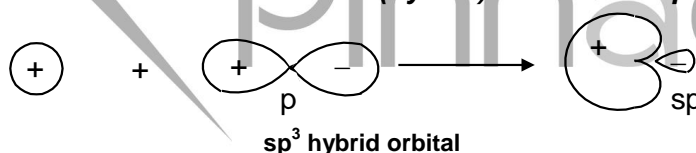
It should be expected to show a covalency of 2. In its millions of compounds carbon shows tetravalency. To explain this the concept of hybridisation is introduced. Consider the formation of methane,  $CH_4$ .

It may be supposed just for the sake of a picture that one of the electrons in the 2s-orbital is promoted to the vacant  $p_x$  orbital (excited state). This is possible because energy released during bond formation will compensate for this. Then the four orbitals, one s and three p-orbitals, get mixed up and form four new 'hybrid' orbitals, of equal energy which are called  $sp^3$  hybrid orbitals, as they are formed by the mixing up (or blending) of one s and three p-orbitals.



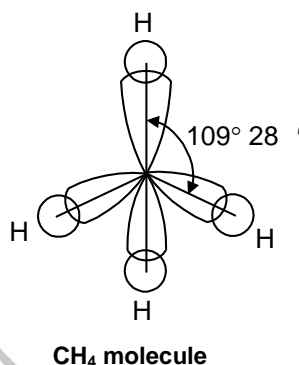
$sp^3$  hybridisation of carbon atom

So hybridisation is nothing but **combination of a certain number of atomic orbitals of slightly different energies to form the same number of new (hybrid) orbitals of equal energy.**



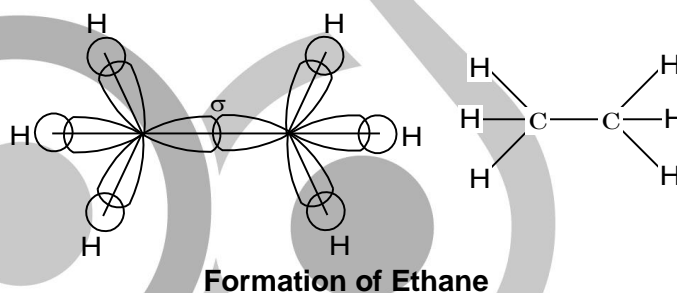
Each  $sp^3$  hybrid orbital is like a p-orbital, but with 2 lobes of unequal size (In figures, the small lobe is usually omitted).

Since the hybrid orbitals repel one another, they orient themselves with an angle of  $109^\circ 28'$  between them and point to the four corners of a regular tetrahedron. Each hybridised orbital overlaps the 1s-orbital of a hydrogen atom and forms a  $\sigma$  bond. Each  $sp^3$  hybridised orbital has one fourth s character and three-fourths p character. Note that a hybrid atomic orbital from s and p-orbitals can form only  $\sigma$  bonds. (4 C – H  $\sigma$  bonds.)



### Formation of Ethane

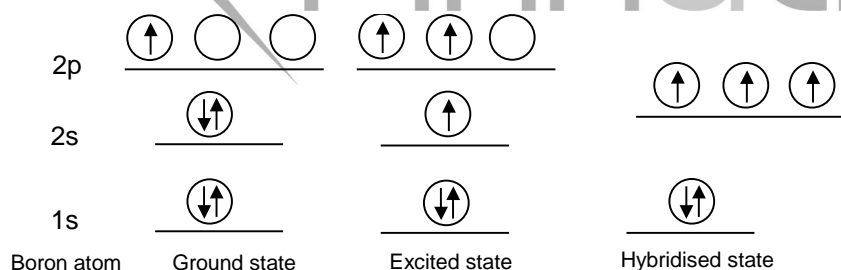
In this case there is  $sp^3-sp^3$  overlap resulting in the formation of the C – C bond and  $sp^3-s$  overlap forming C – H bonds. (1 C – C  $\sigma$  bond and 6 C – H  $\sigma$  bonds.)



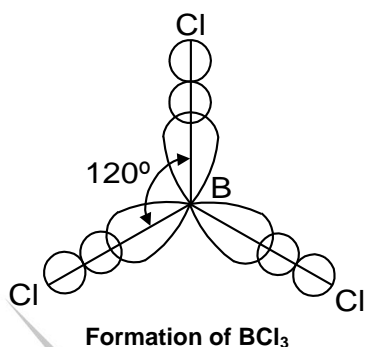
## 9.2 SP<sup>2</sup> HYBRIDISATION

### Formation of boron trifluoride/trichloride, BF<sub>3</sub>/BCl<sub>3</sub>

Boron has the electronic configuration  $1s^2 2s^2 2p^1$ . One of the s electrons is promoted to a vacant  $p_y$ -orbital (excited state). Then **one s-orbital and two p-orbitals** hybridise to form three  $sp^2$  hybrid orbitals of equivalent energy. This kind of hybridisation is called  **$sp^2$  (trigonal) hybridisation**.

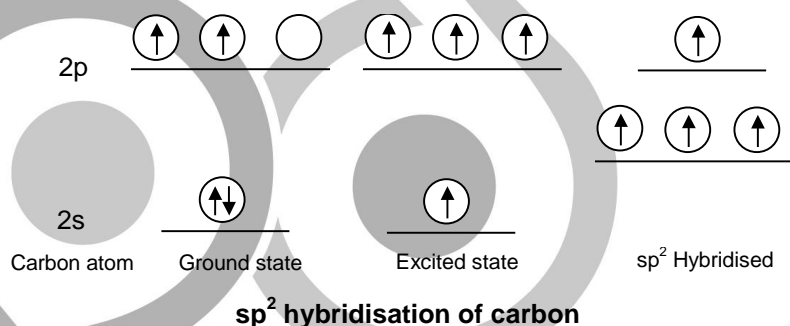


### $sp^2$ hybridisation of Boron



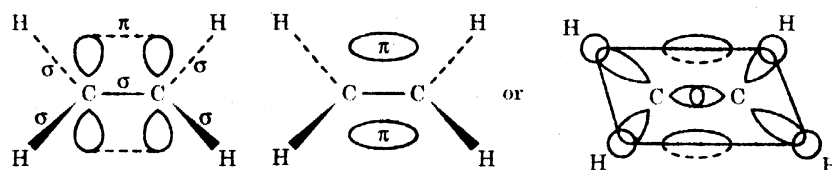
The three  $sp^2$  hybrid orbitals are **co-planar** and are at angles of  $120^\circ$  to each other. Each hybrid orbital overlaps with the vacant p-orbital of the chlorine atom and forms a  $\sigma$  bond. The other halides of Boron have similar structures. An  $sp^2$  hybrid orbital has one-third s character and two-thirds p character. (3  $\sigma$  bonds.)

### Formation of Ethylene molecule



In the formation of ethylene, carbon atom undergoes  $sp^2$  hybridisation. Two of the  $sp^2$  orbitals of each atom form  $\sigma$  bonds with 1s-orbitals of hydrogen atoms by axial overlapping.

The  $sp^2$ - $sp^2$  overlap results in the formation of a C – C  $\sigma$  bond. The two carbon atoms and the four hydrogen atoms are all in the same plane and the bond angles are  $120^\circ$ .



**Ethylene molecule**

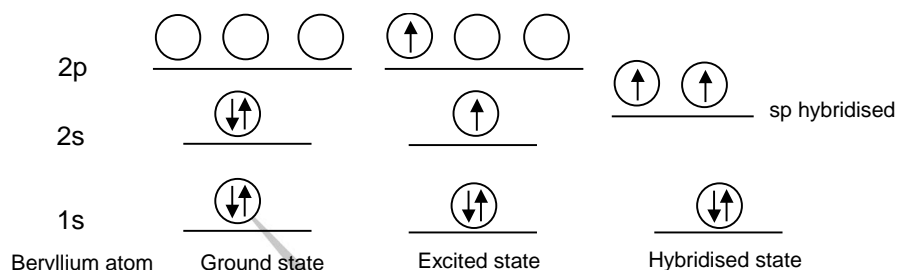
At right angles to this plane there remains the  $2p_z$  orbital of each carbon atom which overlap laterally to form a  $\pi$  bond between the two carbon atoms. The double bond between the two carbon atoms consists of a  $\sigma$  bond and a  $\pi$  bond. (4 C – H  $\sigma$  bonds, 1 C – C  $\sigma$  bond and 1  $\pi$  bond)

## 9.3 SP HYBRIDISATION

One s and one p-orbital combine to form two hybrid orbitals known as sp (or linear or diagonal) orbitals. They are of equal energy and are collinear. Each sp-orbital has one-half s character and one-half p

character.

### (i) Formation of beryllium chloride

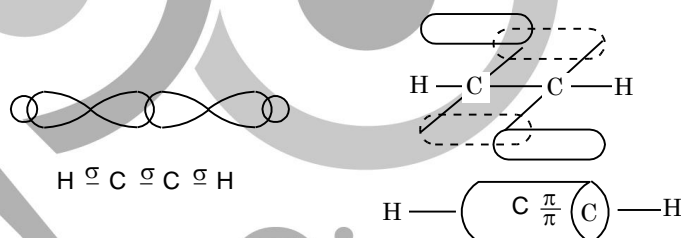


#### sp hybridization of Beryllium

The sp hybrid orbitals of beryllium atom overlap with the vacant  $p_x$ -orbitals of two chlorine atoms and form two  $\sigma$  bonds.  $\text{Cl} \sigma \text{Be} \sigma \text{Cl}$  since sp hybrid orbitals protrude along the axis farther than the corresponding s or p-orbitals they are able to overlap better and form stronger bonds than s or p-orbitals alone.

### (ii) Formation of acetylene

Hybridization of the one 2s and 2p carbon orbitals leads to the formation of two sp hybrid orbitals. sp-sp overlap between two carbon atoms form a  $\sigma$  bond between them. The other sp orbital on each carbon atom forms a  $\sigma$  bond with the 1s orbital of a hydrogen atom.



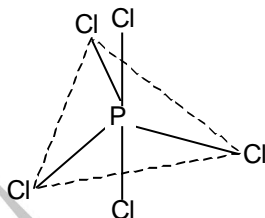
#### Acetylene molecule

Each of the carbon atom has two remaining p-orbitals which are mutually at right angles to each other. They laterally overlap and form two  $\pi$  bonds, sometimes pictured as a cylindrical sheath about the line joining the nuclei. One triple bond between 2 carbon atoms contains one C-C  $\sigma$  bond and 2 $\pi$  bonds. (1 C - C  $\sigma$  bond, 2 C - H  $\sigma$  bonds and 2 $\pi$  bonds)

The carbon-carbon triple bond is made up of one strong  $\sigma$  bond and two weaker  $\pi$  bonds; it has a total strength of 198 kcal/mole. It is stronger than a carbon-carbon double bond (163 kcal) or C-C single bond in ethane (88 kcal) and therefore is shorter than either.

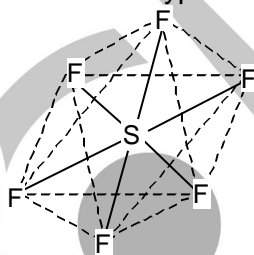
## 9.4 $sp^3d$ HYBRIDIZATION

In this type of hybridization, one 's', three 'p' and one 'd' orbitals of the same shell mix to give five  $sp^3d$  hybrid orbitals. These five  $sp^3d$  hybrid orbitals orient themselves towards the corners of a trigonal bipyramidal.



## 9.5 $sp^3d^2$ HYBRIDIZATION

In this type of hybridization, one 's', three 'p' and two 'd' orbitals of the same/different shell mix to give six  $sp^3d^2$  hybrid orbitals. These six  $sp^3d^2$  hybrid orbitals orient themselves towards the corners of an octahedron. This type of hybridization is exhibited by  $SF_6$ ,  $SCl_6$  etc.



## 9.6 SCHEME FOR DETERMINING HYBRIDIZATION OF THE CENTRAL ATOM OF A SPECIES

- (i) Identify the central atom of the species.
- (ii) Write outermost electronic configuration of the central atom.
- (iii) Determine oxidation state of the central atom.
- (iv) Excite the electrons (if necessary) to the orbitals of higher energy in order to make the number of unpaired electrons equal to the oxidation state of the central atom.
- (v) Now start putting orbitals into your pocket, beginning from 's' orbitals. The number of orbitals added to the pocket must have orbitals with unpaired electrons equal to number of other atoms of the species.
- (vi) All the orbitals added to the pocket (including s orbital, whether it has paired or unpaired electrons) are now summed. If there are one 's' and one 'p' orbital in the pocket, then the hybridization is  $sp$ . If there are one 's' and '2p' orbitals in the pocket, it is  $sp^2$  hybridization, and so on.
- (vii) Each unpaired electron left (outside the pocket) will form a  $\pi$ -bond.
- (viii) Each orbital with paired electrons in the pocket will exist as lone pair on the central atom.

The working of this scheme can be seen in the following illustrations.



**Illustration 7**

**Question:** Find out hybridization of the central atom in  $\text{ClO}_3^-$  and draw its structure.

**Solution:**

(i) The central atom is Cl.

(ii) Outer most electron configuration of Cl =  $[\text{Ne}]3s^23p^5$



(iii) Oxidation state of Cl is +5.

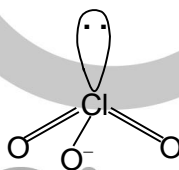
(iv) Outer most electron configuration of Cl after excitation:



(v) Now start adding orbitals into your pocket beginning from s, and thereafter 3p orbitals. We will stop adding orbitals to the pocket after adding all three '3p' orbitals because then the orbitals with unpaired electrons in the pocket would become equal to the number of other atoms in  $\text{ClO}_3^-$ .

(vi) So, the hybridization of Cl in  $\text{ClO}_3^-$  is  $sp^3$ .

(vii) Shape of  $\text{ClO}_3^-$  would be tetrahedral with one lone pair (pyramidal). Each unpaired electron left (outside the pocket) will form a  $\pi$ -bond and there will be one lone pair on Cl.



**Illustration 8**

**Question:** Find out hybridization of the central atom in  $\text{ClO}_2^-$  and draw its structure.

**Solution:**

(i) The central atom is Cl.

(ii) Outer most electron configuration of Cl =  $[\text{Ne}]3s^23p^5$



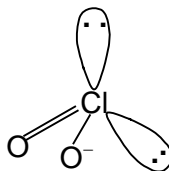
(iii) Oxidation state of Cl is +3.

(iv) Outer most electron configuration of Cl after excitation:



(v) Now start adding orbitals into your pocket beginning from s, and thereafter 3p orbitals. We will stop adding orbitals to the pocket after adding all three '3p' orbitals because then the orbitals with unpaired electrons in the pocket would become equal to the number of other atoms in  $\text{ClO}_2^-$ .

- (vi) So, the hybridization of Cl in  $\text{ClO}_2^-$  is  $\text{sp}^3$ .
- (vii) Shape of  $\text{ClO}_2^-$  would be tetrahedral with two lone pairs (angular). Each unpaired electron left (outside the pocket) will form a  $\pi$ -bond and there will be two lone pair on Cl.

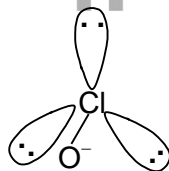


### Illustration 9

**Question:** Find out hybridization of the central atom in  $\text{ClO}^-$  and draw its structure.

**Solution:**

- The central atom is Cl.
  - Outer most electron configuration of Cl =  $[\text{Ne}]3s^2 3p^5$
- | 3s | 3p      | 3d        |
|----|---------|-----------|
| ↑↓ | ↑↓ ↑↓ ↑ | □ □ □ □ □ |
- Oxidation state of Cl is +1.
  - There is no need for excitation of electrons because the number of unpaired electrons is already equal to oxidation state of Cl.
  - Now start adding orbitals into your pocket beginning from s, and thereafter 3p orbitals. We will stop adding orbitals to the pocket after adding all three '3p' orbitals because then the orbitals with unpaired electrons in the pocket would become equal to the number of other atoms in  $\text{ClO}^-$ .
  - So, the hybridization of Cl in  $\text{ClO}^-$  is  $\text{sp}^3$ .
  - Shape of  $\text{ClO}^-$  would be tetrahedral with three lone pairs (linear). Chlorine atom will have three lone pairs.



### Illustration 10

**Question:** Find out hybridization of the central atom in  $\text{SO}_2$  and draw its structure.

**Solution:**

- The central atom is S.
- Outer most electron configuration of S =  $[\text{Ne}]3s^2 3p^4$



- Oxidation state of S is +4.

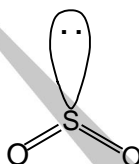
(iv) Outer most electron configuration of S after excitation:



(v) Now start adding orbitals into your pocket beginning from s, and thereafter 3p orbitals. We will stop adding orbitals to the pocket after adding two '3p' orbitals because then the orbitals with unpaired electrons in the pocket would become equal to the number of other atoms in SO<sub>2</sub>.

(vi) So, the hybridization of S in SO<sub>2</sub> is sp<sup>2</sup>.

(vii) Shape of SO<sub>2</sub> would be trigonal planar with one lone pair (angular). Each unpaired electron left (outside the pocket) will form a π-bond and there will be one lone pair on S.



### Illustration 11

**Question:** Find out hybridization of the central atom in XeOF<sub>2</sub> and draw its structure.

**Solution:**

(i) The central atom is Xe.

(ii) Outer most electron configuration of Xe is 5s<sup>2</sup>5p<sup>6</sup>



(iii) Oxidation state of Xe is +4.

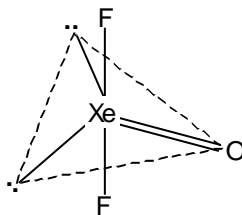
(iv) Outer most electron configuration of Xe after excitation:



(v) Now start adding orbitals into your pocket beginning from s, and thereafter 5p orbitals and then d orbitals. We will stop adding orbitals to the pocket after all three '5p' orbitals and one '5d' orbital because then the orbitals with unpaired electrons in the pocket would become equal to the number of other atoms in XeOF<sub>2</sub>.

(vi) So, the hybridization of Xe in XeOF<sub>2</sub> is sp<sup>3</sup>d.

(vii) Shape of XeOF<sub>2</sub> would be trigonal bipyramidal with two lone pairs (T shaped). Each unpaired electron left (outside the pocket) will form a π-bond and there will be two lone pairs on Xe.



### Illustration 12

**Question:** Find out hybridization of the central atom in  $\text{XeOF}_4$  and draw its structure.

**Solution:**

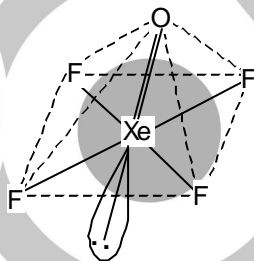
- (i) The central atom is Xe.  
(ii) Outer most electron configuration of Xe is  $5s^2 5p^6$



- (iii) Oxidation state of Xe is +6.  
(iv) Outer most electron configuration of Xe after excitation:



- (v) Now start adding orbitals into your pocket beginning from s, and thereafter 5p orbitals and then d orbitals. We will stop adding orbitals to the pocket after all three '5p' orbitals and two '5d' orbital because then the orbitals with unpaired electrons in the pocket would become equal to the number of other atoms in  $\text{XeOF}_4$ .  
(vi) So, the hybridization of Xe in  $\text{XeOF}_4$  is  $sp^3d^2$ .  
(vii) Shape of  $\text{XeOF}_4$  would be octahedron with one lone pair (square pyramidal). Each unpaired electron left (outside the pocket) will form a  $\pi$ -bond and there will be one lone pair on Xe.



## 10. HYDROGEN BOND

A hydrogen atom normally forms a single bond. In some compounds, however, the hydrogen atom may be located between two atoms acting as a bridge between them. Hydrogen atom is now involved in two bonds, one a normal covalent bond, the other a hydrogen bond. A hydrogen bond is always formed between two small, strongly electronegative atoms as fluorine, oxygen and nitrogen.

### 10.1 INTERMOLECULAR HYDROGEN BONDING – MOLECULAR ASSOCIATION

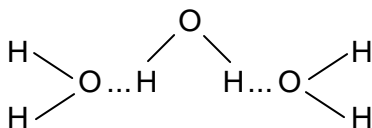
#### (i) Hydrogen fluoride:

From molecular measurements, it is known that hydrogen fluoride is associated (i.e., many molecules are joined together). HF is a polar molecule, with the fluorine atom acquiring a slight negative charge and the hydrogen atom acquiring an equal positive charge. The electrostatic attraction between the oppositely charged ends results in hydrogen bonding as shown below.

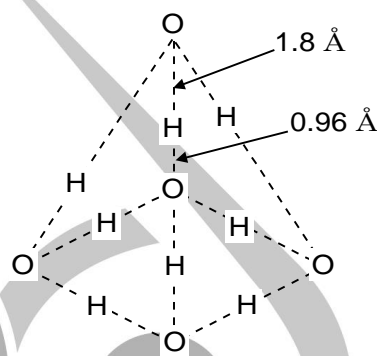


Many  $\text{H} - \text{F}$  units are held together, as  $(\text{HF})_n$ , by hydrogen bonding. The covalent  $\text{H} - \text{F}$  bond is much shorter than the  $\text{F} \cdots \text{H}$  hydrogen bond; so a hydrogen bond is much weaker than a covalent bond. Fluorine, with the highest electronegativity forms the strongest hydrogen bond. The nature of the hydrogen bond is considerably electrostatic.

- (ii) **Water:** The high boiling point compared to that of hydrogen sulphide is due to molecular association through hydrogen bonding.



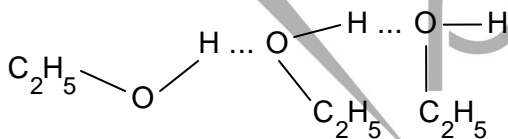
The crystal structure of ice shows a tetrahedral arrangement of water molecules. Each oxygen atom is surrounded tetrahedrally by 4 others. Hydrogen bonds link pairs of oxygen atoms together as shown in Figure 6.19. The arrangement of water molecules in ice is a very open structure and this explains the low density of ice. When ice melts, the structure breaks down and the molecules pack more closely together so that water has a higher density; this packing goes to a maximum upto a temperature of 4°C.



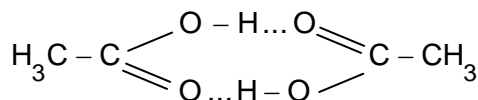
- (iii) **Ammonia** is also associated through hydrogen bonding; hence it has higher boiling point than  $\text{PH}_3$  or  $\text{AsH}_3$ .

**Note:** Methane has normal values for its melting and boiling points. It is not associated as carbon has no lone pairs and is not sufficiently electronegative to be linked by hydrogen bonds.

- (iv) **Alcohols and phenols:** Lower alcohols and phenols are associated due to intermolecular hydrogen bonding. Methanol, ethanol and phenol have relatively much higher boiling points than methane or chloromethane, ethane or chloroethane, benzene or chlorobenzene respectively.



- (v) **Carboxylic acids:** Some carboxylic acids exist as dimers e.g., the dimer of acetic acid is represented as

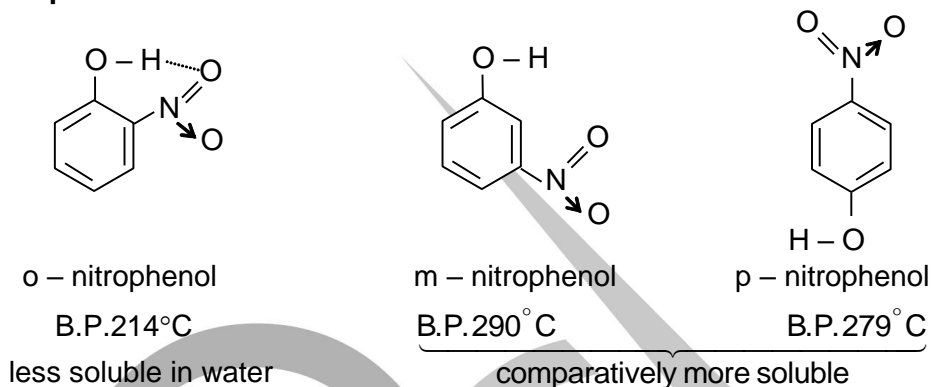


In aqueous solution molecules of a carboxylic acid link up with water molecules through hydrogen bonding rather than form dimers.

## 10.2 INTRAMOLECULAR HYDROGEN BONDING

Sometimes hydrogen bonding may take place within a molecule; this is known as intramolecular (or internal) hydrogen bonding. It may lead to the linkage of two groups to form a ring; such an effect is known as chelation, in the case of complex compounds.

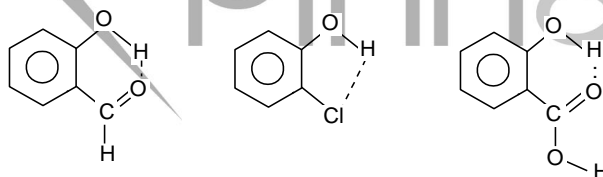
### (i) Nitrophenols



Because of the proximity of  $-OH$  and  $-NO_2$  groups in o-nitrophenol there is intramolecular hydrogen bonding which prevents intermolecular hydrogen bonding between two or more molecules. Since molecular association cannot take place, the boiling point of o-nitrophenol is lower than that of the other two. Because of the distance between  $-OH$  and  $-NO_2$  groups in m- and p-nitrophenols there is no possibility of intramolecular hydrogen bonding. Intermolecular hydrogen bonding may take place to a certain extent which causes some degree of molecular association; this is responsible for the higher boiling points of the two nitrophenols.

Further the formation of intramolecular hydrogen bonding in o-nitrophenol prevents it from entering into intermolecular hydrogen bonding with water and this explains its reduced solubility.

(ii) Other molecules in which intramolecular hydrogen bonding is present are o-hydroxybenzaldehyde, o-chlorophenol and o-hydroxybenzoic acid.



### Illustration 13

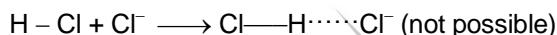
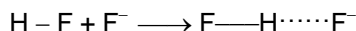
**Question:**  $H_2O$  is a liquid at ordinary temperature while  $H_2S$  is a gas although both O and S belong to the same group of the periodic table.

**Solution:**  $H_2O$  is capable of forming intermolecular hydrogen bond. This is possible due to high electronegativity and small size of oxygen. Due to intermolecular H-bonding, molecular association takes place. As a result the effective molecular weight increases and hence the boiling point increases, so  $H_2O$  exist in liquid phase. But in  $H_2S$ , no hydrogen bonding is possible due to large size and less electronegativity of S. So its boiling point is that of an isolated  $H_2S$  molecule and therefore it is a gas with low boiling point.

### Illustration 14

**Question:** The salt  $\text{KHF}_2$  is known but  $\text{KHCl}_2$  is not known. Explain.

**Solution:** The formation of  $\text{KHF}_2$  involves reaction of  $\text{HF}_2^-$  with  $\text{KOH}$ . Similar is the case with  $\text{KHCl}_2$ . So the main factor is the formation of  $\text{HF}_2^-$  or  $\text{HCl}_2^-$  ion.



Due to higher electronegativity and small size of fluorine, it is capable of forming H-bond resulting in the formation of  $\text{HF}_2^-$  and thereby  $\text{KHF}_2$  exists. But with chlorine, there is no possibility of H-bonding, so there is no possibility of existence of  $\text{KHCl}_2$ .

### Illustration 15

**Question:** o-hydroxy benzaldehyde is more volatile than p-hydroxy benzaldehyde.

**Solution:** More volatility means compound has lower boiling point. p-hydroxy benzaldehyde remains associated through intermolecular hydrogen bonding. But in o-hydroxy benzaldehyde, intramolecular H-bonding takes places, as a result of which there is no association. So p-hydroxy benzaldehyde, which remains as an associated species has got higher boiling and so less volatile while o-hydroxy benzaldehyde is highly volatile.

## 11. MOLECULAR ORBITAL THEORY AND BOND ORDER

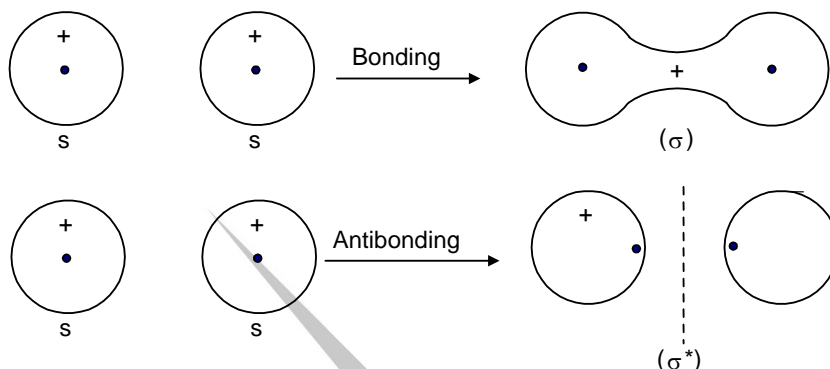
The Valence Bond Theory (V.B. Theory) with the concepts of hybridisation and resonance is used to explain the structure and properties of several molecules, but there are limitations. For example, the V.B. theory in its original form, is not able to explain the paramagnetic behaviour of  $\text{O}_2$  molecule. Hence the Molecular Orbital Theory (or M.O. Theory) due to Hund and Mulliken. The following are the essential features of the M.O. Theory.

1. In the M.O. model, all the electrons are taken together and considered as moving in the field of all the nuclei. (In the V.B. model, only the bonding electrons are considered and they are taken to move in the field of the nuclei involved in bonding.)
2. The atomic orbitals are combined to form, what are called molecular orbitals and electrons are fed into these orbitals. (In the V.B. model, electrons are fed into the atomic orbitals, which are then supposed to overlap.)
3. The number of combining atomic orbitals is equal to the number of molecular orbitals formed.
4. When two atomic orbitals combine, two M.O.'s are formed, of which one has a lower energy, while the other has a higher energy. The former is known as the bonding orbital and the latter antibonding. Mathematically, if  $\psi_1$  represents the wave function corresponding to orbital 1 and  $\psi_2$  for orbital 2, the total function is a linear combination of  $\psi_1$  and  $\psi_2$  i.e.,  $\psi = \psi_1 \pm \psi_2$  (omitting the constants). This is known as linear combination of atomic orbitals (L.C.A.O.). Of these,  $\psi_1 + \psi_2$  corresponds to the bonding M.O., while  $\psi_1 - \psi_2$  corresponds to antibonding M.O. i.e.,  $\psi_b = \psi_1 + \psi_2$  and  $\psi_a = \psi_1 - \psi_2$ . The electron density or probability of finding an electron is directly proportional to  $\psi^2$ .

For the bonding orbital,  $\psi_b^2 = (\psi_1 + \psi_2)^2 = \psi_1^2 + \psi_2^2 + 2\psi_1\psi_2$ , which is greater than  $\psi_1^2 + \psi_2^2$  i.e., the



electron density between the two nuclei is concentrated when the bonding M.O. is formed, than when no such combination of orbitals is made.



For the antibonding orbital,  $\psi_a^2 = (\psi_1 - \psi_2)^2 = \psi_1^2 + \psi_2^2 - 2\psi_1\psi_2$ , which is less than  $\psi_1^2 + \psi_2^2$  i.e. the electron density between the nuclei is withdrawn in an antibonding M.O. In the bonding M.O., since the electron density between the two nuclei is large, it holds the two nuclei together; hence the name bonding orbital; in the antibonding M.O. the bonding of the nuclei is poor.

5. There are different notations for representing bonding and anti-bonding M.O's obtained from A.O's. We give a simple notation below.

Atomic orbitals that are mixed	s and s	$p_z$ and $p_z$	$p_x$ and $p_x$	$p_y$ and $p_y$
Bonding M.O.	$\sigma s$	$\sigma(p_z)$	$\pi(p_x)$	$\pi(p_y)$
Antibonding M.O.	$\sigma^* s$	$\sigma^*(p_z)$	$\pi^*(p_x)$	$\pi^*(p_y)$

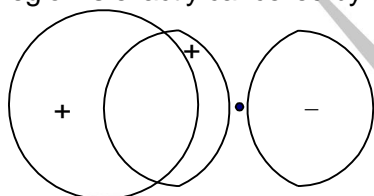
We have assumed the two  $p_z$  orbitals to overlap end to end, so that the M.O. formed is of the ' $\sigma$ ' type (similar to the ' $\sigma$ ' bond in V.B. theory); then the two  $p_x$  atomic orbitals, as also the two  $p_y$  orbitals will overlap laterally to give M.O's of the  $\pi$  type. Some authors follow the convention of choosing two  $p_x$  orbitals for end to end (i.e., axial) overlap, so that the M.O's formed are  $\sigma(p_x)$  and  $\sigma^*(p_x)$ .

6. When electrons are successively placed in the M.O's, Aufbau principle, Hund's rule and Pauli's principle are followed, as in the case of the atomic orbitals.  
Aufbau Principle: M.O's are occupied in the order of increasing energy. The following is the general arrangement of M.O's in the order of increasing energy.  
 $\sigma(1s) < \sigma^*(1s) < \sigma(2s) < \sigma^*(2s) < \sigma(2p_z) < \pi(2p_x) = \pi(2p_y) < \pi^*(2p_x) = \pi^*(2p_y) < \sigma^*(2p_z) \dots$  etc.  
 The above is only a general order and slight variations often occur due to interaction between s and p orbitals. For example, sometimes,  $\pi(2p_x) = \pi(2p_y) < \sigma(2p_z)$ .

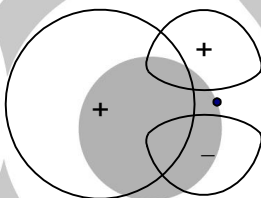
7. Hund's rule of maximum multiplicity:  
The degenerate M.O's are occupied singly, before any pairing could occur. The maximum capacity for each M.O. is 2 electrons.

8. Only atomic orbitals of equal or nearly equal energies combine to give the M.O's. In the case of homonuclear diatomic molecules, energies of corresponding A.O's of the two atoms are equal. So the above condition of combination of A.O's assumes special significance in the case of heteronuclear diatomic molecules and it has to be used with caution.

Further, for effective combination or overlap, the A.O's should have the same symmetry. Thus we have the s-s,  $p_z-p_z$ ,  $p_y-p_y$  and  $p_x-p_x$  overlaps to give bonding and antibonding orbitals as pointed out earlier (see point 5). Regarding the s-p overlap, a 2s orbital may overlap with a  $2p_z$  orbital as shown in Figure (a) below, since both have axial symmetry around the internuclear axis. However, the 2s- $2p_x$  or 2s- $2p_y$  overlap makes no contribution to bonding, as shown in Figure (b) below, where the constructive overlap in one region is exactly cancelled by the effect of the destructive overlap in the other.



(a) Enhanced amplitude



(b) Decreased amplitude

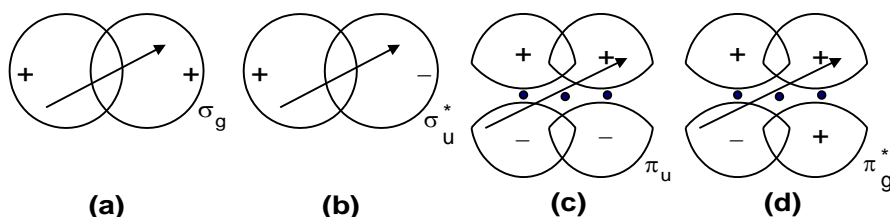
9. It has been already pointed out that electrons in the bonding M.O's tend to pull the nuclei together and that the electrons in the antibonding M.O's tend to separate them. Hence the combined influence of bonding and antibonding electrons may either stabilize or destabilize the molecule, depending on the relative number of these two types of electrons. The stabilizing power is expressed in terms of what is called the "bond order".

$$\text{Bond order} = \frac{1}{2} \left[ \left( \text{number of electrons in the bonding M.O.} \right) - \left( \text{number of electrons in the antibonding M.O.} \right) \right]$$

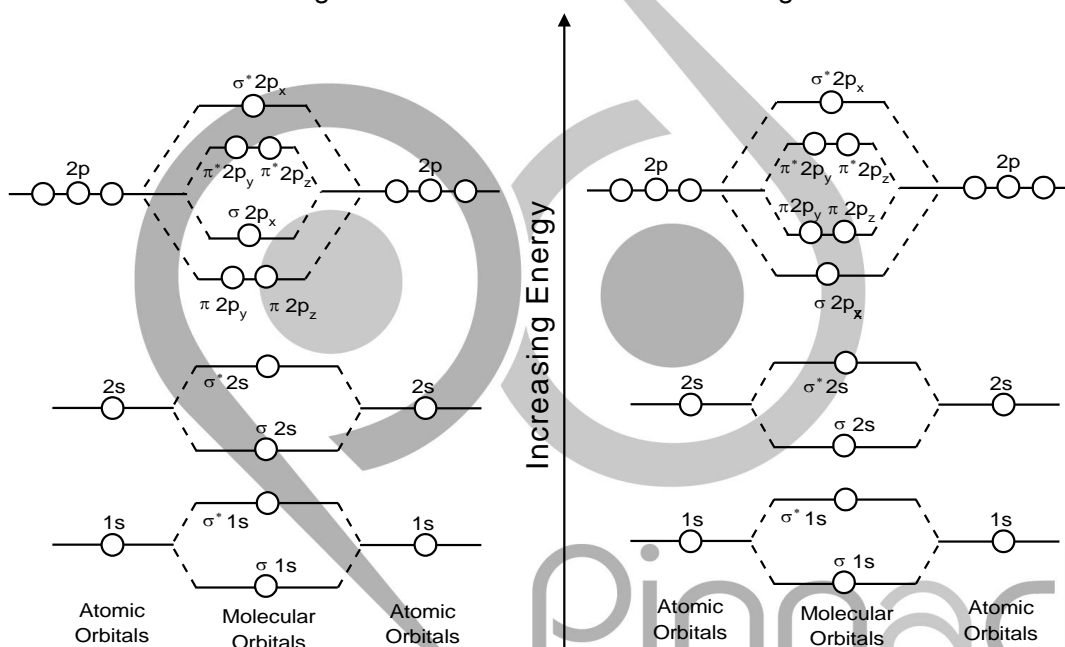
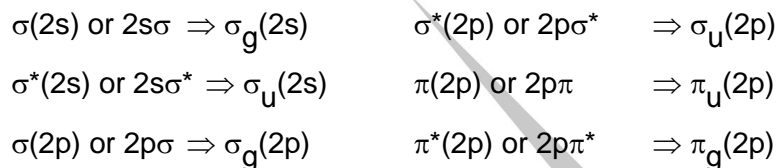
The greater the bond order, the greater the bond stability and the shorter the bond distance.

10. The M.O's are also named on symmetry grounds. For **homonuclear diatomic molecules**, the symbols 'g' ('gerade' meaning 'even') and 'u' ('ungerade' meaning 'uneven' or 'odd') are used. The symbol 'g' is used, if the orbital has a centre of symmetry. i.e., if along any straight line passing through the centre (This is called the centre of inversion), at equal distances from it, the electron densities are equal and the orbital signs are the same (i.e., the wave function has the same amplitude and sign at the two points which are opposite and equidistant from the centre). If the electron densities are equal, but the orbital signs are opposite at the two points mentioned above, the symbol 'u' is used.

In the ' $\sigma$ ' type of orbitals, the bonding orbitals are 'g' and the antibonding 'u'; in the ' $\pi$ ' type of orbitals, the bonding orbitals are 'u' and the antibonding 'g'. Figure (a) and (b) below correspond to overlap of 's' orbitals to form ' $\sigma$ ' type of M.O's; Figure (c) and (d) correspond to the formation of ' $\pi$ ' type of M.O's from 'p'-orbitals.



The equivalence of the different notations used to represent the M.O.'s is shown for a few cases (the x, y, z subscripts in the case of p-orbitals are dropped).



**For elements with  $Z \leq 7$**   
**Molecular orbital energy level diagram**

**For elements with  $Z > 7$**   
**Molecular orbital energy level diagram**

### 11.1 HOMONUCLEAR DIATOMIC MOLECULES

We shall now consider the electronic configuration of a few homonuclear diatomic molecules.

1.  $H_2$ : Electronic configuration of H atom:  $1s^1$ .

$\therefore$  in  $H_2$  molecule there are 2 electrons. M.O. configuration of  $H_2$  is  $\sigma 1s^2$ . There is no electron in antibonding M.O.  $\therefore$  Bond order (B.O.) =  $\frac{2-0}{2} = 1$

2.  $He_2$ : Electronic configuration of He atom:  $1s^2$

$\therefore$  in  $He_2$  molecule there are 4 electrons. The M.O. configuration for He is  $\sigma 1s^2, \sigma^* 1s^2$ .

$\therefore$  Bond order =  $\frac{2-2}{2} = 0$ .

$\therefore$   $He_2$  molecule is not stable.

Taking  $He_2^+$ , the structure is  $\sigma 1s^2, \sigma^* 1s^1$ .

$$\therefore \text{B.O.} = \frac{2-1}{2} = \frac{1}{2}.$$

$\therefore \text{He}_2^+$  is stabler than  $\text{He}_2$ .

3.  $\text{Li}_2$ : (Li atom:  $1s^2 2s^1$ ). Total number of electrons in  $\text{Li}_2 = 6$ . Molecular orbital configuration of  $\text{Li}_2$  is  $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^0$  or [KK]  $\sigma 2s^2, \sigma^* 2s^0$ .

where KK corresponds to filled  $\sigma(1s)$  and  $\sigma^*(1s)$  levels. The contribution of [KK] to bond order is zero. Therefore, we can ignore it and consider only the valence electrons.

$$\text{B.O.} = \frac{2-0}{2} = 1$$

$\therefore \text{Li}_2$  is stable and it is found to exist to some extent in lithium vapour.

4.  $\text{Be}_2$ : (Be:  $1s^2 2s^2$ )

Molecular orbital structure of  $\text{Be}_2$  is:

[KK]  $\sigma 2s^2, \sigma^* 2s^2$ . Ignoring [KK],

$$\text{B.O.} = \frac{2-2}{2} = 0. \therefore \text{Be}_2 \text{ is not stable.}$$

5.  $\text{B}_2$ : (B:  $1s^2 2s^2 2p^1$ )

M.O. structure of  $\text{B}_2$  is

[KK]  $\sigma 2s^2, \sigma^* 2s^2, \pi 2p_y^1, \pi 2p_z^1$ .

Note that though  $\pi(2p)$  orbital is usually more energetic than  $\sigma(2p)$ , there is an inversion of the order here due to mixing (hybridization) of  $\sigma(2s)$  and  $\sigma(2p)$  orbitals. Also note that the degenerate M.O's  $\pi(2p_y)$  and  $\pi(2p_z)$  have one electron in each according to Hund's principle.

$$\text{B.O.} = \frac{4-2}{2} = 1$$

$\therefore \text{B}_2$  is stable.

Since there are two unpaired electrons in the molecule,  $\text{B}_2$  is paramagnetic.

6.  $\text{C}_2$ : (C:  $1s^2 2s^2 2p^2$ ).

M.O. picture of  $\text{C}_2$  is:

[KK]  $\sigma 2s^2, \sigma^* 2s^2, \pi 2p_y^2, \pi 2p_z^2$

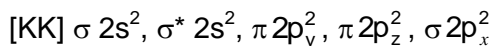
Here again  $\pi(2p)$  orbital is less energetic than  $\sigma(2p)$ .

$$\text{B.O.} = \frac{6-2}{2} = 2 \text{ and the molecule is stable.}$$

Since there is no unpaired electron,  $\text{C}_2$  is diamagnetic. [If  $\pi(p)$  orbital had been more energetic than  $\sigma(p)$ ,  $\text{C}_2$  would have been paramagnetic, which is contrary to experimental observation.]

7.  $N_2$ : (N:  $1s^2 2s^2 2p^3$ ).

M.O. picture of  $N_2$  is:



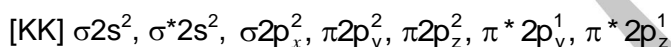
Ignoring the subscripts  $x$ ,  $y$  and  $z$  for the p-orbital and considering, only valence electrons, the representation is  $\{\sigma(2s)\}^2 \{\sigma^*(2s)\}^2 \{\pi(2p)\}^4 \{\sigma(2p)\}^2$ .

$$B.O. = \frac{8-2}{2} = 3$$

The molecule is diamagnetic. For  $N_2^+$ ,  $B.O. = \frac{7-2}{2} = 2.5$ .

8.  $O_2$ : [O:  $1s^2 2s^2 2p^4$ ]

M.O. structure of  $O_2$  is:



$$B.O. = \frac{8-4}{2} = 2$$

Due to the presence of unpaired electrons in the two antibonding orbitals (Hund's rule),  $O_2$  is paramagnetic. The M.O. theory here is superior to the V.B. theory, which does not explain the paramagnetic behaviour of  $O_2$ .

Let us now compare the bond strengths of  $O_2$ ,  $O_2^+$ ,  $O_2^-$ ,  $O_2^{2-}$ . For  $O_2$ ,  $B.O. = 2$ .

For  $O_2^+$ : one electron from the antibonding M.O. has been removed.

$$\therefore B.O. = \frac{8-3}{2} = 2.5$$

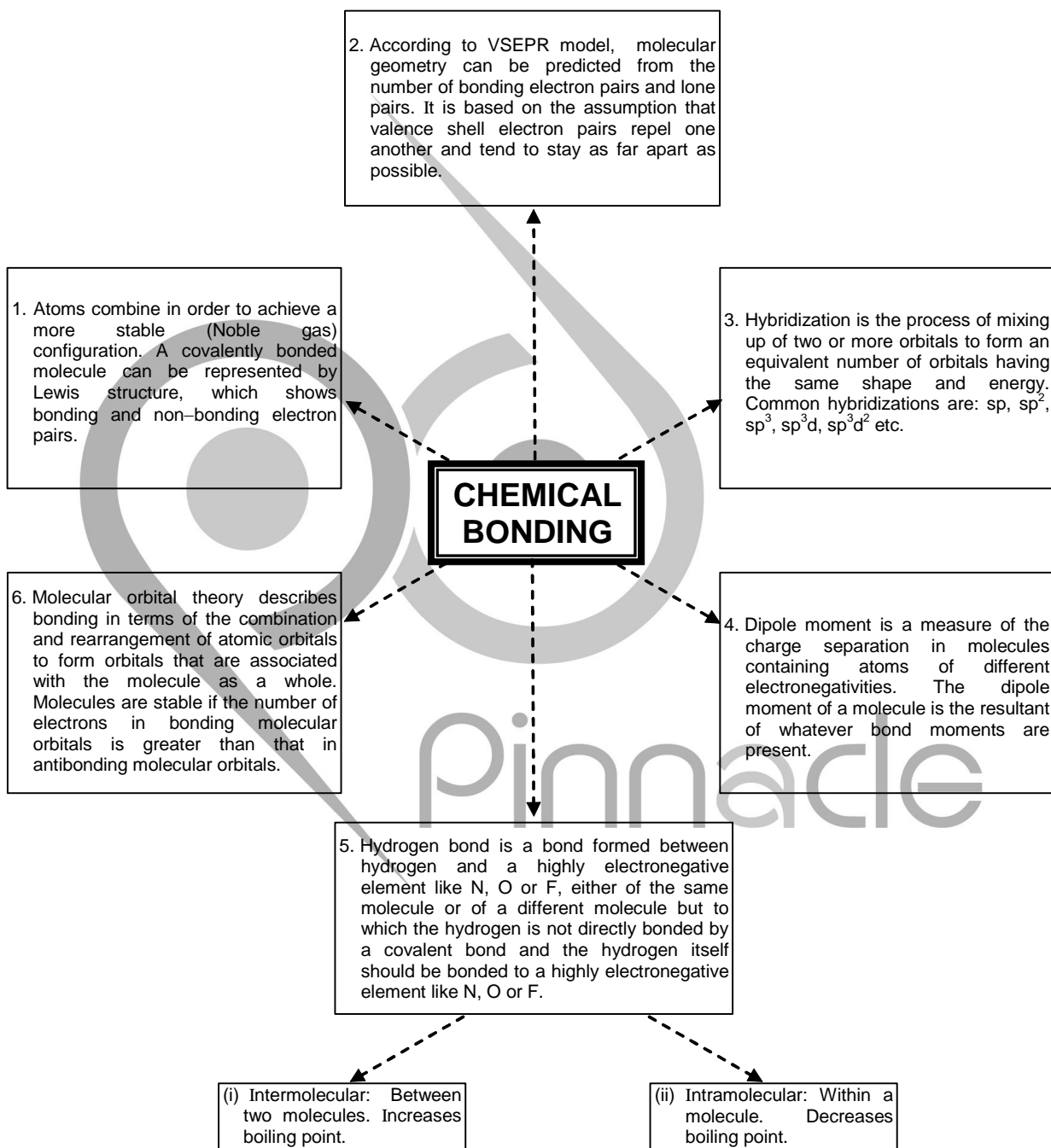
For  $O_2^-$ : one electron is added to the antibonding M.O.

$$\therefore B.O. = \frac{8-5}{2} = 1.5$$

$$\text{For } O_2^{2-}: B.O. = \frac{8-6}{2} = 1$$

$\therefore$  the bond is strongest in  $O_2^+$  and the bond length the least.

## MIND MAP



## SOLVED OBJECTIVE EXAMPLES

**Example 1:**

Which of the following molecules is linear?

- |                    |                     |
|--------------------|---------------------|
| (a) ICl            | (b) SO <sub>3</sub> |
| (c) O <sub>3</sub> | (d) SO <sub>2</sub> |

**Solution:**

A diatomic molecule is always linear. So, option (a) is definitely correct. For more than one correct answer questions one would require to check all the options. SO<sub>3</sub> is triangular planar, O<sub>3</sub> and SO<sub>2</sub> are bent.

∴ (a)

**Example 2:**

Which of the following hydride is ionic?

- |                      |                      |
|----------------------|----------------------|
| (a) H <sub>2</sub> O | (b) NH <sub>3</sub>  |
| (c) CaH <sub>2</sub> | (d) H <sub>2</sub> S |

**Solution:**

Metallic hydrides are ionic while non-metallic hydrides are covalent.

∴ (c)

**Example 3:**

A covalent molecule AB<sub>4</sub> (not a complex) will have which of the following hybridisation if it is square planar.

- |                                    |                       |
|------------------------------------|-----------------------|
| (a) sp <sup>3</sup>                | (b) sp <sup>3</sup> d |
| (c) sp <sup>3</sup> d <sup>2</sup> | (d) dsp <sup>2</sup>  |

**Solution:**

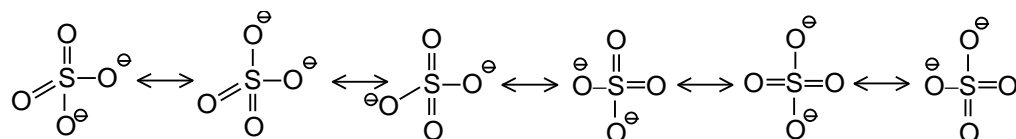
A covalent non-complex molecule does not form dsp<sup>2</sup> hybridization. AB<sub>4</sub> with sp<sup>3</sup> hybridization is tetrahedral, with sp<sup>3</sup>d is T-shaped, with sp<sup>3</sup>d<sup>2</sup> is square planar.

∴ (c)

**Example 4:**

How many types of bond lengths are there in SO<sub>4</sub><sup>2-</sup>?

- |           |          |
|-----------|----------|
| (a) one   | (b) two  |
| (c) three | (d) four |

**Solution:**


Due to resonance all bond length are same.

∴ (a)



**Example 5:**

A molecule  $XY_2$  contains two  $\sigma$ , two  $\pi$ -bonds and one lone pair in valence shell of X. The arrangement of lone pair as well as bond pair is

- (a) square pyramidal (b) linear  
 (c) trigonal planar (d) unpredictable

**Solution:**

There are a total of three bond and lone pairs. The best possible arrangement for this is trigonal planar.

$\therefore$  (c)

## SOLVED SUBJECTIVE EXAMPLES

**Example 1:**

Why does  $CS_2$  have zero dipole moment while  $H_2S$  does not?

**Solution:**

$CS_2$  is linear and symmetrical, while  $H_2S$  is bent.

**Example 2:**

Why  $C_2H_2$  is more acidic than  $C_2H_6$  although the bond energy of C-H bond in  $C_2H_2$  is more than in  $C_2H_6$ ?

**Solution:**

Acidic strength involves heterolytic cleavage while bond energy measures the energy for breaking a bond homolytically. In  $C_2H_2$ , 'C' is  $sp$  hybridized and so is more electronegative than C in  $C_2H_6$  which is  $sp^3$  hybridized.

**Example 3:**

Why is  $H_2^+$  more stable than  $H_2^-$  even though both have equal bond order?

**Solution:**

$H_2^+$  has less antibonding electron (zero) while  $H_2^-$  has more antibonding electron (one).

**Example 4:**

$CuCl$  is more covalent than  $NaCl$  even though both  $Cu^+$  and  $Na^+$  have the same size. Why?

**Solution:**

$Cu^+$  has more nuclear charge.

**Example 5:**

Why is oxygen atom in  $XeOF_2$  in equatorial plane?

**Solution:**

Out of the surrounding atoms the least electronegative atom behaves like a lone pair.

## EXERCISE – I (BUILDING A FOUNDATION)

### Section A: COVALENCY

#### SINGLE CORRECT ANSWER TYPE

1. When two atoms combine to form a molecule.
  - (a) energy is released.
  - (b) Energy is absorbed.
  - (c) Energy is neither released nor absorbed.
  - (d) Energy may release or absorb.
2. Valency expresses:
  - (a) Total number of electrons in an atom.
  - (b) Oxidation number of an element.
  - (c) Combining capacity of an element.
  - (d) No. of electrons in first shell of element.
3. Maximum covalency is observed in
  - (a) H
  - (b) O
  - (c) S
  - (d) N
4. In which of the following species, central atom is NOT surrounded by exactly 8 valence electrons?
  - (a)  $\text{BF}_4^-$
  - (b)  $\text{NCl}_3$
  - (c)  $\text{PCl}_4^+$
  - (d)  $\text{SF}_4$
5. Which statement is true about the most stable Lewis structure for  $\text{CS}_2$ ?
  - (a) There is no lone pair in molecule.
  - (b) Central atom is carbon.
  - (c) The central atom does not have an octet of electrons.
  - (d) A sulfur atom must be the central atom for the structure to be stable.
6. Which of the following species/molecules does not have same number of bond pairs and lone pairs?
  - (a)  $\text{OCN}^-$
  - (b)  $\text{H}_2\text{O}$
  - (c)  $\text{C}_2\text{H}_2\text{Cl}_2$
  - (d)  $\text{O}_3$
7. Covalency of P in  $\text{PO}_4^{3-}$  and  $\text{PH}_3$  respectively are:
  - (a) 5, 4
  - (b) 3, 5
  - (c) 3, 3
  - (d) 5, 3

#### INTEGER ANSWER TYPE

8. What covalency is exhibited by chlorine on first excitation?

9. Maximum covalency of a second period element is.....

### SUBJECTIVE TYPE

10. What are covalencies of following elements in their neutral compounds?

- |        |         |        |        |        |        |
|--------|---------|--------|--------|--------|--------|
| (1) F  | (2) Br  | (3) Xe | (4) N  | (5) P  | (6) Cl |
| (7) S  | (8) Li  | (9) Be | (10) B | (11) O | (12) C |
| (13) H | (14) Si | (15) I |        |        |        |

### LEWIS STRUCTURE

Draw the Lewis structure of following species showing lone pair and bonds:

- |                                  |                                       |                                  |                                  |                                  |
|----------------------------------|---------------------------------------|----------------------------------|----------------------------------|----------------------------------|
| (1) $\text{CO}_2$                | (2) $\text{O}_3$                      | (3) $\text{SO}_2$                | (4) $\text{SO}_3$                | (5) $\text{PCl}_5$               |
| (5) $\text{BrF}_3$               | (7) $\text{BrF}_5$                    | (8) $\text{IF}_7$                | (9) $\text{CrO}_2\text{Cl}_2$    | (10) $\text{XeOF}_2$             |
| (11) $\text{POCl}_3$             | (12) $\text{SO}_2\text{Cl}_2$         | (13) $\text{SOBr}_2$             | (14) $\text{Cl}_2\text{O}$       | (15) $\text{N}_2\text{H}_4$      |
| (16) $\text{N}_2\text{O}_3$      | (17) $\text{N}_2\text{O}_4$           | (18) $\text{N}_2\text{O}_5$      | (19) $\text{P}_4\text{S}_{10}$   | (20) $\text{S}_2\text{F}_{10}$   |
| (21) $\text{N}_2\text{F}_4$      | (22) $\text{H}_2\text{S}_2\text{O}_7$ | (23) $\text{NO}_2^-$             | (24) $\text{NO}_3^-$             | (25) $\text{N}_3^-$              |
| (26) $\text{I}_3^-$              | (27) $\text{PO}_4^{3-}$               | (28) $\text{NO}_2^+$             | (29) $\text{ICl}_2^-$            | (30) $\text{ICl}_2^+$            |
| (31) $\text{ICl}_4^-$            | (32) $\text{PCl}_4^+$                 | (33) $\text{PCl}_6^-$            | (34) $\text{ClO}^-$              | (35) $\text{ClO}_2^-$            |
| (36) $\text{ClO}_3^-$            | (37) $\text{ClO}_4^-$                 | (38) $\text{ClIBr}^-$            | (39) $\text{BrF}_2^+$            | (40) $\text{ClO}_2^+$            |
| (41) $\text{IO}_2^{5-}$          | (42) $\text{S}_2^{2-}$                | (43) $\text{SnCl}_3^-$           | (44) $\text{SO}_3^{2-}$          | (45) $\text{SO}_4^{2-}$          |
| (46) $(\text{XeF}_3^+)$          | (47) $\text{SCN}^-$                   | (48) $\text{IO}_4^-$             | (49) $(\text{NO}^+)$             | (50) $\text{P}_2\text{O}_7^{4-}$ |
| (51) $\text{P}_2\text{O}_6^{4-}$ | (52) $\text{S}_2\text{O}_7^{2-}$      | (53) $\text{S}_2\text{O}_4^{2-}$ | (54) $\text{S}_2\text{O}_6^{2-}$ | (55) $\text{S}_2\text{O}_8^{2-}$ |

### Section B: OVERLAPPING OF ATOMIC ORBITALS

#### SINGLE CORRECT ANSWER TYPE

- Which atom can not form  $\pi$  – bond?
  - O
  - N
  - C
  - H
- The fluorine molecule is formed by overlap of:
  - p – p orbitals (sideways overlap)
  - p – p orbitals (end – to – end overlap)
  - sp – sp orbitals
  - s – s orbitals
- Nuclear repulsion is minimum in which of the following head to head overlap?
  - ns – ns

- (b)  $ns - np$   
 (c)  $np - np$   
 (d)  $ns - nd$
4. choose correct order of relative  $\pi$  - bond strength:
- (a)  $2p\pi - 2p\pi > 2p\pi - 3p\pi > 2p\pi - 3d\pi$   
 (b)  $2p\pi - 2p\pi > 2p\pi - 3d\pi > 2p\pi - 3p\pi$   
 (c)  $2p\pi - 3p\pi > 2p\pi - 2p\pi > 2p\pi - 3d\pi$   
 (d)  $2p\pi - 3d\pi > 2p\pi - 3p\pi > 2p\pi - 2p\pi$
5.  $p_Y$  -Orbital can not form  $\pi$ - bond by lateral overlap with: (INA = Z)
- (a)  $d_{z^2}$ - orbital  
 (b)  $d_{x^2y^2}$ - orbital  
 (c)  $p_Y$  -Orbital  
 (d)  $p_Z$  -Orbital
6. which of the following combination of orbitals do/does not form bond (if Z - axis is internuclear axis)?
- (a)  $s + p_Y$   
 (b)  $s + s$   
 (c)  $p + p_Z$   
 (d)  $p_Z + s$

### COMPREHENSION TYPE

#### Paragraph for Question no. 7 to 9

According to V.B.T., atoms of element form bond only to pair up their unpaired electrons present in ground state or excited state. This pairing of unpaired electron will take place by overlapping of orbitals each one having one unpaired electron with opposite spin.

7. Which of the following orbital combination does not form bond? (Z = internuclear axis)
- (a)  $p_x + p_x$  (sideways overlapping)  
 (b)  $d_{x^2y^2} + p_Y$  (sideways overlapping)  
 (c)  $d_{xy} + d_{xy}$  (sideways overlapping)  
 (d)  $d_{yz} + p_Y$  (sideways overlapping)
8. Which orbital can not form  $\pi$  - bond?
- (a)  $p_x$   
 (b)  $p_Y$   
 (c)  $s$   
 (d)  $d_{z^2}$
9. The strength of  $\sigma$  - bonds formed by  $2s - 2s$ ,  $2p - 2p$  and  $2p - 2s$  overlap has the order:
- (a)  $s - s > p - p > p - s$   
 (b)  $s - s > p - s > p - p$   
 (c)  $p - p > p - s > s - s$   
 (d)  $p - p > s - s > p - s$

#### Paragraph for Question no. 10 and 11

According to valence bond theory for a covalent bond formation, two atoms must come closer to each other

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so that orbitals of one atom overlaps with the other. Overlapping orbitals must have unpaired electron. Greater is the overlapping extent, lesser will be the bond length, more will be attraction and more will be bond energy and bond strength.

10. When two  $P_z$  orbitals overlap in a particular direction and form  $\sigma$  – bond, if two  $d_{yz}$  orbitals overlap in same direction then which of the following type of covalent bond should be formed:

- (a)  $\sigma$  – bond
- (b)  $\pi$  – bond
- (c)  $\delta$  – bond
- (d) either  $\sigma$  or  $\pi$  bond

11. which combination of overlap is non – bonding? (INA = Z)

- (a)  $s + P_z$
- (b)  $s + P_y$
- (c)  $s + P_x$
- (d)  $P_z + P_z$

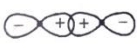
### Matrix Match Type

12. column – I

(Overlapping)



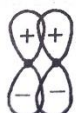
(a)



(b)



(c)



(d)

Column – II

(Result, type of overlap)

(p)  $\sigma$  – bond

(Q)  $\pi$  – bond

(R)  $\delta$  – bond

(s) Axial overlap

(T) Sidewise overlap

### INTEGER ANSWER TYPE

13. How many second period elements can form  $d\pi - d\pi$  bond?

14. How many given elements can form  $\pi$  – bond with other atom of itself?

C, N, O, F, H, He, Li, Na

15. What is minimum principal quantum number of electron that can undergo sidewise overlap to form  $\pi$  – bond?

### Section C: HYBRIDISATION – I

#### SINGLE CORRECT ANSWER TYPE

1. In which of the following, the state of hybridization of the central atom is not same as in the others:

- (a) B in  $BF_3$
- (b) O in  $H_3O^+$

- (c) N in  $\text{NH}_3$   
 (d) P in  $\text{PCl}_3$
2. Which bonds are formed by a carbon atom with  $\text{sp}^2$ -hybridisation?  
 (a) 4  $\pi$  –bonds  
 (b) 2  $\pi$  –bonds and 2  $\sigma$  –bonds  
 (c) 1  $\pi$  –bond and 3  $\sigma$  – bonds  
 (d) 4  $\sigma$  –bonds
3. Which hybrid orbital always forms  $\pi$  –bond?  
 (a)  $\text{sp}$   
 (b)  $\text{sp}^2$   
 (c)  $\text{sp}^3$   
 (d) None
4. In the compound:  
 $\overset{1}{\text{C}}\text{H}_2 = \overset{2}{\text{C}}\text{H}_2 - \overset{3}{\text{C}}\text{H}_2 - \overset{4}{\text{C}}\text{H}_2 - \overset{5}{\text{C}} \equiv \overset{6}{\text{C}}\text{H}$ , the  $\text{C}_2 - \text{C}_3$  bond is formed by the overlapping of:-  
 (a)  $\text{sp} - \text{sp}^2$   
 (b)  $\text{sp}^3 - \text{sp}^3$   
 (c)  $\text{sp} - \text{sp}^3$   
 (d)  $\text{sp}^2 - \text{sp}^3$

### MULTIPLE CORRECT ANSWER TYPE

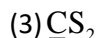
5. In which species, the hybrid state of central atom is/are  $\text{sp}^3\text{d}$ ?  
 (a)  $\text{I}_3^-$   
 (b)  $\text{SF}_4$   
 (c)  $\text{PF}_5$   
 (d)  $\text{IF}_5$
6. Solid  $\text{PCl}_5$  has hybridization:  
 (a)  $\text{sp}^3$   
 (b)  $\text{sp}^3\text{d}$   
 (c)  $\text{sp}^3\text{d}^2$   
 (d)  $\text{sp}^2$

### INTEGER ANSWER TYPE

7. Central atom of  $\text{XeO}_2\text{F}_2$  possesses how many hybrid orbitals?  
 8. How many pair of orbitals do not mix together to form hybrid orbital?  
 (2s, 2p), (1s, 3p) (4s, 4p, 4d), (3s, 3p, 2p), (3d, 4d), (3s, 3p), (3s, 3p, 3d), (1s, 3s), (2s, 4p)  
 9. Hybrid orbital containing lone pair of electrons is/are present in how many given species?  
 $\text{NO}_2^-$ ,  $\text{PCl}_3$ ,  $\text{H}_2\text{O}$ ,  $\text{NF}_3$ ,  $\text{SiH}_4$ ,  $\text{C}_2\text{H}_2$ ,  $\text{NH}_4^+$ ,  $\text{ICl}_2^-$ ,  $\text{XeF}_2$

### SUBJECTIVE TYPE

10. Find hybridisation of underlined atoms in following species:



- |   |   |  |   |  |
|---|---|--|---|--|
| (6) $\underline{\text{S}}\text{O}_2$  | (7) $\underline{\text{S}}\text{F}_4$  | (8) $\underline{\text{S}}\text{O}_3$   | (9) $\underline{\text{N}}\text{O}_2^-$  | (10) $\underline{\text{N}}\text{O}_2^+$  |
| (11) $\underline{\text{N}}\text{O}_3^-$   | (12) $\underline{\text{C}}\text{O}_3^{2-}$                                      | (13) $\underline{\text{S}}\text{O}_3^{2-}$   | (14) $\underline{\text{S}}\text{O}_4^{2-}$  | (15) $\underline{\text{X}}\text{eF}_2$   |
| (16) $\underline{\text{X}}\text{eF}_4$  | (17) $\underline{\text{I}}\text{Cl}_2^-$  | (18) $\underline{\text{I}}\text{Cl}_4^+$   | (19) $\underline{\text{I}}\text{Cl}_4^-$  | (20) $\underline{\text{I}}\text{Cl}_2^+$ |
| (21) $\underline{\text{X}}\text{eOF}_2$   | (22) $\underline{\text{X}}\text{eO}_3\text{F}_2$                                | (23) $\underline{\text{X}}\text{eO}_2\text{F}_2$   | (24) $\underline{\text{X}}\text{eOF}_4$   | (25) $\underline{\text{X}}\text{eO}_3$   |
| (26) $\underline{\text{X}}\text{eO}_4$  | (27) $\underline{\text{X}}\text{eO}_6^{4-}$                                     | (28) $\underline{\text{H}}\text{CN}$   | (29) $\underline{\text{S}}\text{F}_6$   | (30) $\underline{\text{S}}\text{OCl}_2$  |
| (31) $\underline{\text{H}}\text{O}\underline{\text{C}}\underline{\text{I}}\text{O}_3$                             | (32) $\underline{\text{P}}\text{O}(\underline{\text{O}}\underline{\text{H}})_3$ | (33) $\underline{\text{C}}\text{H}_3 - \underline{\text{C}}\text{H}_2 - \underline{\text{C}}\text{I}$  | (34) $\underline{\text{C}}\text{H}_3 - \underline{\text{C}}\text{H} = \underline{\text{C}}\text{H}_2$ |  |
| (35) $\underline{\text{C}}\text{H}_3 - \underset{\text{O}}{\parallel} - \underline{\text{O}}\underline{\text{H}}$ | (36) $\underline{\text{H}}\underline{\text{C}}\text{OOH}$                       | (37) $\underline{\text{H}} - \underline{\text{C}} \equiv \underline{\text{C}} - \underline{\text{C}}\text{H}_2 - \underline{\text{C}}\text{H}_3$ | (38) $\underline{\text{P}}\underline{\text{C}}\text{I}_4^+$   |  |
| (39) $\underline{\text{P}}\underline{\text{C}}\text{I}_6^-$   | (40) $\underline{\text{B}}\underline{\text{F}}_4^-$                             |  |   |  |

## Section D: HYBRIDISATION - II

### SINGLE CORRECT ANSWER TYPE

- Which species has the same electron geometry as  $\text{NH}_3$ ?
  - $\text{SO}_3^{2-}$
  - $\text{CO}_3^{2-}$
  - $\text{NO}_3^-$
  - $\text{SO}_3$
- Among given species identify the iso-structural pairs:
  - $[\text{NF}_3 \text{ and } \text{NO}_3^-]$
  - $[\text{BF}_4^- \text{ and } \text{NH}_4^+]$
  - $[\text{BCl}_2 \text{ and } \text{BrCl}_3]$
  - $[\text{NF}_3 \text{ and } \text{NO}_3^-]$
- Which carbon is most electronegative?
  - $\text{sp}^3$  hybridised carbon
  - $\text{sp}$  hybridised carbon
  - $\text{sp}^2$  hybridised carbon
  - $\text{sp}^3\text{d}$  hybridised carbon.
- The structure of  $\text{IF}_7$  is:-
  - Octahedral
  - Pentagonal bipyramid
  - Square pyramid
  - Trigonal bipyramid
- Percentage p - character is three times to percentage s - character in which hybrid orbital?
  - $\text{sp}$
  - $\text{sp}^2$
  - $\text{sp}^3$



(d)  $dsp^2$

### MATRIX MATCH TYPE

#### 6. Column - I (Electron geometry)

- (P) Linear  
(Q) Octahedral  
(R) Trigonal bipyramidal  
(S) Trigonal planar

	P	Q	R	S
(a)	3	2	4	1
(b)	4	1	2	3
(c)	4	1	2	3
(d)	2	3	1	4

#### Column - II (Hybridisation)

- (1)  $sp^3d^2$   
(2)  $sp^3d$   
(3)  $sp^2$   
(4)  $sp$

#### 7. Column - I (Hybridization)

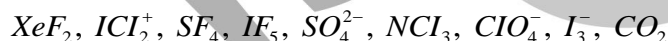
- (a)  $sp^3d$  (Trigonal bipyramidal)  
(b)  $sp^3d^2$  (octahedral)  
(c)  $sp^3d^3$  (Pentagonal bipyramidal)  
(d)  $sp^3$  (Tetrahedral)

#### Column - II (Used orbitals)

- (P)  $d_{xy}$   
(Q)  $d_{z^2}$   
(R)  $d_{x^2-y^2}$   
(S)  $p_z$   
(T)  $p_x$

### INTEGER ANSWER TYPE

8. How many axial orbital (s) are used in hybridization of central atom in  $PCl_5$ ?  
9. How many non-planar species use axial set of d-orbitals for hybridization?



10. Bond angle in a tetratomic species is  $120^\circ$ . Minimum how many hybrid orbitals can be present in it

### Section E: VSEPR THEORY - I

#### SINGLE CORRECT ANSWER TYPE

1. The geometrical arrangement of orbitals and shape of  $I_3^-$  are respectively:  
(a) Trigonal bipyramidal geometry, linear shape  
(b) Hexagonal geometry, T-shape  
(c) Triangular planar geometry, triangular shape  
(d) Tetrahedral geometry, pyramidal shape
2. Which of the following pairs of species have identical shapes?  
(a)  $NO_2^+$  and  $NO_2^-$   
(b)  $PCl_5$  and  $BrF_5$   
(c)  $XeF_4$  and  $ICl_4^-$   
(d)  $TeCl_4$  and  $XeO_4$
3. The shapes of  $XeF_4$  and  $XeF_5^-$  and  $SnCl_2$  are:  
(a) Octahedral, trigonal bipyramidal and bent  
(b) Square pyramidal, pentagonal planar and linear

- (c) Square planar, pentagonal planar and angular  
(d) See - saw, T - shaped and linear
4. Give the correct order of initials T or F for following statements. Use T if statement is true and F if it is false:
- (I) The order of repulsion between different pair of electrons is  $l_p - l_p > l_p - b_p > b_p - b_p$   
(II) In general, as the number of lone pair of electrons on central atom increases, value of bond angle also increases  
(III) The number of lone pair on O in  $H_2O$  is 2 while on N in  $NH_3$  is 1
- (a) TTT  
(b) TFT  
(c) FTF  
(d) TFF
5.  $SbF_5$  reacts with  $XeF_4$  and  $XeF_6$  to form ionic compounds  $[XeF_3^+][SbF_6^-]$  and  $[XeF_5^+][SbF_6^-]$  then molecular shape of  $[XeF_3^+]$  ion and  $[XeF_5^+]$  ion respectively are:
- (a) Square pyramidal, T - shaped  
(b) Bent - T - shape, square pyramidal  
(c) See - saw, square pyramidal  
(d) Square pyramidal, see - saw

### MULTIPLE CORRECT ANSWER TYPE

6. Which of the following is (are) V - shaped?
- (a)  $S_3^{2-}$   
(b)  $ClO_2^-$   
(c)  $ClO_2^+$   
(d)  $SO_2$
7.  $CO_2$  molecule is not isostructural with:
- (a)  $HgCl_2$   
(b)  $SnCl_2$   
(c)  $C_2H_2$   
(d)  $NO_2$

### INTEGER ANSWER TYPE

8. How many species are linear shaped among given?  
 $CS_2$ ,  $S_3^{2-}$ ,  $SO_2$ ,  $N_2O$ ,  $NO_2^+$ ,  $N_3^-$ ,  $ICl_2^-$ ,  $NO_3^-$ ,  $C_2H_2$
9. Lone pair occupies axial position in how many given species?  
 $ClF_3$ ,  $XeF_2$ ,  $I_3^-$ ,  $SF_4$ ,  $XeF_5^-$ ,  $XeO_2F_2$ ,  $XeOF_2$ ,  $XeO_3$ ,  $PCl_3$ ,
10. Predict shape of following species:
- |              |                 |                 |                 |               |
|--------------|-----------------|-----------------|-----------------|---------------|
| (1) $H_2O$   | (2) $NO_2^-$    | (3) $NO_3^-$    | (4) $NO_2^+$    | (5) $PCl_4^+$ |
| (6) $NH_4^+$ | (7) $SO_3^{2-}$ | (8) $SOCl_2$    | (9) $PF_6^-$    | (10) $N_3^-$  |
| (11) $XeF_2$ | (12) $XeOF_2$   | (13) $XeO_2F_2$ | (14) $XeO_3F_2$ | (15) $XeO_4$  |
| (16) $XeF_4$ | (17) $XeO_3$    | (18) $XeF_3^+$  | (19) $ClO_2^-$  | (20) $OF_2$   |

- |                  |                |                  |                |               |
|------------------|----------------|------------------|----------------|---------------|
| (21) $I_3^-$     | (22) $ICI_2^+$ | (23) $ICI_4^-$   | (24) $ICI_4^+$ | (24) $SF_4$   |
| (26) $BF_3$      | (27) $XeF_5^-$ | (28) $BrF_5$     | (29) $HCN$     | (30) $SO_2$   |
| (31) $SO_4^{2-}$ | (32) $N_2O$    | (33) $PO_4^{3-}$ | (34) $PCl_3$   | (35) $I_3^+$  |
| (36) $IF_7$      | (37) $ClO_2^+$ | (38) $C_3O_2$    | (39) $O_3$     | (40) $HgCl_2$ |

## Section F: VSEPR THEORY – II

### SINGLE CORRECT ANSWER TYPE

- In which of the following species maximum number of atoms lie in one plane?
  - $XeF_2O_2$
  - $PCl_5$
  - $AsH_4^+$
  - $XeF_4$
- Which molecular geometry is least likely to result from a trigonal bipyramidal electron geometry?
  - Trigonal planar
  - See – saw
  - Linear
  - T – shaped
- Which of the following statements is correct in the context of the allene molecule,  $C_3H_4$ ?
  - The central carbon is  $sp$  hybridized.
  - The terminal carbon atoms are  $sp^2$  hybridized.
  - The planes containing the  $CH_2$  groups are mutually perpendicular to permit the formations two separate  $\pi$  – bonds.
  - All the correct.

### MULTIPLE CORRECT ANSWER TYPE

- Which of the following combination of bond pair (b.p.) and lone pair (l.p.) give same shape.
 

(i) 3b.p. + 1 l.p.	(ii) 2 b.p. + 2 l.p.	(iii) 2 b.p. + 1 l.p.	(iv) 2 b.p. + 0 l.p.
(v) 3 b.p. + 2 l.p.	(vi) 2 b.p. + 3 l.p.		

  - (ii) and (iii)
  - (iv) and (v)
  - (iv) and (vi)
  - (iii) and (vi)
- If  $AB_n^{\pm}$ , type species are tetrahedral, then which of the following is/are correctly matched? (where A is central atom, B is surrounding atom and n is charge on species.)
 

A	B	n
(a) Xe	O	0
(b) Se	F	0
(c) P	O	- 3
(d) N	H	+ 1

## COMPREHENSION TYPE

### Paragraph for Question no. 6 and 7

According to VEPER model, molecules adopt geometries in which their valence electron pair position themselves as far from each other as possible. The VSEPR model considers double and triple bonds to have slightly greater repulsive effects than single bonds because of the repulsive effect of  $\pi$  – electrons. The lone pair of electrons occupies more space around the central atom

6. Which of the following statement is false?
  - (a)  $\text{SbF}_4^-$  and  $\text{SF}_4$  are isostructural.
  - (b) In  $\text{IOF}_5$  the hybridization of central atom is  $\text{sp}^3\text{d}^2$
  - (c) Double bond (s) in  $\text{SOF}_4$  and  $\text{XeO}_2\text{F}_2$ , is/are occupying equatorial position(s) of their respective geometry.
  - (d) None of these
7. Which of the following does not represent the pair of species having same shape?
  - (a)  $\text{SF}_5^-$  and  $\text{IF}_5$
  - (b)  $\text{XeO}_3\text{F}_2$  and  $\text{PCl}_5$
  - (c)  $\text{SeF}_3^+$  and  $\text{XeO}_3$
  - (d)  $\text{NO}_3^-$ ,  $\text{SO}_3^{2-}$

### Paragraph for Question No. 8 and 9

The space model is obtained by joining the points representing various bonded atoms gives due shape of the molecule. The geometry of the molecule is definite relative arrangement of the bonded pair repulsion theory by Gillespie and Nyholm.

8. Which molecules have identical shape and hybridization of central atom?
 

(I)  $\text{SnCl}_2$     (II)  $\text{NH}_3$     (III)  $\text{PCl}_6^-$     (IV)  $\text{SF}_6$

  - (a) I, II & IV
  - (b) II, III & IV
  - (c) III & IV
  - (d) All
9. Which electron geometry of covalent molecules has not been observed?
  - (a) Pentagonal bipyramidal
  - (b) Octahedral
  - (c) Hexagonal
  - (d) Tetrahedral

## MATRIX MATCH TYPE

### 10. Column – I

(Species)

- (a)  $\text{AsO}_4^{3-}$
- (b)  $\text{ICl}_2^+$
- (c)  $\text{SOF}_4$
- (d)  $\text{XeOF}_4$

### Column – II

(Property)

- (P) All the three p – orbitals of valance shell are used in hybridization
- (Q) Tetrahedral shape
- (R) Axial d – orbital with two nodal cones used in hybridization
- (S) All bond lengths are equal

(T)  $p\pi - d\pi$  (s) present

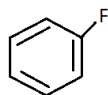
## Section G: RESONANCE AND EQUIVALENT HYBRID ORBITALS

### SINGLE CORRECT ANSWER TYPE

- The correct order of Cl – O bond order is:
  - $\text{ClO}_3^- < \text{ClO}_4^- < \text{ClO}_2^- < \text{ClO}^-$
  - $\text{ClO}^- < \text{ClO}_4^- < \text{ClO}_3^- < \text{ClO}_2^-$
  - $\text{ClO}^- < \text{ClO}_2^- < \text{ClO}_3^- < \text{ClO}_4^-$
  - $\text{ClO}_4^- < \text{ClO}_3^- < \text{ClO}_2^- < \text{ClO}^-$
- The correct order of increasing C – O bond strength of  $\text{CO}$ ,  $\text{CO}_3^{2-}$ ,  $\text{CO}_2$  is:
  - $\text{CO}_3^{2-} < \text{CO}_2 < \text{CO}$
  - $\text{CO}_2 < \text{CO}_3^{2-} < \text{CO}$
  - $\text{CO} < \text{CO}_3^{2-} < \text{CO}_2$
  - $\text{CO} < \text{CO}_2 < \text{CO}_3^{2-}$
- N – O bond length is maximum in:
  - $\text{NO}_2^+$
  - $\text{NO}_3^-$
  - $\text{NO}^+$
  - $\text{NO}_2^-$
- All bonding orbitals of central atom are equivalent in:
  - $\text{CH}_3\text{F}$
  - $\text{BF}_2\text{Cl}$
  - $\text{XeO}_3$
  - None of these

### MULTIPLE CORRECT ANSWER TYPE

- All hybrid orbitals of central atom are equivalent in:
  - $\text{NH}_3$
  - $\text{CO}_2$
  - $\text{NO}_3^-$
  - $\text{O}_3$
- Long pair is delocalized in:



- 
- $\text{SO}_4^{2-}$
- $\text{NH}_4^+$
- $\text{N}_3^-$

### INTEGER ANSWER TYPE

- All hybrid orbitals of central atom are equivalent in how many species?  
 $\text{XeO}_4$ ,  $\text{NO}_2^-$ ,  $\text{NO}_2^+$ ,  $\text{SO}_3^{2-}$ ,  $\text{XeOF}_4$ ,  $\text{NF}_3$ ,  $\text{CH}_2\text{F}_2$ ,  $\text{SO}_4^{2-}$ ,  $\text{NO}_3^-$

8.  $\pi$  – bond order in  $\text{CO}_3^{2-}$  is  $x$ . The value of  $3x$  is nearly equal to....
9. Long pair of oxygen is delocalised in how many given species?  
 $\text{NO}_2^-$ ,  $\text{NO}_3^-$ ,  $\text{ClO}_4^-$ ,  $\text{O}_3$ ,  $\text{CH}_3$ ,  $\text{OH}$ ,  $\text{HCOO}^-$ ,  $\text{CH}_3$ ,  $\text{COOH}$ ,  $\text{CO}_3^{2-}$ ,  $\text{PO}_4^{3-}$

### SUBJECTIVE TYPE

10. Find bond order in following species:

- |                        |                      |                     |                         |                        |                      |
|------------------------|----------------------|---------------------|-------------------------|------------------------|----------------------|
| (1) $\text{CO}_3^{2-}$ | (2) $\text{NO}_2^+$  | (3) $\text{NO}_3^-$ | (4) $\text{NO}_2^-$     | (5) $\text{PO}_4^{3-}$ | (6) $\text{ClO}_4^-$ |
| (7) $\text{SO}_4^{2-}$ | (8) $\text{ClO}_3^-$ | (9) $\text{O}_3$    | (10) $\text{SO}_3^{2-}$ |                        |                      |

### Section H: BENT'S RULE

#### SINGLE CORRECT ANSWER TYPE

- Which of the following statements is incorrect for  $\text{PCl}_5$ ?
  - Its three P – Cl bond lengths are equal.
  - It involves  $\text{sp}^3 \text{d}$  hybridization.
  - Each bond angle is more than  $120^\circ$ .
  - Its shape is trigonal bipyramidal.
- The shortest O – O bond length is present in:
  - $\text{O}_2\text{F}_2$
  - $\text{O}_2$
  - $\text{H}_2\text{O}_2$
  - $\text{O}_3$
- In which of the following molecules/ions are all the bonds not of equal length?
  - $\text{XeF}_4$
  - $\text{BF}_4^-$
  - $\text{ClF}_3$
  - $\text{SiF}_4$

#### MULTIPLE CORRECT ANSWER TYPE

- Correct statement(s) regarding  $\text{As}(\text{CH}_3)_2\text{Cl}_2$  molecules is/are:
  - Maximum three halogen atoms can lie in same plane.
  - Both axial and equatorial position may have equal number of atoms.
  - As – Cl bond length is longer than As – F bond length.
  - $\text{CH}_3$  is at equatorial position.
- Which of the following statements is/are not correct for following compounds?
 

(I) $\text{SCL}_2(\text{OCH}_3)_2$ and	(II) $\text{SF}_3(\text{OCH}_3)_2$
--	------------------------------------

  - $-\text{CH}_3$  groups in both cases occupy the same position
  - Cl – atoms occupy equatorial position in case of (I) and F – atoms occupy equatorial position in case of (II)
  - Cl – atoms occupy axial position in case of (I) & F – atoms occupy equatorial position in case of (II)
  - Both Cl in (I) and both F in (II) occupy axial positions.
- Select correct comparison(s):
 

(a) $\text{N}_2\text{H}_4 > \text{N}_2\text{F}_4$	(N – N bond length)
---	---------------------

- (b)  $C_2H_6 > C_2F_6$  (C – C bond length)  
 (c) In  $PCl_4F$  :  $P - F_{axial} > P - Cl_{equatorial}$  (both length)  
 (d) In  $PCl_3F_2$  :  $P - Cl_{equatorial} > P - F_{axial}$  (bond length)
7. In molecule of  $PF_3Cl_2$ ,
- (a)  $P - Cl_{equatorial}$  is longest bond.  
 (b)  $P - F_{axial}$  bond length is more than  $P - F_{equatorial}$  bond length  
 (c)  $P - Cl_{equatorial}$  bond length is more than  $P - F_{equatorial}$  bond length  
 (d)  $P - Cl_{equatorial}$  bond length is more than  $P - F_{axial}$  bond length.

## COMPREHENSION TYPE

### Paragraph for Question no. 08 to 10

Bent's rule states that a lone pair prefers to occupy that hybrid which has greater % of s – character and a strong electron negative atom prefers to overlap with that hybrid orbital of central atom which has relatively smaller percentage of s – character.

8. The correct statement is:
- (a) In  $CH_2F_2$  the bond angle is larger  $109^\circ 28'$   
 (b) In  $CH_2F_2$  the C – F bond has more than 25% s – character  
 (c) In  $CH_2F_2$  the H – C – H bond angle is larger than  $109^\circ 28'$   
 (d) In  $CH_2F_2$  the C – H bond has than 25% s – character
9. The incorrect statement regarding  $PCl_2F_3$  molecule will be:
- (a) Given compound is polar  
 (b) Both axial positions are occupied by F – atoms  
 (c) Both Cl atoms are present in equatorial positions  
 (d) One Cl atom present at axial and other Cl atom is present at equatorial position of geometry
10. The highest H – C – H bond angle is present in:
- (a)  $CH_2F_2$   
 (b)  $CH_4$   
 (c)  $CH_3Cl$   
 (d)  $CH_3F$

## Section I: BOND PARAMETERS

### SINGLE CORRECT ANSWER

1. The correct order of H – M – H bonds angle is:
- (a)  $NH_3 < PH_3 < SbH_3 < BiH_3$   
 (b)  $AsH_3 < SbH_3 < PH_3 < NH_3$   
 (c)  $NH_3 < PH_3 < BiH_3 < SbH_3$   
 (d)  $BiH_3 < SbH_3 < AsH_3 < PH_3$
2. The correct increasing order of bond angle among  $BF_3$ ,  $PF_3$  and  $ClF_3$ :
- (a)  $BF_3 < PF_3 < ClF_3$   
 (b)  $PF_3 < BF_3 < ClF_3$   
 (c)  $ClF_3 < PF_3 < BF_3$



- (d)  $\text{BF}_3 = \text{PF}_3 = \text{ClF}_3$
3. Among the following species, which has the least bond angle around the central atom?
- $\text{O}_3$
  - $\text{I}_3^-$
  - $\text{NO}_2^-$
  - $\text{XeF}_5^-$
4. The bond angles of  $\text{NH}_3$ ,  $\text{NH}_2^-$ ; and  $\text{NH}_4^+$ ; are in the order;
- $\text{NH}_2^- > \text{NH}_3 > \text{NH}_4^+$
  - $\text{NH}_4^+ > \text{NH}_3 > \text{NH}_2^-$
  - $\text{NH}_3 > \text{NH}_2^- > \text{NH}_4^+$
  - $\text{NH}_3 > \text{NH}_4^+ > \text{NH}_2^-$
5. The  $\text{H}-\text{C}-\text{H}$  bond angle in  $\text{CH}_4$  is  $109.5^\circ$ . Due to lone pair repulsion, the  $\text{H}-\text{O}-\text{H}$  angle in  $\text{H}_2\text{O}$  will:
- Remain the same
  - Increase
  - Decrease
  - Become  $180^\circ$
6. The molecule having the largest bond angle is:
- $\text{H}_2\text{O}$
  - $\text{H}_2\text{S}$
  - $\text{H}_2\text{Se}$
  - $\text{H}_2\text{Te}$
7. The compound  $\text{MX}_4$  is tetrahedral. The number of  $\angle \text{XMX}$  angles in the compound is:
- Three
  - Four
  - Five
  - Six
8. Bond dissociation energy is maximum in:
- $\text{NO}_2^-$
  - $\text{N}_2\text{O}$
  - $\text{NO}_2$
  - $\text{NO}_3^-$
9. Bond dissociation energy is maximum in:
- $\text{C}-\text{F}$
  - $\text{C}-\text{Cl}$
  - $\text{C}-\text{Br}$
  - $\text{C}-\text{I}$
10.  $\text{C}-\text{C}$  bond length is minimum in:
- $\text{C}_2\text{H}_6$
  - $\text{C}_2\text{H}_4$
  - $\text{C}_2\text{H}_2$
  - Benzene
11. Select the correct comparison of bond dissociation energy:

- (a)  $F_2 > Cl_2$   
 (b)  $Cl_2 > Br_2$   
 (c)  $I_2 > Cl_2$   
 (d)  $I_2 > Br_2$

12. C – C bond length is shortest in:

- (a)  $\text{>C}=\text{C}<$   
 (b)  $\equiv C - C \equiv$   
 (c)  $\text{>C}-C \equiv$   
 (d)  $\text{>C}-C \leq$

### MULTIPLE CORRECT ANSWER TYPE

13. Which of the following statements is/are correct?

- (a)  $ClF_4$  molecule is bent T – shaped.  
 (b) In  $SF_4$  molecule, F – S – F equatorial bond angle is less than  $120^\circ$  due to lp – lp repulsion.  
 (c) In  $[ICl_4]^-$  molecular ion, Cl – I – Cl bond angle is  $90^\circ$ .  
 (d) In  $OBr_2$ , the bond angle is less than of  $OCl_2$ .

14. All bond angles are identical in:

- (a)  $CF_4$   
 (b)  $SF_6$   
 (c)  $CH_3Cl$   
 (d)  $C_2H_4$

15. Bond dissociation energy order not according to bond length factor in which pair(s):

- (a) C – C, Si – Si  
 (b) N – N, P – P  
 (c) O – O, S – S  
 (d) C – H, C – Si

### MATRIX MATCH TYPE

16. Column – I (Species)

- (a) – SCN  
 (b)  $O_3$   
 (c)  $SO_4^{2-}$   
 (d)  $CO_3^{2-}$

Column – II (Bond angle)

- (P)  $180^\circ$   
 (Q)  $120^\circ$   
 (R)  $109.5^\circ$   
 (S)  $117^\circ$

17. Column – I

- (a)  $XeF_5^-$   
 (b)  $PBr_4^+$   
 (c)  $IOF_3$   
 (d)  $NH_2^-$

Column – II

- (P) d – orbital with zero nodal plane is used in hybridization  
 (Q) Non – axial d – orbital is used in hybridization  
 (R) Planar species  
 (S) Non – planar species  
 (T) Bond angle  $109^\circ 28'$  or less than  $109^\circ 28'$

18. How many right angle, bond angles are present in a  $ClF_3$  molecule?

19. How many possible  $\angle$  FSeF angles are present in a  $\text{SeF}_4$  molecule?
20. How many species contain atleast one bond angle of  $180^\circ$ ?
- $\text{ClO}_2^+$ ,  $\text{C}_3\text{O}_2$ ,  $\text{CS}_2$ ,  $\text{N}_3^-$ ,  $\text{H}_2\text{S}$ ,  $\text{OF}_2$ ,  $\text{I}_3^+$ ,  $\text{C}_3\text{H}_4$ ,  $\text{N}_2\text{O}$

### Section J: DRAGO'S GENERALISATION

- Central atom does not use hybrid orbital for bond formation in:
  - $\text{H}_2\text{O}$
  - $\text{H}_2\text{S}$
  - $\text{NH}_3$
  - $\text{CH}_4$
- Bond angle in  $\text{PH}_3$  is close to:
  - $120^\circ$
  - $90^\circ$
  - $180^\circ$
  - $109^\circ 28'$
- S – character is maximum in lone pair of central atom in:
  - $\text{H}_2\text{O}$
  - $\text{SbH}_3$
  - $\text{OF}_2$
  - $\text{CH}_4$

### COMPREHENSION TYPE

#### Paragraph for Question no. 4 and 5

If the central atom is of third row or below this in the periodic table, then lone pair will occupy s – orbital and bonding will be through almost pure p – orbitals and bond angles are nearly  $90^\circ$ , if the substituents' electronegativity value is  $\leq 2.5$ .

- In which of the following option, the covalent bond has maximum %s character?
  - S – H bond in  $\text{H}_2\text{S}$
  - P – H bond in  $\text{PH}_3$
  - N – H bond in  $\text{NH}_3$
  - S – F bond in  $\text{SF}_6$
- In which molecule, lone pair of central atom forms strongest sigma bond with  $\text{H}^+$ ?
  - $\text{H}_2\text{O}$
  - $\text{H}_2\text{Se}$
  - $\text{H}_2\text{S}$
  - $\text{H}_2\text{Te}$

#### Paragraph for Question no.6 to 8

Hybridization is a theoretical concept, it explains observed facts about structure of species, however in some molecular species, there is no need to consider hybridization to explain observed structural facts.

6. If one would have considered hybridization in  $\text{PH}_3$ , then which of the following experimental fact can't be explained satisfactorily?
  - (a) Geometry
  - (b) Bond length
  - (c) Bond angle
  - (d) None of these
7. In which of the following transformation, maximum change in bond angle is expected?
  - (a)  $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$
  - (b)  $\text{PH}_3 + \text{H}^+ \rightarrow \text{PH}_4^+$
  - (c)  $\text{AsH}_3 + \text{H}^+ \rightarrow \text{AsH}_4^+$
  - (d)  $\text{SbH}_3 + \text{H}^+ \rightarrow \text{SbH}_4^+$
8. Strongest Lewis among given is:
  - (a)  $\text{NH}_3$
  - (b)  $\text{PH}_3$
  - (c)  $\text{AsH}_3$
  - (d)  $\text{SbH}_3$

### INTEGER ANSWER TYPE

9. Drago's generalization is not applicable to how many species given below?  
 $\text{AsH}_3$ ,  $\text{GeH}_4$ ,  $\text{PH}_3$ ,  $\text{H}_2\text{Se}$ ,  $\text{SiH}_4$ ,  $\text{NH}_3$ ,  $\text{H}_2\text{S}$ ,  $\text{H}_2\text{O}$ ,  $\text{SbH}_3$
10. How many hybrid orbital(s) are used for bond formation in a molecule of  $\text{H}_2\text{Te}$ ?

### Section K: BACK BONDING & BRIDGE BONDING

#### SINGLE CORRECT ANSWER TYPE

1. The state of hybridization of central atom in dimer form of both  $\text{BH}_3$  and  $\text{BeH}_2$  is respectively.
  - (a)  $\text{sp}^2$ ,  $\text{sp}$
  - (b)  $\text{sp}^3$ ,  $\text{sp}^2$
  - (c)  $\text{sp}^3$ ,  $\text{sp}^3$
  - (d)  $\text{sp}^2$ ,  $\text{sp}^3$
2. The geometry with respect to the central atom of the following molecules are:  
 $\text{N}(\text{SiH}_3)_3$ ;  $\text{Me}_3\text{N}$ ;  $(\text{SiH}_3)_3\text{P}$ 
  - (a) Planar, pyramidal, planar
  - (b) Planar, pyramidal, pyramidal
  - (c) Pyramidal, pyramidal, pyramidal
  - (d) Pyramidal, planar, pyramidal
3. In which of the following cases, the strength of  $\pi$  – bond (back bond) maximum?
  - (a)  $\text{PF}_3$
  - (b)  $\text{BF}_3$
  - (c)  $(\text{H}_3\text{Si})_3\text{N}$
  - (d)  $(\text{SiH}_3)_2\text{O}$

#### MULTIPLE CORRECT ANSWER TYPE

4. Which of the following statement is/are correct for  $\text{CCl}_3^-$  and  $\text{:CCl}_2$ 
  - (a) Back bonding in both cases is from Cl to C – atom.
  - (b) Back bonding in both cases is from C to Cl – atom.
  - (c) Back bonding in  $\text{CCl}_3^-$  is from C to Cl but reverse in  $\text{:CCl}_2$
  - (d)  $\text{:CCl}_2$  is planar.
5. Which of the following molecule(s) is/are having  $p\pi - p\pi$  bonding?
  - (a)  $\text{BF}_3$
  - (b)  $\text{BeF}_2$
  - (c)  $\text{B}_3\text{N}_3\text{H}_6$
  - (d)  $\text{BCl}_3$
6. Choose the incorrect statements:
  - (a)  $\text{CH}_3\text{NCS}$  molecule is linear.
  - (b) In  $\text{SiH}_3$ ,  $\text{NCS}$ ,  $\text{SiNC}$  bond angle is  $180^\circ$ .
  - (c)  $\text{N}_3^-$  molecule is bent.
  - (d)  $\text{P}(\text{SiH}_3)_3$  molecule is planar.
7. Which of the following statement(s) is/are incorrect about  $\text{I}_2\text{Cl}_6$  molecule?
  - (a) All I – Cl bonds are equivalent.
  - (b) Molecule contains non – polar bonds.
  - (c) Molecule is non planar.
  - (d) All bond angles are equal.

### Paragraph for Question No.s. 8 & 9

Bridge bonding is a specific kind of bonding in pages of chemistry. In general  $\sigma$  – bond pair delocalization is very difficult. But electron deficiency of the central atom forces to delocalise and forms this kind of bond.

8. The  $\text{B}_2\text{H}_6$  molecule is dissolved in tetrahydrofuran. Which atom(s) is/are have change in hybridization with respect to reactant and final product of the process given?
  - (a) B only
  - (b) B and O
  - (c) B, O and C
  - (d) None of these
9. In which of the dimerization process, the achievement of the octet is not the driving force?
  - (a)  $2\text{AlCl}_3 \rightarrow \text{Al}_2\text{Cl}_6$
  - (b)  $\text{BeCl}_2 \rightarrow \text{BeCl}_2$  (solid)
  - (c)  $2\text{ICl}_3 \rightarrow \text{I}_2\text{Cl}_6$
  - (d) All of these

### MATRIX MATCH TYPE

10. Column I

(a)  $\text{B}_2\text{H}_6$

Column II

(P)  $(3\text{C} - 4\text{e}^-)$  bond

- (b)  $\text{Be}_2\text{H}_4$   
 (c)  $\text{Be}_2\text{Cl}_4$   
 (d)  $\text{Al}_2(\text{CH}_3)_6$
- (Q)  $(3\text{C} - 2e^-)$  bond  
 (R)  $(2\text{C} - 2e^-)$  bond  
 (S)  $sp^3$  hybridisation of central atom  
 (T)  $sp^2$  hybridisation of central atom

**11. Column – I**

- (a)  $\text{N}(\text{SiH}_3)_3$   
 (b)  $\text{N}(\text{CH}_3)_3$   
 (c)  $\text{B}_2\text{H}_6$   
 (d)  $\text{BF}_3$

**Column – II**

- (P)  $p\pi - d\pi$  back bonding  
 (Q)  $sp^3$  hybridisation for underlined atom  
 (R)  $p\pi - p\pi$  back bonding  
 (S) neither  $p\pi - p\pi$  nor  $p\pi - d\pi$  back bonding  
 (T) Underlined atom combine with electron rich molecule

**INTEGER ANSWER TYPE**

12. How many vacant orbitals of central atom are involved in hybridisation in formation of a molecule of  $\text{Be}_2\text{H}_4$ ?  
 13. How many hybrid orbitals are possessed by a central atom in solid  $\text{BeCl}_2$ ?

**SUBJECTIVE TYPE**

14. Assign planar/non – planar to the following species:

- |                               |                                  |   |                               |                              |  |
|-------------------------------|----------------------------------|---|-------------------------------|------------------------------|--|
| (1) $\text{XeF}_2$            | (2) $\text{ClF}_3$               | (3) $\text{SF}_4$                                 | (4) $\text{SiF}_4$            | (5) $\text{PCl}_3$           | (6) $\text{H}_2\text{C} = \text{CH}_2$ |
| (7) $\text{NH}_3$             | (8) $\text{HC} \equiv \text{CH}$ | (9) $\text{H}_2\text{C} = \text{C} = \text{CH}_2$ | (10) $\text{BrF}_5$           | (11) $\text{XeF}_5^-$        | (12) $\text{XeF}_5^+$                  |
| (13) $\text{CH}_2\text{Cl}_2$ | (14) $\text{ClO}_2^+$            | (15) $\text{B}_2\text{H}_6$                       | (16) $\text{Al}_2\text{Cl}_6$ | (17) $\text{I}_2\text{Cl}_6$ | (18) $\text{B}_3\text{N}_3\text{H}_6$  |
| (19) $\text{ICl}_4^-$         | (20) $\text{NO}_3^-$             |   |                               |                              |  |

15. Back – bonding is present in:

- |                               |                               |  |  |                   |                   |
|-------------------------------|-------------------------------|--|--|-------------------|-------------------|
| (1) $\text{B}(\text{OH})_3$   | (2) $\text{CF}_3$             | (3) $\text{F}_2\text{B} - \text{NH}_2$             | (4) $\text{OF}_2$                                      | (5) $\text{CF}_3$ | (6) $\text{PF}_3$ |
| (7) $\text{B}(\text{CH}_3)_3$ | (8) $\text{CH}_2 - \text{OH}$ | (9) $\text{CH}_3 - \text{N} = \text{C} = \text{O}$ | (10) $\text{Al} - \text{O} - \text{Al}(\text{CF}_3)_2$ |                   |                   |

**Section L: ODD ELECTRON SPECIES AND NON – EXISTING SPECIES**

1. Odd electron does not occupy hybrid orbital in:

- (a)  $\dot{\text{C}}\text{F}_3$   
 (b)  $\dot{\text{N}}\text{O}_2$   
 (c)  $\dot{\text{C}}\text{H}_3$   
 (d)  $\dot{\text{C}}\text{IO}_3$

2.  $\text{PH}_5$  does not exist due to:

- (a) Steric crowding  
 (b) Lack of d – orbital contraction  
 (c) p – orbital contraction  
 (d) Inert pair effect

3.  $\text{SiF}_6^{2-}$  is known while  $\text{SiCl}_6^{2-}$  is not known. It can be explained on the basis of:

- (a) Inert pair effect

- (b) Lack of valency  
 (c) Steric crowding  
 (d) All of these
4. Bond angle is more than  $120^\circ$  in:

- (a)  $\dot{\text{C}}\text{F}_3$   
 (b)  $\dot{\text{N}}\text{O}_2$   
 (c)  $\dot{\text{C}}\text{H}_3$   
 (d) None of these

### MULTIPLE CORRECT ANSWER TYPE

5. Which of the following species form dimer?

- (a)  $\dot{\text{C}}\text{F}_3$   
 (b)  $\dot{\text{N}}\text{O}_2$   
 (c)  $\dot{\text{C}}\text{H}_3$   
 (d)  $\dot{\text{C}}\text{IO}_3$

6. Which of the following is/are not known?

- (a)  $\text{OF}_6$   
 (b)  $\text{SF}_4$   
 (c)  $\text{NCl}_5$   
 (d)  $\text{SH}_6$

7. In which of the following dimerization processes, hybridisation of central atom is unchanged?

- (a)  $2\dot{\text{C}}\text{F}_3 \rightarrow \text{C}_2\text{F}_6$   
 (b)  $2\dot{\text{N}}\text{O}_2 \rightarrow \text{N}_2\text{O}_4$   
 (c)  $2\dot{\text{C}}\text{H}_3 \rightarrow \text{C}_2\text{H}_6$   
 (d)  $2\dot{\text{C}}\text{IO}_3 \rightarrow \text{Cl}_2\text{O}_6$

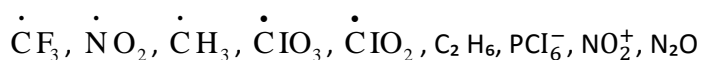
8. Steric factor does not allow existence of:

- (a)  $\text{PH}_5$   
 (b)  $\text{PCl}_6^-$   
 (c)  $\text{PI}_6^-$   
 (d)  $\text{PCl}_4^+$

### INTEGER ANSWER TYPE

9. In how many paramagnetic species, electrons of central atom occupy hybrid orbital?





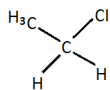
## SUBJECTIVE TYPE

10. Select the species which do not exist?

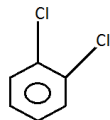
- |                      |                              |                     |                     |                           |                          |
|----------------------|------------------------------|---------------------|---------------------|---------------------------|--------------------------|
| (i) $\text{NCl}_5$   | (ii) $\text{OF}_4$           | (iii) $\text{HOFO}$ | (iv) $\text{XeH}_2$ | (v) $\text{PCl}_4^+$      | (vi) $\text{SiI}_6^{2-}$ |
| (vii) $\text{PH}_5$  | (viii) $\text{BF}_6^{3-}$    | (ix) $\text{SH}_6$  | (x) $\text{SH}_2$   | (xi) $\text{BeCl}_6^{4-}$ | (xii) $\text{PH}_4^+$    |
| (xiii) $\text{IH}_7$ | (xiv) $\text{B}_2\text{H}_6$ | (xv) $\text{XeF}_2$ |                     |                           |                          |

## Section M: DIPOLE MOMENT

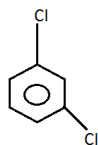
- Which of the following has least polarity in bond?
  - $\text{H}-\text{F}$
  - $\text{H}-\text{Cl}$
  - $\text{H}-\text{O}$
  - $\text{H}-\text{S}$
- $\text{BeF}_2$  has zero dipole moment whereas  $\text{H}_2\text{O}$  has a dipole moment because:-
  - Water is linear.
  - $\text{H}_2\text{O}$  is bent
  - F is more electronegative than O.
  - Hydrogen bonding is present in  $\text{H}_2\text{O}$ .
- The correct order of dipole moment is:
  - $\text{CH}_4 < \text{NF}_3 < \text{NH}_3$
  - $\text{NH}_3 < \text{CH}_4 < \text{NH}_3$
  - $\text{NH}_3 < \text{NF}_3 < \text{CH}_4$
  - $\text{NF}_3 < \text{NH}_3 < \text{CH}_4$
- Which one of the following molecules has highest dipole moment?
  - $\text{H}_2\text{S}$
  - $\text{CO}_2$
  - $\text{CCl}_4$
  - $\text{BF}_3$
- Which one of the following hydrocarbons has the lowest dipole moment?



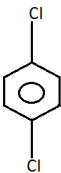
- $\text{CH}_3\text{C} \equiv \text{CCH}_3$
  - $\text{CH}_3\text{CH}_2\text{C} \equiv \text{CH}$
  - $\text{CH}_2 = \text{CH} - \text{C} \equiv \text{CH}$
6. Dipole moment is maximum in:



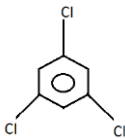
(a)



(b)



(c)



(d)

7. A diatomic molecule  $H - X$  has a dipole moment of 1.2 D. If the bond length is  $1.0 \times 10^{-8} \text{ cm}$ , what fraction of charge does exist on each atom?
- 0.25
  - 0.75
  - 0.50
  - 0.33

### MULTIPLE CORRECT ANSWER TYPE




8. Dipole moment of  $AX_4$  type of molecule is zero. The geometry of it can be:
- Tetrahedral
  - Square planar
  - See – saw
  - Bent T
9. Which molecule is/are planar as well as non polar?
- $XeF_4$
  - $CF_4$
  - $CO_2$
  - $SF_6$
10. Select the incorrect order of dipole moment?
- $XeF_2 > XeO_3F_2$
  - $BF_3 > BCl_3$
  - $CF_4 > CH_4$
  - $NH_3 > NF_3$
11.  $XY_3$  type molecule has:
- [X = central atom, Y = monovalent surrounding atom]
- Planarity when they are polar.
  - Polarity when they are planar.
  - Polarity when are non – planar.
  - Planarity when they are non – polar.

### INTEGER ANSWER TYPE

12. 1 Debye =  $1 \times 10^{-y}$  esu – cm. The value of  $\left(\frac{y}{2}\right)$  is .....
13. Measured dipole moment of HCl is  $3.436 \times 10^{-30}$  coulomb – meter. Bond length in HCl is  $2.29 \times 10^{-10}$  meter. The percentage ionic character in HCl is x. The approximate value of x is:
14. How many given species are non – polar as well as non – planar?  
 $\text{PCl}_3\text{F}_2$ ,  $\text{IF}_7$ ,  $\text{SO}_3$ ,  $\text{XeF}_4$ ,  $\text{CH}_2\text{Cl}_2$ ,  $\text{CCl}_4$ ,  $\text{ClF}_3$ ,  $\text{PCl}_2\text{F}_3$ ,  $\text{XeF}_2$

### SUBJECTIVE TYPE

15. Assign polar/non – polar to the following species:

- |                     |                               |                              |                     |                               |                               |                             |
|---------------------|-------------------------------|------------------------------|---------------------|-------------------------------|-------------------------------|-----------------------------|
| (1) $\text{CO}_2$   | (2) $\text{CH}_2\text{Cl}_2$  | (3) $\text{CCl}_4$           | (4) $\text{BF}_3$   | (5) $\text{NO}_2$             | (6) $\text{XeF}_4$            | (7) $\text{SF}_6$           |
| (8) $\text{SiF}_4$  | (9) $\text{SF}_4$             | (10) $\text{XeF}_2$          | (11) $\text{SO}_3$  | (12) $\text{SO}_2\text{Cl}_2$ | (13) $\text{N}_2\text{O}$     | (14) $\text{SO}_2$          |
| (15) $\text{PCl}_5$ | (16) $\text{XeO}_3\text{F}_2$ | (17) $\text{I}_2\text{Cl}_6$ | (18) $\text{ClF}_3$ | (19) $\text{XeO}_4$           | (20) $\text{XeO}_2\text{F}_2$ | (21) $\text{C}_2\text{O}_2$ |
- 
- |  |  |                               |  |                               |
|--|--|-------------------------------|--|-------------------------------|
| (22)  | (23)  | (24) $\text{ClO}_2^+$         | (25)  | (26) $\text{PCl}_2\text{F}_3$ |
| (27) $\text{H}_2\text{O}_2$  | (28) $\text{BeF}_2$  | (29) $\text{PCl}_3\text{F}_2$ | (30) $\text{H}_2\text{S}$  |                               |

### Section N: MOLECULAR ORBITAL THEORY – I

#### SINGLE CORRECT ANSWER TYPE

- The paramagnetic property of oxygen is well explained by:-
  - Molecular orbital theory
  - Resonance theory
  - Valence bond theory
  - VSEPR theory
- According to molecular orbital theory, which of the following statement about the magnetic character and bond order is correct regarding  $\text{O}_2^+$ ?
  - Paramagnetic and bond order  $< \text{O}_2$
  - Paramagnetic and bond order  $> \text{O}_2$
  - Diamagnetic and bond order  $< \text{O}_2$
  - Diamagnetic and bond order  $> \text{O}_2$
- Which of the following has fractional bond order?
  - $\text{O}_2^{2+}$
  - $\text{O}_2^{2-}$
  - $\text{F}_2^{2-}$
  - $\text{H}_2^-$
- What of the following does not exist?
  - $\text{He}_2$
  - $\text{N}_2$
  - $\text{O}_2$
  - $\text{H}_2$

5. Which of the following statement is not correct regarding bonding molecular orbitals?
- Bonding molecular orbitals possess less energy than the combining atomic orbitals.
  - Bonding molecular orbitals have low electron density between the two nuclei.
  - Electron in bonding molecular contributes to the attraction between atoms.
  - They are formed when the lobes of the combining atomic orbitals have the same sign.
6. Which of the following combination of orbitals is correct?



(a)



(b)



(c)



(d)

7. How many nodal planes are present in a  $\sigma_s$  bonding molecular orbital?

- Zero
- 1
- 2
- 3

8. If  $z$  - axis is the molecular axis, then  $\sigma$  - molecular orbitals can be formed by the overlap of:

- $s + p_z$
- $p_x + p_y$
- $p_z + p_y$
- $p_x + p_x$

### MULTIPLE CORRECT ANSWER TYPE

9. In which of the following,  $\pi$  - bond order is greater than  $\sigma$ -bond order?

- $N_2$
- $O_2$
- $O_2^+$
- $O_2^-$

10. Which of the following changes involve the change in magnetic moment?

- $N_2 \rightarrow N_2^+$
- $H_2^+ \rightarrow H_2^-$
- $O_2 \rightarrow O_2^{2-}$
- $N_2^+ \rightarrow N_2^-$

11. If  $N_B$  is the number of bonding electrons and  $N_A$  is the number of antibonding electrons of a molecule. Then chose the incorrect statement(s) for the relationship,  $N_B > N_A$ :

- (a) Molecule will always be diamagnetic.
- (b) Molecule may have integral, fractional or zero value of bond order.
- (c) Molecule is only paramagnetic species.
- (d) Molecule does not exist.

## COMPREHENSION TYPE

### Paragraph for question no. 12 & 13

According to MOT, two atomic orbitals overlap resulting in the formation of molecular orbitals. Number of atomic orbitals overlapping together is equal to the molecular orbital formed. The two atomic orbitals thus formed by LCAO (linear combination of atomic orbital) in the same phase or in the different phase are known as bonding and antibonding molecular orbitals respectively. The energy of bonding molecular orbital is lower than that of the pure atomic orbital by an amount  $\Delta$ . This is known as the stabilization energy. The energy of antibonding molecular orbital is increased by  $\Delta'$  (destabilization energy).

12. The bond order of  $N_2$  is equal to that of:
- (a)  $O_2$
  - (b)  $O_2^{2-}$
  - (c)  $O_2^{2+}$
  - (d) None
13. Which among the following pairs contain paramagnetic species?
- (a)  $O_2^{2-}$  and  $N_2$
  - (b)  $O_2^+$  and  $N_2$
  - (c)  $O_2$  and  $N_2$
  - (d)  $O_2$  and  $N_2^-$

## MATRIX MATCH TYPE

### 14. Column – I

(Axial/sideways combination of appropriate (type of molecular orbital) / inappropriate pure orbitals)

- (a) P + P pure orbitals
- (b) S + P pure orbitals
- (c) (non – axial) d + p pure orbitals
- (d) (axial) d + p pure orbitals

### Column – II

- (P)  $\sigma$  – bonding molecular orbital
- (Q)  $\pi$  – anti – bonding molecular orbital
- (R)  $\sigma$  – anti – bonding molecular orbital
- (S)  $\pi$  – bonding molecular orbital
- (T) Non – bonding molecular orbital

## SUBJECTIVE TYPE

15. Find bond order and magnetic behaviour of following diatomic species using molecular orbital theory:

- |              |               |                 |                 |                 |              |
|--------------|---------------|-----------------|-----------------|-----------------|--------------|
| (1) $H_2^+$  | (2) $He_2^+$  | (3) $N_2$       | (4) $O_2^-$     | (5) $Li_2$      | (6) $O_2^+$  |
| (7) $C_2$    | (8) $B_2$     | (9) $N_2^+$     | (10) $O_2^{2-}$ | (11) $O_2^{2+}$ | (12) $N_2^-$ |
| (13) $H_2^-$ | (14) $F_2^+$  | (15) $C_2^{2-}$ | (16) $C_2^+$    | (17) $B_2^+$    | (18) $O_2$   |
| (19) $F_2$   | (20) $Li_2^+$ | (21) $F_2^-$    | (22) $Be_2^+$   | (23) $N_2^{2-}$ | (24) – CH    |

(25) NO

 (26) NO<sup>+</sup>

(27) CO

(28) CN

 (29) NO<sup>-</sup>

(30) BN

**Section O: MOLECULAR ORBITAL THEORY – II**

- In which of the following ionization processes, the bond order has increased and the magnetic behaviour has changed?
  - NO → NO<sup>+</sup>
  - O<sub>2</sub> → O<sub>2</sub><sup>+</sup>
  - N<sub>2</sub> → N<sub>2</sub><sup>+</sup>
  - C<sub>2</sub> → C<sub>2</sub><sup>+</sup>
- Which of the following species will have the minimum bond energy?
  - N<sub>2</sub>
  - N<sub>2</sub><sup>-</sup>
  - N<sub>2</sub><sup>+</sup>
  - N<sub>2</sub><sup>-2</sup>
- Which of the following species exhibits the diamagnetic behaviour?
  - O<sub>2</sub><sup>+</sup>
  - O<sub>2</sub>
  - NO
  - O<sub>2</sub><sup>2-</sup>
- Which of the following statements is correct about N<sub>2</sub> molecule:-
  - It has bond order of 3.
  - The number of unpaired electrons present in it is zero.
  - The order of filling of MOs is  $\pi(2p_x) = \pi(2p_y), \sigma(2p_z)$ .
  - All the above three statements are correct.
- Using MO theory, predict which of the following species has the shortest bond length?
  - O<sub>2</sub><sup>-</sup>
  - O<sub>2</sub><sup>2-</sup>
  - O<sub>2</sub><sup>2+</sup>
  - O<sub>2</sub><sup>+</sup>
- Increasing order of bond length in NO, NO<sup>+</sup> and NO<sup>-</sup> is:-
  - NO > NO<sup>-</sup> > NO<sup>+</sup>
  - NO<sup>+</sup> < NO < NO<sup>-</sup>
  - NO < NO<sup>+</sup> < NO<sup>-</sup>
  - NO < NO<sup>+</sup> = NO<sup>-</sup>

**MULTIPLE CORRECT ANSWER TYPE**

- Which of the following species are repelled in the external magnetic field?
  - N<sub>2</sub>
  - O<sub>2</sub>
  - C<sub>2</sub>
  - H<sub>2</sub>

8. Which of the following is/are correct order of energy of molecular orbitals?
- $\sigma_{2pz} > \pi_{2px}$  ( $O_2$  molecule)
  - $\sigma_{2pz} > \pi_{2py}$  ( $N_2$  molecule)
  - $\sigma_{2pz} > \pi_{2px}$  ( $B_2$  molecule)
  - $\sigma_{2pz} > \pi_{2px}$  ( $F_2$  molecule)
9. Which of the following has more no. of bonding molecular orbital electron(s) than antibonding molecular orbital electrons?
- $B_2$
  - $N_2$
  - $O_2$
  - $H_2$
10. Select the correct statement(s) about the compound  $NO[BF_4]$ :
- It has  $5\sigma$  and  $2\pi$  bond.
  - Nitrogen – oxygen bond length is higher than nitric oxide ( $NO$ ),
  - It is a diamagnetic species.
  - B – F bond length in this compound is lower than in  $BF_3$ .
11. Which of the following molecular species is/are having  $\pi_{2p}^*$  as H.O.M.O. (highest occupied molecular orbital):
- $Li_2$
  - $N_2$
  - $F_2$
  - $C_2^{2-}$

### MATRIX MATCH TYPE

#### 12. Column I

##### (Species)

- $O_2[BF_4]$
- $KO_2$
- $Na_2 O_2$
- $O_2$

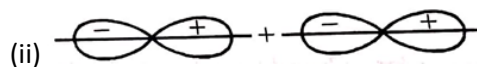
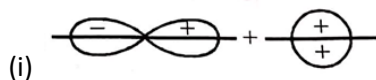
#### Column II

##### (Properties)

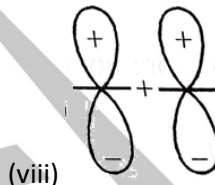
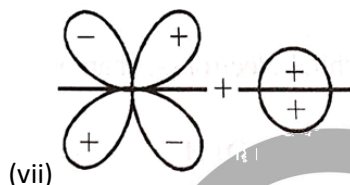
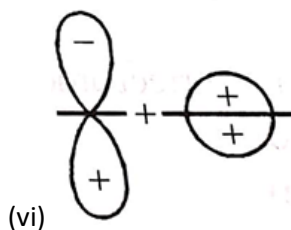
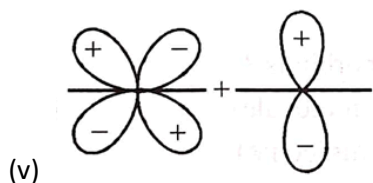
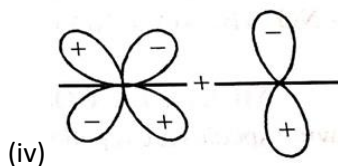
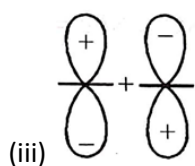
- Paramagnetic ion
- Diamagnetic ion
- Fractional bond order
- Bond order  $\geq 2$
- HOMO of diatomic ion contains 4 electrons

### INTEGER ANSWER TYPE

13. If Hund's rule is violated then find the total number of species among following which will be diamagnetic?  $B_2$ ,  $O_2$ ,  $N_2^-$ ,  $C_2^+$ ,  $N_2^+$ ,  $F_2^+$ ,  $N_2^{2-}$ ,  $Li_2^+$ ,  $O_2^+$
14. Total number of unpaired electron(s) present in both cationic and anionic part of compound  $O_2[PF_6]$ .
15. There are some arrangements of atomic orbitals which are given below:







Then calculate the value of " $Q \times R - P$ " where, P, Q and R are no. of arrangements which give bonding molecular orbitals (positive overlap), antibonding molecular orbitals (negative overlap) and non – bonding molecular orbitals (zero overlap) respectively.

## Section P: INTRODUCTION OF IONIC BOND

### SINGLE CORRECT ANSWER TYPE

- An electrovalent bond or ionic bond is formed generally between:-
  - Two electronegative atoms
  - Two metals
  - Electropositive and electronegative atoms.
  - Two electropositive atoms.
- The electronegativity of cesium is 0.7 and that of fluorine is 4.0. The bond formed between the two is
  - Covalent
  - Electrovalent/ionic
  - Coordinate
  - Metallic
- Element A has 3 electrons in the outermost orbit and element B has 6 electrons in the outermost orbit. The formula of the compound formed between A and B would be:-
  - $A_2B_3$
  - $A_2B_6$
  - $A_2B$
  - $A_3B_2$

4. Ionic compounds in general possess both:-
  - (a) High melting points and non – directional bonds.
  - (b) High melting points and low boiling points
  - (c) Directional bonds and low boiling points
  - (d) High solubilities in polar and non – polar solvents.
5. An ionic compound  $A^+B^-$  is most likely to be formed when –
  - (a) Ionization energy of A is low.
  - (b) Electron affinity of B is high.
  - (c) Electron affinity of B is low.
  - (d) Both (a) and (b)
6. Electrovalent compounds or ionic compounds do not show isomerism. The reason is –
  - (a) Presence of ions
  - (b) Strong electrostatic force of attraction
  - (c) Brittleness
  - (d) Non – directional nature of ionic bond
7. Oxide of a metal 'M' has formula  $M_2O_3$ . The formula of nitride of this metal 'M' will be:-
  - (a)  $M_3N$
  - (b)  $MN$
  - (c)  $M_3N_2$
  - (d)  $M_2N_3$
8. Solid NaCl is a bad conductor of electricity because:
  - (a) In solid NaCl, there are no ions.
  - (b) Solid NaCl is a covalent compound.
  - (c) In solid NaCl, there is no mobility in ions.
  - (d) In solid NaCl, there are no attractive forces.

#### INTEGER ANSWER TYPE

9. How many species contain ionic as well as covalent bond(s)?  
 $CsBr_3$ ,  $NaOH$ ,  $NH_4Cl$ ,  $Na_2CO_3$ ,  $CaSO_4$ ,  $KNO_3$ ,  $KI_3$ ,  $CaC_2$ ,  $Ba(OH)_2$
10. The magnitude of electrovalency of group – 1 metals is.....

#### Section P: LATTICE ENTHALPY AND FAJAN'S RULE

#### SINGLE CORRECT ANSWER TYPE

1. Which set of compounds in the following pairs has the higher lattice energy?
  - (i)  $KCl$  or  $MgO$       (ii)  $LiF$  or  $LiBr$       (iii)  $Mg_3N_2$  or  $NaCl$
  - (a)  $KCl$ ,  $LiBr$ ,  $Mg_2N_2$
  - (b)  $MgO$ ,  $LiBr$ ,  $Mg_3N_2$
  - (c)  $MgO$ ,  $LiF$ ,  $NaCl$
  - (d)  $MgO$ ,  $LiF$ ,  $Mg_3N_2$
2. The incorrect order of lattice energy is:

- (a)  $\text{AlF}_3 > \text{MgF}_2$   
 (b)  $\text{Li}_3\text{N} > \text{Li}_2\text{O}$   
 (c)  $\text{NaCl} > \text{LiF}$   
 (d)  $\text{TiC} > \text{ScN}$
3. From the following sequence, calculate the lattice energy of  $\text{AB(s)}$  in  $\text{KJ/mole}$  :-  
 $\text{A(s)} \rightarrow \text{A}^+ (\text{g}) + \text{e}^- ; \Delta H = 610 \text{ KJ mol}^{-1}$   
 $\text{B(g)} + \text{e}^- \rightarrow \text{B}^- (\text{g}) ; \Delta H = -260 \text{ KJ mol}^{-1}$   
 $\text{A(s)} + \text{B(g)} \rightarrow \text{AB(s)} ; \Delta H = -569 \text{ KJ mol}^{-1}$   
 (a)  $-219$   
 (b)  $-919$   
 (c)  $+1539$   
 (d)  $+301$
4. The magnitude of the lattice energy of an ionic solid increases if:-  
 (a) The ions are large.  
 (b) The ions are small.  
 (c) The ions are of equal size.  
 (d) Charges on the ions are small.
5. Lattice energy is maximum in which of the following?  
 (a)  $\text{NaF}$   
 (b)  $\text{MgO}$   
 (c)  $\text{ScN}$   
 (d)  $\text{CsI}$
6. Among  $\text{LiCl}$ ,  $\text{BeCl}_2$ ,  $\text{BCl}_3$  and  $\text{CCl}_4$ , the covalent character follows the order:  
 (a)  $\text{LiCl} < \text{BeCl}_2 > \text{BCl}_3 > \text{CCl}_4$   
 (b)  $\text{LiCl} > \text{BeCl}_2 < \text{BCl}_3 < \text{CCl}_4$   
 (c)  $\text{LiCl} < \text{BeCl}_2 < \text{BCl}_3 < \text{CCl}_4$   
 (d)  $\text{LiCl} > \text{BeCl}_2 > \text{BCl}_3 > \text{CCl}_4$
7. The correct order of decreasing polarizability of ions is:  
 (a)  $\text{Cl}^- > \text{Br}^- > \text{I}^- > \text{F}^-$   
 (b)  $\text{F}^- > \text{I}^- > \text{Br}^- > \text{Cl}^-$   
 (c)  $\text{F}^- > \text{Cl}^- > \text{Br}^- > \text{I}^-$   
 (d)  $\text{I}^- > \text{Br}^- > \text{Cl}^- > \text{F}^-$
8. According to Fajan's rules, necessary condition to have more covalent character is:  
 (a) Small cation and large anion  
 (b) Small cation and small anion  
 (c) Large cation and large anion  
 (d) Large cation and small anion
9. Which of the following is correct order of ionic character?  
 (a)  $\text{NaCl} < \text{CuCl}$   
 (b)  $\text{NaCl} < \text{KCl}$   
 (c)  $\text{Li}_2\text{CO}_3 > \text{Ag}_2\text{CO}_3$   
 (d)  $\text{BeSO}_4 > \text{ZnSO}_4$

10. Which of the following has maximum covalent character?
- $\text{PbF}_4$
  - $\text{BiF}_5$
  - $\text{AlF}_3$
  - $\text{TlF}_3$
11. Among the following, the maximum covalent character is shown by the compound:-
- $\text{AlCl}_3$
  - $\text{MgCl}_2$
  - $\text{KCl}$
  - $\text{SnCl}_2$

### MULTIPLE CORRECT ANSWER TYPE

12. Polarisation may be called as the distortion of the shape of an anion an adjacently placed cation. Which of the following statements is/are not correct?
- Low cation is likely to bring about a large degree of low radius.
  - A large cation is likely to bring about a large degree of polarisation.
  - High polarization is brought about by a cation of high charge.
  - A small anion is likely to undergo a large degree of Polarisation.
13. Which statement is/are correct?
- Higher is the Polarisation, higher will be relative solubility in non – polar solvent.
  - Higher is the Polarisation, more will be the electrical conductivity.
  - With increase in Polarisation, covalent character increases.
  - Higher is the Polarisation in metal oxide, more will be ionic character.

### INTEGER ANSWER TYPE

14. How many given ions pseudo noble gas configuration?  
 $\text{Li}^+, \text{K}^+, \text{Mg}^{2+}, \text{Ga}^{3+}, \text{Zn}^{2+}, \text{Hg}^{2+}, \text{Cu}^+, \text{Pb}^{2+}, \text{Ca}^{2+}$
15. What is period number of cation of element of group number 2, which has maximum polarizing power in its group?

### Section Q: INERT PAIR EFFECT

#### SINGLE CORRECT ANSWER TYPE

1. Inert pair effect is the reluctance of electrons of:
- s – orbital to take part in bonding.
  - p – orbital to take part in bonding.
  - d – orbital to take part in bonding.
  - f – orbital to take part in bonding.
2. Which of the following can act as an oxidiser?
- $\text{PbO}_2$
  - $\text{Bi}_2\text{O}_3$

- (c) SnO  
(d) None of these
3. The stability of dihalides of Si, Ge, Sn and Pb increases in the sequence
- (a)  $\text{GeX}_2 < \text{SiX}_2 < \text{SnX}_2 < \text{PbX}_2$   
(b)  $\text{SiX}_2 < \text{GeX}_2 < \text{PbX}_2 < \text{SnX}_2$   
(c)  $\text{SiX}_2 < \text{GeX}_2 < \text{SnX}_2 < \text{PbX}_2$   
(d)  $\text{PbX}_2 < \text{SnX}_2 < \text{GeX}_2 < \text{SiX}_2$
4. Which of the following order is incorrect?
- (a) Oxidizing power order :  $\text{SiCl}_4 < \text{SnCl}_4 < \text{PbCl}_4$   
(b) Reducing power order :  $\text{CO} > \text{CO}_2$   
(c) Oxidizing power order :  $\text{PbO}_2 > \text{SnO}_2$   
(d) Reducing power order :  $\text{PbO} > \text{SnO}$
5. Which of the following does not exist?
- (a)  $\text{BI}_3$   
(b)  $\text{BiI}_3$   
(c)  $\text{PbI}_4$   
(d)  $\text{HgI}_2$

### MULTIPLE CORRECT ANSWER TYPE

6. Correct stability order of metal cation is/are:
- (a)  $\text{Pb}^{2+} < \text{Sn}^{2+}$   
(b)  $\text{Pb}^{4+} < \text{Pb}^{2+}$   
(c)  $\text{Sn}^{4+} < \text{Sn}^{2+}$   
(d)  $\text{Pb}^{4+} < \text{Sn}^{4+}$
7. Choose the incorrect option(s):
- (a) PbO is a powerful oxidiser.  
(b)  $\text{SnCl}_4$  is a powerful reducer.  
(c)  $\text{BiF}_3$  unstable.  
(d)  $\text{PbI}_4$  is more stable than  $\text{PbI}_2$

### MATRIX MATCH TYPE

8. Match the Column:

#### Column – I

#### (Property)

- (a) Cation with pseudo inert gas configuration  
(b) Cation with  $(18 + 2)$  electronic configuration  
(c) Strongest reducing agent out of given four ions  
(d) Cation which form most covalent chloride.

#### Column – II

#### (Ions)

- (P)  $\text{Sn}^{+2}$   
(Q)  $\text{Sn}^{+4}$   
(R)  $\text{Pb}^{+2}$   
(S)  $\text{Bi}^{+3}$

### INTERGER ANSWER TYPE

9. How many mole of diatomic gas is produced in reaction :  $2\text{PbO}_2 \xrightarrow{\Delta}$

10. The magnitude of oxidation state of Tl in  $TlI_3$  is....

## Section R: THERMAL STABILITY & SOLUBILITY

### SINGLE CORRECT ANSWER TYPE

- Which of the following is thermally most stable?
  - $BeSO_4$
  - $CaSO_4$
  - $SrSO_4$
  - $BaSO_4$
- When following compounds are heated separately in closed container of same volume at nearly same temperature such that each of the following decomposes at least to some extent, then one mole of which compound generates maximum pressure at same temperature?
  - $BaCO_3$
  - $LiNO_3$
  - $(NH_4)_2CO_3$
  - $(NH_4)_2C_2O_4$
- Which of the compound is most soluble in water
  - $AgF$
  - $AgCl$
  - $AgBr$
  - $AgI$
- Solubility of alkali metal fluorides:
  - Increases down the group.
  - Decreases down the group.
  - First increases then decreases down the group
  - First decreases then increases down the group.
- Extent of hydration is maximum in which ion?
  - $Ce_{(aq)}^{4+}$
  - $La_{(aq)}^{3+}$
  - $Ba_{(aq)}^{2+}$
  - $Cs_{(aq)}^{+}$
- Select correct order of thermal stability:
  - $Li_2O > Na_2O > Rb_2O$
  - $Rb_2O > Na_2O > Li_2O$
  - $Na_2O > Rb_2O > Li_2O$
  - $Rb_2O > Li_2O > Na_2O$

### MULTIPLE CORRECT ANSWER TYPE

- Which of the following give  $O_2$  (g) on thermal decomposition?
  - $Pb_2O_3$

- (b)  $\text{Pb}_3\text{O}_4$   
 (c)  $\text{PbO}_4$   
 (d)  $\text{Na}_2\text{O}$
8. Which of the following give  $\text{CO}_2(\text{g})$  on thermal decomposition?
- (a)  $\text{Ag}_2\text{CO}_3$   
 (b)  $\text{ZnCO}_3$   
 (c)  $\text{CaCO}_3$   
 (d)  $\text{Li}_2\text{CO}_3$

### INTEGER ANSWER TYPE

9. How many moles of  $\text{CO}_2$  are obtained on thermal decomposition of 1 mole  $\text{Na}_2\text{CO}_3$ ?
10. How many given metal carbonates are water soluble?  
 $\text{Ag}_2\text{CO}_3$ ,  $\text{Na}_2\text{CO}_3$ ,  $\text{FeCO}_3$ ,  $\text{ZnCO}_3$ ,  $\text{K}_2\text{CO}_3$ ,  $\text{Rb}_2\text{CO}_3$ ,  $\text{BaCO}_3$ ,  $\text{Cs}_2\text{CO}_3$ ,  $\text{CaCO}_3$

### WEAK FORCES

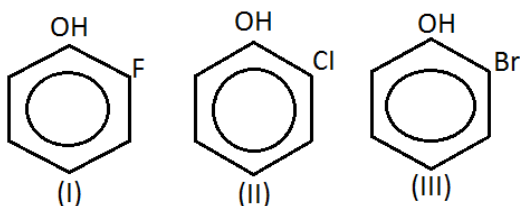
1. Which of the following gas maximum boiling point?
- (a)  $\text{BF}_3$   
 (b)  $\text{BCl}_3$   
 (c)  $\text{BBr}_3$   
 (d)  $\text{BI}_3$
2. Correct order of boiling points
- (a)  $\text{BF}_3 > \text{BMe}_3$   
 (b)  $\text{NF}_3 > \text{NMe}_3$   
 (c)  $\text{CCl}_4 > \text{SiCl}_4$   
 (d)  $\text{He} > \text{H}_2$
3. Which of the following interaction is strongest?
- (a)  $\text{Na}^+ \dots \text{OH}_2$   
 (b)  $\text{Na}^+ \dots \text{C}_6\text{H}_6$   
 (c)  $\text{Li}^+ \dots \text{OH}_2$   
 (d)  $\text{Li}^+ \dots \text{C}_6\text{H}_6$
4. Solubility of  $\text{Cl}_2$  in water is due to:
- (a) Dipole – Dipole interaction.  
 (b) Dipole – Induced dipole interaction.  
 (c) Ion – Dipole interaction.  
 (d) Ion – Induced dipole interaction.
5. Which of the following does not form clathrates?
- (a) Xe  
 (b) Rn  
 (c) Kr  
 (d) He
6. The molecular size of ICl and  $\text{Br}_2$  is approximately same, but b.p. of ICl is about  $40^\circ\text{C}$  higher than that of  $\text{Br}_2$ . It is



because:

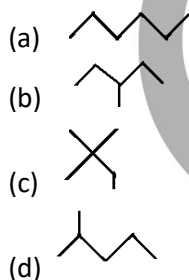
- (a) ICl bond is stronger than Br – Br bond.
- (b) IE of iodine < IE of bromine.
- (c) ICl is polar while Br<sub>2</sub> is nonpolar.
- (d) I has smaller size than Br.

7. The correct order of boiling point is:



- (a) I > II > III
- (b) III > II > I
- (c) II > I > III
- (d) III > I > II

8. Which is most volatile among given?



### INTEGER ANSWER TYPE

- 9. The magnitude of Dipole – induced dipole interaction energy is inversely proportional to  $r^x$  ( $r$  = distance between polar and non – polar molecule).  $x$  is equal to.....
- 10. What is period number of least water soluble noble gas?

### Section R: HYDROGEN BONDING

- 1. The correct order of strength of H – bond in the following compound is:
  - (a) H<sub>2</sub>O > H<sub>2</sub>O<sub>2</sub> > HF
  - (b) HF > H<sub>2</sub>O<sub>2</sub> > H<sub>2</sub>O
  - (c) HF > H<sub>2</sub>O < H<sub>2</sub>O<sub>2</sub>
  - (d) HF > H<sub>2</sub>O > H<sub>2</sub>O<sub>2</sub>
- 2. Which of the following is most volatile?
  - (a) HF
  - (b) HCl
  - (c) HBr
  - (d) HI

3. Density of water is more than that of ice at 25°C. Which of the following is responsible for the above order?
- Intermolecular H – bonding
  - Intermolecular H – bonding
  - Ion – ion interaction
  - Ion – dipole interaction.
4. Which among the following has maximum boiling point?
- $\text{CH}_3 - \text{CH}_2 - (\text{CH}_2)_2 - \text{CH}_2\text{OH}$   
 $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \underset{\text{OH}}{\text{CH}} - \text{CH}_3$
  - $\text{CH}_3 - \text{CH}_2 - \underset{\text{OH}}{\text{CH}} - \text{CH}_2 - \text{CH}_3$
  - $\text{CH}_3 - \underset{\text{OH}}{\text{C}}(\text{CH}_3) - \text{CH}_2 - \text{CH}_3$
  - $\text{CH}_3 - \underset{\text{OH}}{\text{C}}(\text{CH}_3) - \text{CH}_2 - \text{CH}_3$

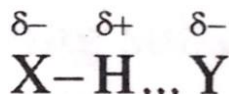
### MULTIPLE CORRECT ANSWER TYPE

5. The correct order of decreasing boiling points is/are:
- $\text{NH}_3 > \text{PH}_3$
  - $\text{H}_2\text{O} > \text{H}_2\text{Te}$
  - $\text{CH}_3\text{COOH} > \text{CH}_3\text{OCH}_3$
  - $\text{CH}_4 > \text{GeH}_4 > \text{SiH}_4$

### COMPREHENSION TYPE

#### Paragraph for Question no 6 & 7

Hydrogen bond is the term given to the relatively weak secondary interaction between a hydrogen atom bound to an electronegative atom and another atom which is also generally electronegative and which has one or more lone pairs and can thus act as a base. We can give the following generalized representation of a hydrogen bond.



Bond dissociation energy of H – bond ranges from 8 to 42 kJ/mol, and the most commonly encountered hydrogen bonds are O – H .... O, N – H....O and F – H...F.

6. Strongest H – bond among given is:
- $\text{H} - \text{Cl} \cdots \text{H} - \text{Cl}$
  - $\text{F}^- \cdots \text{H} - \text{F}$
  - $\text{S} \cdots \text{H} - \text{O}$
  - $\text{N} \cdots \text{H} - \text{N}$

7. Which of the following interaction has energy between 8 – 42 kJ/mol?

- (a)  $\text{Na}^+ : \text{CCl}_4$
- (b)  $\text{CHCl}_3 : \text{Br}^-$
- (c)  $\text{C}_6\text{H}_6 : \text{CCl}_4$
- (d)  $\text{H}_2\text{O} : \text{HCN}$

### INTEGER ANSWER TYPE

- 8. How many membered chelate ring is formed in ortho – nitrophenol due to intermolecular hydrogen – bonding?
- 9. How many atoms are capable of forming hydrogen bond in a molecule of Boric acid,  $\text{B}(\text{OH})_3$  with other molecules of  $\text{B}(\text{OH})_3$ ?

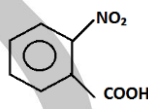
### SUBJECTIVE TYPE

10. Select the species forming intermolecular hydrogen bond with their other molecules, forming intermolecular hydrogen bond and forming no hydrogen bond:

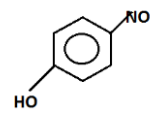
(1)  $\text{C}_2\text{H}_5\text{OH}$

(2)  $\text{Cl}_3\text{C} - \text{CH}(\text{OH})_2$

(3)



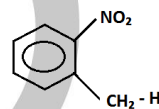
(4)



(5)  $\text{CH}_3\text{NH}_2$

(6)  $\text{CH}_3\text{COOH}$

(7)



(8)  $\text{H}_2\text{O}$

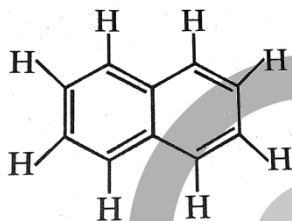
(9) Ortho – fluorophenol

(10)  $\text{CH}_3 - \text{CHO}$

### Section S:

- 1. Polarity in a molecule and hence the dipole moment depends primarily on electronegativity of the constituent atoms and shape of a molecule. Which of the following has the highest dipole moment?
  - (a)  $\text{CO}_2$
  - (b)  $\text{HI}$
  - (c)  $\text{H}_2\text{O}$
  - (d)  $\text{SO}_2$
- 2. Hydrogen bonds are formed in many compounds e.g.,  $\text{H}_2\text{O}$ ,  $\text{HF}$ ,  $\text{NH}_3$ . The boiling point of such compounds depends to a large extent on the strength of hydrogen bond and the number of hydrogen bonds. The correct decreasing order of the boiling points of above compounds is:
  - (a)  $\text{HF} > \text{H}_2\text{O} > \text{NH}_3$
  - (b)  $\text{H}_2\text{O} > \text{HF} > \text{NH}_3$
  - (c)  $\text{NH}_3 > \text{HF} > \text{H}_2\text{O}$
  - (d)  $\text{NH}_3 > \text{H}_2\text{O} > \text{HF}$
- 3. In  $\text{PO}_4^{3-}$  ion the formal charge on the oxygen atom of P – O bond is
  - (a) + 1
  - (b) – 1
  - (c) – 0.75

- (d) + 0.75
4. In  $\text{NO}_3^-$  ion, the number of bond pairs and lone pairs of electrons on nitrogen atom are
- 2, 2
  - 3, 1
  - 1, 3
  - 4, 0
5. Which of the following species has tetrahedral geometry?
- $\text{BH}_4^-$
  - $\text{NH}_2^-$
  - $\text{CO}_3^{2-}$
  - $\text{H}_3\text{O}^+$
6. Number of  $\pi$  bond and  $\sigma$  bond in the following structure is:



- 6, 19
  - 4, 20
  - 5, 19
  - 5, 20
7. Which molecule/ion out of the following does not contain unpaired electrons?
- $\text{N}_2^+$
  - $\text{O}_2$
  - $\text{O}_2^{2-}$
  - $\text{B}_2$
8. In which of the following substances will hydrogen bond be strongest?
- HCl
  - $\text{H}_2\text{O}$
  - HI
  - $\text{H}_2\text{S}$
9. Which of the following order of energies of molecular orbitals of  $\text{N}_2$  is correct?
- $(\pi_{2p_y}) < (\sigma_{2p_z}) < (\pi^*_{2p_x}) \cong (\pi^*_{2p_y})$
  - $(\pi_{2p_y}) > (\sigma_{2p_z}) > (\pi^*_{2p_x}) \cong (\pi^*_{2p_y})$
  - $(\pi_{2p_y}) < (\sigma_{2p_z}) > (\pi^*_{2p_x}) \cong (\pi^*_{2p_y})$
  - $(\pi_{2p_y}) > (\sigma_{2p_z}) < (\pi^*_{2p_x}) > (\pi^*_{2p_y})$
10. Which of the following statement is not correct from the view point of molecular orbital theory?
- $\text{Be}_2$  is not a stable molecule.
  - $\text{He}_2$  is not stable but  $\text{He}_2^+$  is expected to exist.

- (c) Bond strength of  $N_2$  is maximum amongst the mononuclear diatomic molecules belonging to the second period.
- (d) The order of energies of molecular orbitals in  $N_2$  molecule is:
- $$\sigma_{2s} < \sigma_{2s}^* < \sigma_{2p_z} < (\pi_{2p_x} = \pi_{2p_y}) < (\pi_{2p_x}^* = \pi_{2p_y}^*) < \sigma_{2p_z}^*$$

### MULTIPLE CORRECT ANSWER TYPE

11. Which of the following species have the same shape?
- $CO_2$
  - $CCl_4$
  - $O_3$
  - $NO_2^-$
12. Which of the following statements are correct about  $CO_3^{2-}$ ?
- The hybridisation of central atom is  $sp^3$ .
  - Its resonance structure has one C – O single bond and two C = O double bonds.
  - The average formal charge on each oxygen atom is 0.67 units.
  - All C – O bond lengths are equal.
13. Which of the following statements are not correct?
- NaCl being an ionic compound is a good conductor of electricity in the solid state.
  - In canonical structures, there is a difference in the arrangement of atoms.
  - Hybrid orbitals form stronger bonds than pure orbitals.
  - VSEPR Theory can explain the square planar geometry of  $XeF_4$ .
14. Species having same bond order are:
- $N_2$
  - $N_2^-$
  - $F_2^+$
  - $O_2^-$
15. CO is isoelectronic with:
- $NO^+$
  - $N_2$
  - $SnCl_2$
  - $NO_2^-$

### COMPREHENSION

#### Paragraph for question no. 16 to 19

The electronic configurations of three elements, A, B and C are given below. Answer the questions 14 to 17 on the basis of these configurations.

A	$1s^2$	$2s^2$	$2p^6$		
B	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^3$
C	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^5$

16. Stable form of A may be represented by the formula:

- (a) A  
(b) A<sub>2</sub>  
(c) A<sub>3</sub>  
(d) A<sub>4</sub>
17. Stable form of C may be represented by the formula:  
(a) C  
(b) C<sub>2</sub>  
(c) C<sub>3</sub>  
(d) C<sub>4</sub>
18. The molecular formula of the compound formed from B and C will be  
(a) BC  
(b) B<sub>2</sub>C  
(c) BC<sub>2</sub>  
(d) BC<sub>3</sub>
19. The bond between B and C will be  
(a) Ionic  
(b) Covalent  
(c) Hydrogen  
(d) Coordinate

### Paragraph for question no. 20 to 22

Comprehension given is followed by some multiple choice questions. Each question has one correct option. Choose the correct option.

Molecular orbitals are formed by the overlap of atomic orbitals. Two atomic orbitals combine to form two molecular orbitals called bonding molecular orbital (BMO) and anti bonding molecular orbital (ABMO). Energy of anti bonding orbital is raised above than the parent atomic orbitals. Energies of various orbitals for elements hydrogen to nitrogen increase in the order:

$\sigma_{1s} < \sigma_{1s}^* < \sigma_{2s} < \sigma_{2s}^* < (\pi_{2p_x} \approx \pi_{2p_y}) < \sigma_{2p_z} < (\pi_{2p_x}^* \approx \pi_{2p_y}^*) < \sigma_{2p_z}^*$  and for oxygen fluorine order of energy of molecular orbitals is given below:

$\sigma_{1s} < \sigma_{1s}^* < \sigma_{2s} < \sigma_{2s}^* < \sigma_{2p_x} < (\pi_{2p_x} \approx \pi_{2p_y}) < (\pi_{2p_x}^* \approx \pi_{2p_y}^*) < \sigma_{2p_z}^*$  different atomic orbitals of one atom combine with those atomic orbitals of the second atom which have comparable energies and proper orientation. Further, if the overlapping is head on, the molecular orbital is called 'Sigma'. ( $\sigma$ ) and if the overlap is lateral, the molecular orbital is called pi', ( $\pi$ ). The molecular orbitals are filled with electrons according to the same rules as followed for filling of atomic orbitals. However, the order for filling is not the same for all molecule or their ions. Bond order is one of the most important parameters to compare the strength of bonds.

20. Which of the following statements is correct?  
(a) In the formation of dioxygen from oxygen atoms, 10 molecular orbitals will be formed.  
(b) All the molecular orbitals in the dioxygen will be completely filled.

- (c) Total number of bonding molecular orbitals will not be same as total number of anti-bonding molecular orbitals in dioxygen.
- (d) Number of filled bonding orbitals will be same as number of filled anti bonding orbitals.
21. Which of the following pair is expected to have the same bond order?
- (a)  $O_2$ ,  $N_2$
- (b)  $O_2^+$ ,  $N_2^-$
- (c)  $O_2^-$ ,  $N_2^+$
- (d)  $O_2^-$ ,  $N_2^-$
22. In which of the following molecules,  $\sigma_{2p_z}$  molecular orbital is filled after  $\pi_{2p_x}$  and  $\pi_{2p_y}$  molecular orbitals?
- (a)  $O_2$
- (b)  $Ne_2$
- (c)  $N_2$
- (d)  $F_2$

### MATRIX MATCH TYPE

23. Match the species in Column I with the type of hybrid orbitals in Column II.

Column I

- (a)  $SF_4$
- (b)  $IF_5$
- (c)  $NO_2^+$
- (d)  $NH_4^+$

Column II

- (P)  $sp^3d^2$
- (Q)  $sp^3d$
- (R)  $sp^3$
- (S)  $sp$

24. Match the species in Column I with the geometry/shape in Column II.

Column I

- (a)  $H_3O^+$
- (b)  $HC \equiv CH$
- (c)  $ClO_2^-$
- (d)  $NH_4^+$

Column II

- (P) Linear
- (Q) Angular
- (R) Tetrahedral
- (S) Trigonal bipyramidal
- (T) Pyramidal

25. Match the items given in Column I with examples given in Column II.

Column I

- (a) Hydrogen bond
- (b) Resonance
- (c) Ionic solid
- (d) Covalent solid

Column II

- (P) C
- (Q) LiF
- (R)  $H_2$
- (S) HF
- (T)  $O_3$



## EXERCISE – II (READY FOR CHALLENGES)

### LEVEL-I

1. When two atoms combine to form a molecules:
  - (A) Energy is released
  - (B) Energy is absorbed
  - (C) Energy is neither released nor absorbed
  - (D) Energy may either released or aborbed
2. The combination of atoms occur because they want :
  - (A) To decrease number of electrons in the outer most orbit
  - (B) To attain an inert gas configuration
  - (C) To increase number of electrons in the outer most orbit
  - (D) To attain 18 electrons in the outermost orbit
3. Which condition favours the bond formation:
  - (A) Maximum attraction and maximum potential energy
  - (B) Minimum attraction and minimum potential energy
  - (C) Minimum potential energy and maximum attraction
  - (D) None of the above
4. An electrovalent bond or ionic bond is formed between:
  - (A) Two electronegative atoms
  - (B) Two metals
  - (C) Electropositive and electronegative atoms
  - (D) Two electropositive atoms
5. Most favourable conditions for electrovalent bonding are:
  - (A) Low ionisation potential of one atom and high electron affinity of the other atom
  - (B) High electron affinity and high ionisation potential of both the atoms
  - (C) low electron affinity and low ionisation potential of both the atoms
  - (D) High ionisation potential of one atom and low electrons afinity of the other atom
6. The electronegativity of cesium is 0.7 and fluorine is 4.0. The bond formed between the two is:
  - (A) Covalent
  - (B) Electrovalent/ionic
  - (C) Coordinate
  - (D) Metallic
7. Electrovalent bond or ionic bond is formed by:
  - (A) Sharing of electons
  - (B) Donaton of electrons
  - (C) Transfer of electrons
  - (D) None of these
8. Element X is strongly electropositive and is strongly electronegative. Both are univalent. The compound formed would be:
 

(A) $X^+ Y^-$	(B) $X - Y$
(C) $X^- Y^+$	(D) $X \rightarrow Y$
9. Element a has 3 electrons in the outermost orbit and element B has 6 eletrons in the outermost orbit. The formula of the compound formed between A and B would be:

- (A) High melting points and non-directional bonds  
 (B) High melting points and low boiling points  
 (C) Directional bonds and low boiling points  
 (D) High solubilities in polar and non-polar point
10. Ionic compounds in general possess both:  
 (A) High melting points and non-directional bonds  
 (B) High melting points and low boiling points  
 (C) Directional bonds and low boiling points  
 (D) High solubilities in polar and non-polar solvents
11. Conditions for ionic bond formation is/are:  
 (1) Small cation, large anion  
 (2) low IP of cation, high electron affinity of anion  
 (3) Large cation, small anion and less charge  
 (4) Less lattice energy  
 Correct answer is:  
 (A) 1, 4 (B) 2, 3 and 4  
 (C) 2 and 3 (D) 1, 2
12. An ionic compound  $A^+ B^-$  is most likely to be formed when:  
 (A) Ionization energy of A is low  
 (B) Electron affinity of B is high  
 (C) Electron affinity of B is low  
 (D) Both (A) and (B)
13. Electrovalent compounds or ionic compounds do not show stereoisomerism. The reason is:  
 (A) Presence of ions  
 (B) Strong electro static force of attraction  
 (C) Brittleness  
 (D) Non-directional nature of ionic bond
14. Compound of a metal 'M' is  $M_2O_3$ . The formula of its nitride will be:  
 (A)  $M_3N$  (B) MN  
 (C)  $M_3N_2$  (D)  $M_2N_3$
15. Polarity between two atoms is depend on:  
 (A) Ionization potential of element  
 (B) Electronegativity of element  
 (C) Electronic configuration of element  
 (D) No. of unpaired electrons
16. As compared to covalent compounds electrovalent compounds generally possess:  
 (A) High m.p. and high b.p.  
 (B) Low m.p. and low b.p.  
 (C) Low m.p. and high b.p.  
 (D) high m.p. and low b.p.
17. An ionic compound  $A^+ B^-$  is most likely to be formed when:  
 (A) The ionization energy of A is high and electron affinity of B is low  
 (B) The ionization energy of A is low and electron affinity of B is high  
 (C) Both, the ionization energy of A and electron affinity of B are high  
 (D) Both, the ionization energy of A and electron affinity of B are low
18. In which of the following solvents, KI has highest solubility? The dielectric constant ( $\epsilon$ ) of each liquid is given in parentheses:

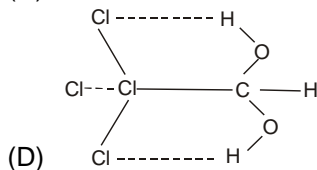
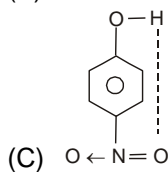
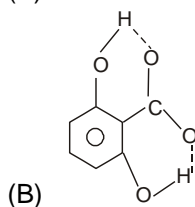
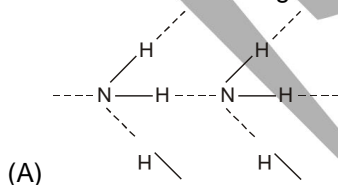
- (A)  $C_6H_6$  ( $\epsilon = 0$ )  
 (B)  $(CH_3)_2CO$  ( $\epsilon = 2$ )  
 (C)  $CH_3OH$  ( $\epsilon = 32$ )  
 (D)  $CCl_4$  ( $\epsilon = 0$ )
19. Solid NaCl is a bad conductor of electricity because:  
 (A) In solid NaCl there are no ions  
 (B) Solid NaCl is covalent  
 (C) In solid NaCl there is no mobility of ions  
 (D) In solid NaCl there are no electrons
20. Molten sodium chloride conducts electricity due to the presence of :  
 (A) Free electrons  
 (B) Free molecules  
 (C) Free sodium and chloride atoms  
 (D) Free sodium and chloride ions
21. Ionic reactions occur in :  
 (A) Aqueous solution and organic solvents of high polarity  
 (B) Non-polar or solvents of low polarity  
 (C) Gaseous state  
 (D) Solid state
22. Which of the following statements about LiCl and NaCl is wrong:  
 (A) LiCl has lower melting point than NaCl  
 (B) LiCl dissolves more in organic solvents whereas NaCl does not  
 (C) LiCl would ionise in water more than NaCl  
 (D) Fused LiCl would be less conducting than fused NaCl
23. A covalent bond is possible between:  
 (A) Similar atoms  
 (B) Dissimilar atoms  
 (C) Similar and dissimilar atoms  
 (D) Similar molecules
24. Which of the following statement is not true about covalent compounds?  
 (A) They may exhibit space isomerism  
 (B) They have low melting and boiling points  
 (C) They show ionic reactions  
 (D) They show molecular reactions
25. In a triple bond there is sharing of:  
 (A) 3 electrons (B) 4 electrons (C) Several electrons (D) 6 electrons
26. Which of the following bonds will have directional character?  
 (A) Ionic bond  
 (B) Metallic bond  
 (C) Covalent bond  
 (D) Both covalent & metallic
27. The boiling point of ICl is nearly  $40^\circ C$  higher than that of  $Br_2$  although the two substances have the same relative molecular mass. This is because:  
 (A) ICl is ionic compound  
 (B) I-Cl bond is stronger than Br-Br bond  
 (C) ICl is polar covalent molecular while  $Br_2$  is non polar

- (D) Ionization energy IP of Iodine is less than that of Br
28. The type of bond formed between two electronegative atoms would be:  
 (A) Covalent (B) Ionic (C) Coordinate (D) All the above
29. Correct order of covalent character of alkaline earth metal chloride in:  
 (A)  $\text{BeCl}_2 < \text{MgCl}_2 < \text{CaCl}_2 < \text{SrCl}_2$   
 (B)  $\text{BeCl}_2 < \text{CaCl}_2 < \text{SrCl}_2 < \text{MgCl}_2$   
 (C)  $\text{BeCl}_2 > \text{MgCl}_2 > \text{CaCl}_2 > \text{SrCl}_2$   
 (D)  $\text{SrCl}_2 > \text{BeCl}_2 > \text{CaCl}_2 > \text{MgCl}_2$
30. Which of the compound is least soluble in water:  
 (A) AgF (B) AgCl (C) AgBr (D) AgI
31. Which pair in the following has maximum and minimum ionic character respectively:  
 (A) LiCl, RbCl (B) RbCl,  $\text{BeCl}_2$  (C)  $\text{BeCl}_2$ , RbCl (D) AgCl, RbCl
32. The M.P. of  $\text{SnCl}_4$  is less than of  $\text{SnCl}_2$ , the suitable reason for the observed fact is:  
 (A) There is more charge on  $\text{Sn}^{+4}$   
 (B) The size of  $\text{Sn}^{+4}$  is small  
 (C) Ionic potential ( $\phi$ ) of  $\text{Sn}^{+4}$  is high  
 (D) The shape of  $\text{SnCl}_4$  is tetrahedral
33. Which of the following statements regarding covalent bond is not true?  
 (A) The electrons are shared between atoms  
 (B) The bond is non-directional  
 (C) The strength of the bond depends upon the extent of overlapping  
 (D) The bond formed may or may not be polar
34. Which of the following has least polarity in bond?  
 (A) H-F (B) H-Cl (C) H-O (D) H-S
35. Among LiCl,  $\text{BeCl}_2$ ,  $\text{BCl}_3$  and  $\text{CCl}_4$ , the covalent bond character follows the order:  
 (A)  $\text{LiCl} < \text{BeCl}_2 > \text{BCl}_3 > \text{CCl}_4$   
 (B)  $\text{LiCl} > \text{BeCl}_2 < \text{BCl}_3 < \text{CCl}_4$   
 (C)  $\text{LiCl} < \text{BeCl}_2 < \text{BCl}_3 < \text{CCl}_4$   
 (D)  $\text{LiCl} > \text{BeCl}_2 > \text{BCl}_3 > \text{CCl}_4$
36. Elements have electronegativities 1.5 and 3.0, bond formed between them would be:  
 (A) Ionic (B) Covalent (C) Co-ordinate (D) Metallic
37. The correct order of decreasing polarisable ions is:  
 (A)  $\text{Cl}^- > \text{Br}^- > \text{I}^- > \text{F}^-$   
 (B)  $\text{F}^- > \text{I}^- > \text{Br}^- > \text{Cl}^-$   
 (C)  $\text{F}^- > \text{Cl}^- > \text{Br}^- > \text{I}^-$   
 (D)  $\text{I}^- > \text{Br}^- > \text{Cl}^- > \text{F}^-$
38. According to Fajan's rules necessary condition to form covalent bond is:  
 (A) small cation and large anion  
 (B) small cation and small anion  
 (C) large cation and large anion  
 (D) large cation and small anion
39. Which is most ionic:  
 (A)  $\text{P}_2\text{O}_5$  (B) MnO  
 (C)  $\text{CrO}_3$  (D)  $\text{Mn}_2\text{O}_7$

40. The correct order of increasing covalent character of the following is:  
 (A)  $\text{SiCl}_4 < \text{AlCl}_3 < \text{CaCl}_2 < \text{KCl}$   
 (B)  $\text{KCl} < \text{CaCl}_2 < \text{AlCl}_3 < \text{SiCl}_4$   
 (C)  $\text{AlCl}_3 < \text{CaCl}_2 < \text{KCl} < \text{SiCl}_4$   
 (D) None of these
41. Which of the following is not a characteristic of a covalent compound:  
 (A) It has low melting point and boiling point  
 (B) It is formed between two atoms having not very small electronegativity difference  
 (C) They have no definite geometry  
 (D) They are generally insoluble in water
42. The order of decreasing polarity in the compounds:  
 $\text{CaO}$ ,  $\text{CsF}$ ,  $\text{KCl}$ ,  $\text{MgO}$  is-  
 (A)  $\text{CaO} > \text{CsF} > \text{KCl} > \text{MgO}$   
 (B)  $\text{MgO} > \text{KCl} > \text{CaO} > \text{CsF}$   
 (C)  $\text{KCl} > \text{CaO} > \text{CsF} > \text{MgO}$   
 (D)  $\text{CsF} > \text{KCl} > \text{CaO} > \text{MgO}$
43. Correct order of polarising power is:  
 (A)  $\text{Cs}^+ < \text{K}^+ < \text{Mg}^{2+} < \text{Al}^{3+}$   
 (B)  $\text{Al}^{3+} < \text{Mg}^{2+} < \text{K}^+ < \text{Cs}^+$   
 (C)  $\text{Mg}^+ < \text{Al}^{3+} < \text{K}^+ < \text{Cs}^+$   
 (D)  $\text{K}^+ < \text{Cs}^+ < \text{Mg}^{2+} < \text{Al}^{3+}$
44. Which of the following statements is correct :  
 (A)  $\text{HCl}$  is covalent both in aqueous solution and in the gaseous state  
 (B)  $\text{HCl}$  is covalent in the gaseous state but ionic in aqueous solution  
 (C)  $\text{HCl}$  is ionic both in the gaseous state and in aqueous solution  
 (D) None of the above
45. The cyanide ion,  $\text{CN}^-$  and  $\text{N}_2$  are isoelectronic But in contrast to  $\text{CN}^-$ ,  $\text{N}_2$  is chemically inert, because of:  
 (A) Low bond energy  
 (B) Absence of bond polarity  
 (C) Unsymmetrical electron distribution  
 (D) Presence of more number of electrons in bonding orbitals
46. In Co-ordinate bond, the acceptor atoms must essentially contain in its valency shell an orbital:  
 (A) With paired electron  
 (B) With single electron  
 (C) With no electron  
 (D) With three electron
47. Correct statement regarding this reaction  

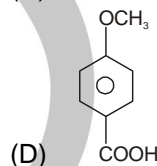
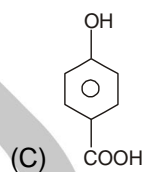
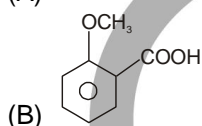
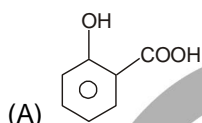
$$\text{BF}_3 + \text{NH}_3 \longrightarrow [\text{F}_3\text{B} \leftarrow \text{NH}_3]$$
  
 (A) Hybridisation of N is changed  
 (B) Hybridisation of B is changed  
 (C)  $\text{NH}_3$  act as a Lewis base  
 (D) (B) & (C) both
48. The correct statement for the reaction  
 (A) Hybridisation state is changed  
 (B) Bond angle increases  
 (C)  $\text{NH}_3$  act as a Lewis acid

- (D) Regular geometry is changed
49. The bonds present in  $\text{N}_2\text{O}_5$  are:  
 (A) Only ionic  
 (B) Covalent & coordinate  
 (C) Only covalent  
 (D) Covalent & ionic
50. The pair of compounds which can form a co-ordinate bond is:  
 (A)  $(\text{C}_2\text{H}_5)_3\text{B}$  and  $(\text{CH}_3)_3\text{N}$   
 (B)  $\text{HCl}$  and  $\text{HBr}$   
 (C)  $\text{BF}_3$  and  $\text{NH}_3$   
 (D) (A) and (C) both
51. In which compound coordinate bond is present:  
 (A)  $\text{NH}_3$  (B)  $\text{NH}_4\text{OH}$  (C)  $\text{H}_2\text{O}$  (D)  $\text{HCl}$
52. The hydrogen bond is strongest in:  
 (A)  $\text{O}-\text{H} \cdots \text{S}$   
 (B)  $\text{S}-\text{H} \cdots \text{O}$   
 (C)  $\text{F}-\text{H} \cdots \text{F}$   
 (D)  $\text{O}-\text{H} \cdots \text{O}$
53.  $\text{H}_2\text{O}$  boils at higher temperature than  $\text{H}_2\text{S}$ , because it is capable of forming:  
 (A) Ionic bonds  
 (B) Covalent bonds  
 (C) Hydrogen bonds  
 (D) Metallic bonds
54. In which of the following molecule, the shown hydrogen bond is not possible:



55. The correct order of volatility is :

- (A)  $\text{NH}_3 < \text{H}_2\text{O}$   
 (B) p-nitro phenol < o-nitro phenol  
 (C)  $\text{CH}_3\text{OH} > \text{CH}_3 - \text{O} - \text{CH}_3$   
 (D)  $\text{HF} > \text{HCl}$
56. Intramolecular H-bond:  
 (A) Decreases Volatility  
 (B) Increases melting point  
 (C) Increases viscosity  
 (D) Increases vapour pressure
57. Weakest hydrogen bond is:  
 (A)  $\text{O} - \text{H} \cdots \text{N}$  (B)  $\text{S} - \text{H} \cdots \text{S}$   
 (C)  $\text{F} - \text{H} \cdots \text{F}$  (D)  $\text{N} - \text{H} \cdots \text{N}$
58. Which of the following has strongest intra molecular hydrogen bonding:

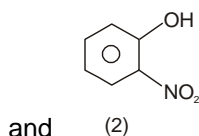
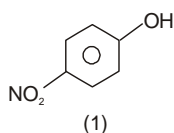


59. Acetic acid exists as dimer in benzene due to :  
 (A) Condensation reaction  
 (B) Hydrogen bonding  
 (C) Presence of carboxyl group  
 (D) None of the above
60. In which case hydrogen bond will not be observed -  
 (A)  $\text{H}_3\text{O}_2^-$  (B)  $\text{H}_2\text{O}$   
 (C)  $\text{HF}$  (D)  $\text{H}_3\text{O}^+$
61. The high boiling point of water is due to the presence of:  
 (A) Dative bond  
 (B) Covalant bond  
 (C) Hydrogen bond  
 (D) Vander waals bond
62. Maximum no. of hydrogen bonds formed by a water molecule in ice is:  
 (A) 4 (B) 3  
 (C) 2 (D) 1
63. Hydrogen bonding is formed in compounds containing hydrogen and:  
 (A) Highly electro-negative atoms  
 (B) Highly electro-positive atoms  
 (C) Metal atoms with d-orbitals occupies  
 (D) Metalloids
64. The intermolecular force in hydrogen fluoride is due to :



- (A) Dipole induced dipole interactions  
 (B) Dipole-dipole interactions  
 (C) Hydrogen bond  
 (D) None of these
65. The boiling point of p-nitrophenol is higher than that of o-nitrophenol because:  
 (A)  $\text{NO}_2$  group at p-position behaves in a different way from that at o-position  
 (B) intramolecular hydrogen bonding exists in p-nitrophenol  
 (C) there is intermolecular hydrogen bonding in p-nitrophenol  
 (D) p-nitrophenol has a higher molecular weight than o-nitrophenol
66. Density of ice is less than that of water because of :  
 (A) presence hydrogen bonding  
 (B) crystal modification of ice  
 (C) open porous structure of ice due to hydrogen bonding  
 (D) different physical states of these
67. KF combines with HF to form  $\text{KHF}_2$ . The compound contains the species:  
 (A)  $\text{K}^+$ ,  $\text{F}^-$  and  $\text{H}^+$   
 (B)  $\text{K}^+$ ,  $\text{F}^-$  and HF  
 (C)  $\text{K}^+$  and  $[\text{HF}_2]^-$   
 (D)  $[\text{KHF}]^+$  and  $\text{F}_2$
68. Which of the following compounds show intramolecular hydrogen bonding:  
 (1) o-nitrophenol                      (2) p-nitrophenol  
 (3) phenol                                (4) salicylaldehyde  
 (A) 1 & 2                                (B) 1 & 3  
 (C) 1 & 4                                (D) 2 & 3
69. The pair of molecules forming strongest hydrogen bonds are :  
 (A)  $\text{SiH}_4$  and  $\text{SiF}_6$   
 (B)  $\begin{array}{c} \text{CH}_3 - \text{C} - \text{CH}_3 \\ \parallel \\ \text{O} \end{array}$  and  $\text{CHCl}_3$   
 (C)  $\begin{array}{c} \text{H} - \text{C} - \text{OH} \\ \parallel \\ \text{O} \end{array}$  and  $\begin{array}{c} \text{CH}_3 - \text{C} - \text{OH} \\ \parallel \\ \text{O} \end{array}$   
 (D)  $\text{H}_2\text{O}$  and  $\text{H}_2$
70. The boiling point of a compound is raised by:  
 (A) intermolecular hydrogen bonding  
 (B) high volatility  
 (C) intramolecular hydrogen bonding  
 (D) non-polarity
71. The boiling point of methanol is greater than of methyl thiol because:  
 (A) There is intramolecular hydrogen bonding in methanol and ntermolecular hydrogen bonding in methyl thiol  
 (B) there is intermolecular hydrogen hydrogen bonding in methanol and no hydrogen bonding in methythiol  
 (C) There is no hydrogen bonding in methanol and intermolecular hydrogen bonding in methylthiol  
 (D) There is intramolecular hydrogen bonding in methanol and no hydrogen bonding in methythiol
72. Out of the two compounds shown below, the vapour pressure of (2) at a particular temperature is

expected to be:



- (A) Higher than that of (1)  
(B) Lower than that of (1)  
(C) Same as that of (1)  
(D) Can be higher or lower depending upon the size of the vessel
73. A sigma bond is formed by the overlapping of :  
(A) s-s orbital alone  
(B) s and p orbitals alone  
(C) s-s, s-p or p-p orbitals along internuclear axis  
(D) p-p orbital along the sides
74. Which overlapping is involved in HCl molecule:  
(A) s-s overlap (B) p-p overlap  
(C) s-d overlap (D) s-p overlap
75. Which of the following configuration shows second excitation state of Iodine:  
(A)   
(B)   
(C)   
(D)
76. Which of the following compound is formed in the second excitation state of sulphur atom:  
(A) SF<sub>4</sub> (B) SF<sub>6</sub> (C) SF<sub>2</sub> (D) None
77. The strength of bonds by 2s – 2s, 2p – 2p and 2p – 2s overlapping has the order:  
(A) s – s > p – p > s – p  
(B) s – s > p – s > p – p  
(C) p – p > s – p > s – s  
(D) p – p > s – s > p – s
78. The triple bond in ethyne is made up of:  
(A) Three sigma bonds  
(B) Three  $\pi$  - bonds  
(C) One sigma and two  $\pi$  bonds  
(D) Two sigma and one  $\pi$  bond
79. In which of the excitation state of chlorine ClF<sub>3</sub> is formed:  
(A) In ground state  
(B) In third excitation state  
(C) In first excitation state  
(D) In second excitation state
80. Variable covalency is exhibited by:  
(A) P and S (B) N and O  
(C) N and P (D) F and Cl
81. Which is not characteristic of  $\pi$ -bond:  
(A)  $\pi$ -bond is formed when a sigma bond already formed  
(B)  $\pi$ -bond are formed from hybrid orbitals

- (C)  $\pi$ - bond may be formed by the overlapping of p-orbitals  
 (D)  $\pi$ - bond results from lateral overlap of atomic orbital
82.  $\pi$  bond is formed:  
 (A) By overlapping of hybridised orbital  
 (B) Overlapping of co-axial p-p orbitals  
 (C) Head on overlapping of s – s orbitals  
 (D) By p – p collateral overlapping
83. Number of  $\sigma$  and  $\pi$  bonds present in are:  
 (A)  $10\sigma, 3\pi$  (B)  $10\sigma, 2\pi$  (C)  $9\sigma, 2\pi$  (D)  $8\sigma, 3\pi$
84. Which of the following fluorides does not exist?  
 (A)  $\text{NF}_5$  (B)  $\text{PF}_5$  (C)  $\text{AsF}_5$  (D)  $\text{SbF}_5$
85. p-p overlapping will be observed in the molecules of -  
 (A)  $\text{H}_2$  (B)  $\text{HBr}$  (C)  $\text{HCl}$  (D)  $\text{Cl}_2$
86. Nitrogen does not form  $\text{NF}_5$  because:  
 (A) Nitrogen is member of V group  
 (B) It contains no empty d-orbital  
 (C) The bond energy of is very high  
 (D) Inert pair effect exists in the molecule
87. The correct order of bond length is  
 (A)  $\text{C}-\text{C} < \text{C}=\text{C} < \text{C}\equiv\text{C}$   
 (B)  $\text{C}\equiv\text{C} < \text{C}=\text{C} < \text{C}-\text{C}$   
 (C)  $\text{C}=\text{C} < \text{C}\equiv\text{C} < \text{C}-\text{C}$   
 (D)  $\text{C}=\text{C} < \text{C}-\text{C} < \text{C}\equiv\text{C}$
88. Which of the following statements is not correct?  
 (A) Double bond is shorter than a single bond  
 (B)  $\sigma$  – bond is weaker than  $\pi$  bond  
 (C) Double bond is stronger than a single bond  
 (D) Covalent bond is stronger than a hydrogen bond
89. Fluorine molecule is formed by:  
 (A) the axial p-p orbital overlap  
 (B) the sideways p-p orbital overlap  
 (C) the s-s orbital overlap  
 (D) the s-p orbital overlap
90. Consider the following statements:  
 I. A sigma ( $\sigma$ ) bond is formed when two s-orbitals overlap  
 II. A pi ( $\pi$ ) bond is formed when two p-orbitals overlap axially  
 III. A  $\sigma$  - bond is weaker than  $\pi$  -bond  
 Which of the above statements is/are correct?  
 (A) I and II (B) II and III  
 (C) I alone (D) II alone
91. In the protonation of  $\text{H}_2\text{O}$ , change occurs in :  
 (A) Hybridisation state of oxygen  
 (B) Shape of molecule

- (C) Hybridisation and shape both  
 (D) None
92. The d-orbitals involved in  $sp^3d$  hybridisation is :  
 (A)  $d_{x^2-y^2}$  (B)  $d_{z^2}$   
 (C)  $d_{xy}$  (D)  $d_{xz}$
93. A  $sp^3$  hybrid orbital contains:  
 (A)  $\frac{3}{4}$  s-character (B)  $\frac{1}{4}$  p-character (C)  $\frac{3}{4}$  p-character (D)  $\frac{1}{2}$  s-character
94. In the compound  

$${}^1\text{CH}_2 = {}^2\text{CH} - {}^3\text{CH}_2 - {}^4\text{CH}_2 - {}^5\text{C} \equiv {}^6\text{CH}$$
  
 , the  $\text{C}_2 - \text{C}_3$  bond formed by the overlapping of :  
 (A)  $sp - sp^2$  (B)  $sp^3 - sp^3$   
 (C)  $sp - sp^3$  (D)  $sp^2 - sp^3$
95. Which of the set of species have same hybridisation state but different shapes:  
 (A)  $\text{NO}_2^+$ ,  $\text{NO}_2$ ,  $\text{NO}_2^-$   
 (B)  $\text{ClO}_4^-$ ,  $\text{SF}_4$ ,  $\text{XeF}_4$   
 (C)  $\text{NH}_4^+$ ,  $\text{H}_3\text{O}^+$ ,  $\text{OF}_2$   
 (D)  $\text{SO}_4^{2-}$ ,  $\text{PO}_4^{3-}$ ,  $\text{ClO}_4^-$
96. Which of the following elements can not exhibit  $sp^3d$  hybridisation state:  
 (1) C (2) P  
 (3) Cl (4) B  
 Correct answer is:  
 (A) 1, 3 (B) 1, 4  
 (C) 2, 3 (D) 2, 4
97. In the protonation of  $\text{NH}_3$  molecule, following statement is true:  
 (A) A covalent bond is formed  
 (B) Hydrogen bond is formed  
 (C) Hybridisation state of N is changed  
 (D) Shape of  $\text{NH}_3$  molecule is changed
98. The shape of sulphate ion is:  
 (A) Hexagonal  
 (B) Square planar  
 (C) Trigonal bipyramidal  
 (D) Tetrahedral
99. In which following set of compound/ion has linear geometry:  
 (A)  $\text{CH}_4$ ,  $\text{NH}_4^+$ ,  $\text{BH}_4^-$   
 (B)  $\text{CO}_3^{2-}$ ,  $\text{NO}_3^-$ ,  $\text{BF}_3$   
 (C)  $\text{NO}_2^+$ ,  $\text{CO}_2$ ,  $\text{N}_3^-$   
 (D)  $\text{BeCl}_2$ ,  $\text{BCl}_3$ ,  $\text{CH}_4$
100. Which of the molecule is trigonal bipyramidal :  
 (A)  $\text{BF}_3$  (B)  $\text{CH}_4$  (C)  $\text{PCl}_5$  (D)  $\text{SF}_6$

101. The type of hybrid orbitals used by chlorine atom in  $\text{PCl}_5$  is/are:  
 (A)  $sp$ ,  $sp^2$ ,  $sp^3$  and  $sp^3d$   
 (B)  $sp$  and  $sp^3$   
 (C) Only  $sp^3$   
 (D) Only  $sp$
102. Which of the following having a square planar structure is:  
 (A)  $\text{NH}_4^+$  (B)  $\text{BF}_4^-$  (C)  $\text{XeF}_4$  (D)  $\text{CCl}_4$
103. The shape of  $\text{XeF}_4$  is :  
 (A) Tetrahedral (B) Linear (C) T-shape (D) Trigonal
104. Among the following compounds the one that is polar and has the central atom with  $sp^3$  hybridisation is:  
 (A)  $\text{H}_2\text{CO}_3$  (B)  $\text{SiF}_4$  (C)  $\text{BF}_3$  (D)  $\text{HClO}_2$
105. Which of the following will be octahedral:  
 (A)  $\text{SF}_6$  (B)  $\text{BF}_4^-$  (C)  $\text{PCl}_5$  (D)  $\text{XeF}_6$
106. The number of antibonding electron pairs in  $\text{O}_2^{2-}$  molecular ion on the basis of molecular orbital theory is (at no.  $O = 8$ ):  
 (A) 2 (B) 3 (C) 4 (D) 5
107. Which of the following option w.r.t. increasing bond order is correct?  
 (A)  $\text{NO} < \text{C}_2 < \text{O}_2^- < \text{He}_2^+$   
 (B)  $\text{C}_2 < \text{NO} < \text{He}_2^+ < \text{O}_2^-$   
 (C)  $\text{He}_2^+ < \text{O}_2^- < \text{NO} < \text{C}_2$   
 (D)  $\text{He}_2^+ < \text{O}_2^- < \text{C}_2 < \text{NO}$
108. The bond-orders of the given species are such that -  
 (A)  $\text{O}_2^- > \text{O}_2^+ > \text{O}_2^{2-} > \text{O}_2$   
 (B)  $\text{O}_2^+ > \text{O}_2 > \text{O}_2^- > \text{O}_2^{2-}$   
 (C)  $\text{O}_2^+ > \text{O}_2^{2-} > \text{O}_2^- > \text{O}_2$   
 (D)  $\text{O}_2^{2-} > \text{O}_2 > \text{O}_2^+ > \text{O}_2^-$
109. The ion that is isoelectronic with  $\text{CO}$  and having same bond order is:  
 (A)  $\text{CN}^-$  (B)  $\text{O}_2^+$  (C)  $\text{O}_2^-$  (D)  $\text{N}_2^+$
110. Which of the following is paramagnetic:  
 (A)  $\text{O}_2^-$  (B)  $\text{CN}^-$  (C)  $\text{CO}$  (D)  $\text{NO}^+$
111. In the following which of the two are paramagnetic:  
 (1)  $\text{N}_2$  (2)  $\text{CO}$  (3)  $\text{B}_2$  (4)  $\text{NO}_2$   
 correct answer is:  
 (A) (1) & (3) (B) (2) & (3)  
 (C) (3) & (4) (D) (2) & (4)
112. Increasing order of bond length in  $\text{NO}$ ,  $\text{NO}^+$  and  $\text{NO}^-$  is:  
 (A)  $\text{NO} > \text{NO}^- > \text{NO}^+$   
 (B)  $\text{NO}^+ < \text{NO} < \text{NO}^-$   
 (C)  $\text{NO} < \text{NO}^+ < \text{NO}^-$   
 (D)  $\text{NO} < \text{NO}^+ = \text{NO}^-$

113. In which of the following set, the value of bond order will be 2.5:
- (A)  $O_2^+$ , NO,  $NO^{+2}$ , CN  
 (B) CN,  $NO^{+2}$ ,  $CN^-$ ,  $F_2$   
 (C)  $O_2^+$ ,  $NO^{+2}$ ,  $O_2^{+2}$ ,  $CN^-$   
 (D)  $O_2^{-2}$ ,  $O_2^-$ ,  $O_2^+$ ,  $O_2$
114. The paramagnetic property of oxygen is well explained by:
- (A) Molecular orbital theory  
 (B) Resonance theory  
 (C) Valence bond theory  
 (D) VSEPR theory
115. Which of the following has fractional bond order?
- (A)  $O_2^{2+}$  (B)  $O_2^{2-}$  (C)  $F_2^{2-}$  (D)  $H_2^-$
116. Higher is the bond order, greater is:
- (A) Bond dissociation energy  
 (B) Covalent character  
 (C) Bond length  
 (D) Paramagnetism
117. The bond order of CO molecule on the basis of molecular orbital theory is:
- (A) Zero (B) 2 (C) 3 (D) 1
118. Which of the following group of molecules have bond order:
- (A)  $N_2^{-2}$ ,  $O_2^{-2}$ , CO  
 (B)  $N_2^+$ ,  $O_2^+$ , NO  
 (C)  $C_2^{-2}$ , BN,  $O_2$   
 (D)  $CN^-$ ,  $NO^+$ ,  $O_2^{+2}$
119. Which of the following species will have the minimum bond energy:
- (A)  $N_2$  (B)  $N_2^-$  (C)  $N_2^+$  (D)  $N_2^{-2}$
120. Which of the following ion has not bond order of 2.5 ?
- (A)  $O_2^-$  (B)  $O_2^+$  (C)  $N_2^+$  (D)  $N_2^-$
121.  $N_2$  and  $O_2$  are converted into monoanions, and respectively. Which of the following statements is wrong?
- (A) In , N-N bond weakens  
 (B) In , O-O bond order increases  
 (C) In , O-O bond order decreases  
 (D) becomes paramagnetic
122. Which of the following has maximum bond strength -
- (A)  $O_2$  (B)  $O_2^+$   
 (C)  $O_2^-$  (D)  $O_2^{2-}$
123.  $BeF_2$  has zero dipole moment whereas  $H_2O$  has dipole moment because:
- (A) Water is linear  
 (B)  $H_2O$  is bent  
 (C) F is more electronegative than O

- (D) Hydrogen bonding is present in  $\text{H}_2\text{O}$
124. Which of the following molecule have zero dipole moment:  
 (A)  $\text{BF}_3$  (B)  $\text{CH}_2\text{Cl}_2$  (C)  $\text{NF}_3$  (D)  $\text{SO}_2$
125. The dipole moment of  $\text{NH}_3$  is:  
 (A) Less than dipole moment of  $\text{NCl}_3$   
 (B) Higher than dipole moment of  $\text{NCl}_3$   
 (C) Equal to the dipole moment of  $\text{NCl}_3$   
 (D) None of these
126. Which of the following species are polar:  
 (1)  $\text{C}_6\text{H}_6$  (2)  $\text{XeF}_2$   
 (3)  $\text{SO}_2$  (4)  $\text{SF}_4$   
 (5)  $\text{SF}_6$   
 correct answer is:  
 (A) (2) & (4) (B) (1), (2) & (5)  
 (C) (1) & (5) (D) (3) & (4)
127. Which set of molecules is polar:  
 (A)  $\text{XeF}_4$ ,  $\text{IF}_7$ ,  $\text{SO}_3$   
 (B)  $\text{PCl}_5$ ,  $\text{C}_6\text{H}_6$ ,  $\text{SF}_6$   
 (C)  $\text{SnCl}_2$ ,  $\text{SO}_2$ ,  $\text{NO}_2$   
 (D)  $\text{CO}_2$ ,  $\text{CS}_2$ ,  $\text{C}_2\text{H}_6$
128. Dipole moment is shown by:  
 (A) 1, 4 - dichlorobenzene  
 (B) Cis 1, 2 - dichloro ethene  
 (C) Trans -1, 2 - dichloro ethene  
 (D) benzene
129. The correct order of dipole moment is:  
 (A)  $\text{CH}_4 < \text{NF}_3 < \text{NH}_3 < \text{H}_2\text{O}$   
 (B)  $\text{NF}_3 < \text{CH}_4 < \text{NH}_3 < \text{H}_2\text{O}$   
 (C)  $\text{NH}_3 < \text{NF}_3 < \text{CH}_4 < \text{H}_2\text{O}$   
 (D)  $\text{H}_2\text{O} < \text{NH}_3 < \text{NF}_3 < \text{CH}_4$
130. Which contains both polar and non-polar bonds?  
 (A)  $\text{NH}_4\text{Cl}$  (B)  $\text{HCN}$  (C)  $\text{H}_2\text{O}_2$  (D)  $\text{CH}_4$
131. What conclusion can be draw from the fact that  $\text{BF}_3$  has no dipole moment but  $\text{PF}_3$  is does:  
 (A)  $\text{BF}_3$  is not symmetrical but  $\text{PF}_3$  is symmetrical  
 (B)  $\text{BF}_3$  molecule must be linear  
 (C) Atomic radius of P is larger than that of B  
 (D)  $\text{BF}_3$  molecule must be planar triangular
132. Species having zero dipole moment:  
 (A)  $\text{XeF}_4$  (B)  $\text{SO}_2$  (C)  $\text{SF}_4$  (D)  $\text{CH}_2\text{Cl}_2$
133. The correct order of decreasing polarity is:  
 (A)  $\text{HF} > \text{SO}_2 > \text{H}_2\text{O} > \text{NH}_3$



- (B)  $\text{HF} > \text{H}_2\text{O} > \text{SO}_2 > \text{NH}_3$   
 (C)  $\text{HF} > \text{NH}_3 > \text{SO}_2 > \text{H}_2\text{O}$   
 (D)  $\text{H}_2\text{O} > \text{NH}_3 > \text{SO}_2 > \text{HF}$

134. The dipole moment of given molecules are such that -

- (A)  $\text{BF}_3 > \text{NF}_3 > \text{NH}_3$   
 (B)  $\text{NF}_3 > \text{BF}_3 > \text{NH}_3$   
 (C)  $\text{NH}_3 > \text{NF}_3 > \text{BF}_3$   
 (D)  $\text{NH}_3 > \text{BF}_3 > \text{NF}_3$

### LEVEL-2

- The correct order of dipole moment is  
 (a)  $\text{CH}_4 < \text{NF}_3 < \text{NH}_3 < \text{H}_2\text{O}$   
 (b)  $\text{NF}_3 < \text{CH}_4 < \text{NH}_3 < \text{H}_2\text{O}$   
 (c)  $\text{NH}_3 < \text{NF}_3 < \text{CH}_4 < \text{H}_2\text{O}$   
 (d)  $\text{H}_2\text{O} < \text{NH}_3 < \text{NF}_3 < \text{CH}_4$
- Carbon atoms in  $\text{C}_2(\text{CN})_4$  are  
 (a)  $\text{sp}$ -hybridised  
 (b)  $\text{sp}^2$ -hybridised  
 (c)  $\text{sp}$ - and  $\text{sp}^2$ -hybridised  
 (d)  $\text{sp}$ ,  $\text{sp}^2$  and  $\text{sp}^3$ -hybridised
- Which of the following is paramagnetic?  
 (a)  $\text{O}_2^-$   
 (b)  $\text{CN}^-$   
 (c)  $\text{CO}$   
 (d)  $\text{NO}^+$
- Which of the molecule is T-shaped?  
 (a)  $\text{BeF}_2$   
 (b)  $\text{BCl}_3$   
 (c)  $\text{NH}_3$   
 (d)  $\text{ClF}_3$
- Which shows a change in the type of hybridization when  
 (a)  $\text{NH}_3$  combines with  $\text{H}^+$   
 (b)  $\text{AlH}_3$  combines with  $\text{H}^-$   
 (c) In both cases  
 (d) In none cases
- Molecular shapes of  $\text{SF}_4$ ,  $\text{CF}_4$  and  $\text{XeF}_4$  are  
 (a) the same, with 2, 0 and 1 lone pair of electrons on central atom, respectively.  
 (b) the same, with 1, 1 and 1 lone pair of electrons on central atom, respectively.  
 (c) different, with 0, 1 and 2 lone pair of electrons on central atom, respectively.  
 (d) different with 1, 0 and 2 lone pair of electrons on central atom, respectively.
- The hybridization of atomic orbitals of nitrogen in  $\text{NO}_2^+$ ,  $\text{NO}_3^-$  and  $\text{NH}_4^+$  are  
 (a)  $\text{sp}$ ,  $\text{sp}^3$  and  $\text{sp}^2$  respectively  
 (b)  $\text{sp}$ ,  $\text{sp}^2$  and  $\text{sp}^3$  respectively  
 (c)  $\text{sp}^2$ ,  $\text{sp}$  and  $\text{sp}^3$  respectively  
 (d)  $\text{sp}^2$ ,  $\text{sp}^3$  and  $\text{sp}$  respectively
- Which of the following contains both polar and non-polar bonds?  
 (a)  $\text{NH}_4\text{Cl}$   
 (b)  $\text{HCN}$   
 (c)  $\text{H}_2\text{O}_2$   
 (d)  $\text{CH}_4$
- The bond order of the superoxide ( $\text{O}_2^-$ ) is  
 (a) 1  
 (b) 1.5

- (c) 2 (d) 2.5
10. Amongst  $\text{LiCl}$ ,  $\text{BeCl}_2$ ,  $\text{MgCl}_2$  and  $\text{RbCl}$  the compounds with greatest and least ionic character, respectively are  
 (a)  $\text{LiCl}$  and  $\text{RbCl}$  (b)  $\text{RbCl}$  and  $\text{BeCl}_2$   
 (c)  $\text{RbCl}$  and  $\text{MgCl}_2$  (d)  $\text{MgCl}_2$  and  $\text{BeCl}_2$
11. In which molecule sulphur atom is not  $sp^3$ -hybridised  
 (a)  $\text{SO}_4^{2-}$  (b)  $\text{SF}_4$   
 (c)  $\text{SF}_2$  (d)  $\text{S}_8$
12. Which of the following orders regarding the bond order is correct?  
 (a)  $\text{O}_2^- > \text{O}_2 > \text{O}_2^+$  (b)  $\text{O}_2^- < \text{O}_2 < \text{O}_2^+$   
 (c)  $\text{O}_2^- > \text{O}_2 < \text{O}_2^+$  (d)  $\text{O}_2^- < \text{O}_2 > \text{O}_2^+$
13. The shape of  $\text{XeF}_4$  is  
 (a) tetrahedral (b) square planar  
 (c) pyramidal (d) nearly linear
14. The shape of  $\text{XeOF}_4$  is  
 (a) square pyramidal (b) square antiprismatic  
 (c) distorted octahedral (d) pentagonal bipyramidal
15. The geometry of  $\text{XeO}_2\text{F}_2$  is  
 (a) plane triangular (b) see-saw  
 (c) square planar (d) tetrahedral
16. The Xe atom in  $\text{XeOF}_4$  involves the hybridization  
 (a)  $sp^3$  (b)  $sp^3d$   
 (c)  $sp^3d^2$  (d)  $sp^3d^3$
17. The hybridisation of P in  $\text{PO}_4^{3-}$  is same as that of  
 (a) I in  $\text{ICl}_4^-$  (b) S in  $\text{SO}_3$   
 (c) N in  $\text{NO}_3^-$  (d) S in  $\text{SO}_4^{2-}$
18. Which of the following statement is correct?  
 (a) Polarisation of an anion is maximum by high charged cation.  
 (b) Small sized cation minimizes the polarisation.  
 (c) A small anion brings about a large degree of polarisation.  
 (d) A large anion undergoes a small degree of polarisation.
19. Which of the following have identical bond order :  
 (i)  $\text{CN}^-$  (ii)  $\text{O}_2^-$  (iii)  $\text{NO}^+$  (iv)  $\text{CN}^+$   
 (a) (i) and (iii) (b) (ii) and (iv)  
 (c) (i) and (iv) (d) (ii) and (iii)
20. Which of the following has highest ionic character?  
 (a)  $\text{MgCl}_2$  (b)  $\text{CaCl}_2$   
 (c)  $\text{BaCl}_2$  (d)  $\text{BeCl}_2$

**ONE OR MORE THAN ONE CHOICE CORRECT**

- PCl<sub>3</sub> and PCl<sub>5</sub> both exist but only PH<sub>3</sub> exists while PH<sub>5</sub> does not exist. This is because
  - H is nearly as electronegative as P.
  - the activation energy for the formation of PH<sub>5</sub> is very high.
  - PH<sub>5</sub> immediately decomposes to PH<sub>3</sub> and H<sub>2</sub> because its equilibrium constant for the decomposition is very high.
  - An element is able to utilize its d-orbitals for bonding only with elements which are more electronegative than it.
- Which of the following molecules or ions is not linear?
  - BeCl<sub>2</sub>
  - ICl<sub>2</sub><sup>-</sup>
  - CS<sub>2</sub>
  - ICl<sub>2</sub><sup>+</sup>
- Which of the following is/are non-polar but contain(s) polar bonds?
  - HCl
  - H<sub>2</sub>O
  - SO<sub>3</sub>
  - CO<sub>2</sub>
- The 90° angles between bond pair-bond pair of electrons exists in
  - ClF<sub>3</sub>
  - I<sub>3</sub><sup>-</sup>
  - BrF<sub>5</sub>
  - PCl<sub>4</sub><sup>+</sup>
- Which of the following species is paramagnetic?
  - CN<sup>-</sup>
  - NO
  - O<sub>2</sub><sup>2-</sup>
  - O<sub>2</sub>
- Shape of I<sub>3</sub><sup>-</sup> is
  - tetrahedral
  - triangular bipyramidal
  - distorted trigonal bipyramidal
  - linear
- Which among the following has bond order zero?
  - F<sub>2</sub><sup>2-</sup>
  - Ar<sub>2</sub>
  - He<sub>2</sub><sup>+1</sup>
  - H<sub>2</sub><sup>+1</sup>
- KF combines with HF to form KHF<sub>2</sub>. The compound contains the species
  - K<sup>+</sup>, F<sup>-</sup> and H<sup>+</sup>
  - K<sup>+</sup>, F<sup>-</sup> and HF
  - K<sup>+</sup> and [HF<sub>2</sub>]<sup>-</sup>
  - one cation and one anion
- Among the following species, identify the isostructural pairs  
 NF<sub>3</sub>, NO<sub>3</sub><sup>-</sup>, BF<sub>3</sub>, H<sub>3</sub>O<sup>+</sup>, N<sub>3</sub><sup>-</sup>, I<sub>3</sub><sup>-</sup>
  - [NF<sub>3</sub>, NO<sub>3</sub><sup>-</sup>] and [BF<sub>3</sub>, H<sub>3</sub>O<sup>+</sup>]
  - [I<sub>3</sub><sup>-</sup>, N<sub>3</sub><sup>-</sup>] and [NO<sub>3</sub><sup>-</sup>, BF<sub>3</sub>]
  - [NF<sub>3</sub>, H<sub>3</sub>O<sup>+</sup>] and [NO<sub>3</sub><sup>-</sup>, BF<sub>3</sub>]
  - [NF<sub>3</sub>, H<sub>3</sub>O<sup>+</sup>] and [N<sub>3</sub><sup>-</sup>, I<sub>3</sub><sup>-</sup>]

10.  $\text{NH}_3$  and  $\text{BF}_3$  form adduct readily
- (a) Hybridization of  $\text{NH}_3$  remains same.
  - (b) through co-ordinate bond between B and N
  - (c) Hybridization of  $\text{NH}_3$  changes from  $\text{sp}^3$  to  $\text{sp}^2$ .
  - (d) Hybridization of B changes from  $\text{sp}^2$  to  $\text{sp}^3$ .



### MATCH THE FOLLOWING

1.

Column I (Compounds)	Column II (Shape / Property)
I. $\text{CH}_4$	(A) tetrahedral
II. $\text{NH}_3$	(B) hydrogen bonding
III. $\text{HF}$	(C) see-saw
IV. $\text{SF}_4$	(D) linear
	(E) polar molecule

### REASONING TYPE

**Directions:** Read the following questions and choose

- (A) If both the statements are true and statement-2 is the correct explanation of statement-1.  
 (B) If both the statements are true but statement-2 is not the correct explanation of statement-1.  
 (C) If statement-1 is True and statement-2 is False.  
 (D) If statement-1 is False and statement-2 is True.

- Statement-1:** Xe atom in  $\text{XeF}_2$  assumes sp hybrid state.  
**Statement-2:**  $\text{XeF}_2$  molecule does not follow octet rule.  
 (a) (A) (b) (B) (c) (C) (d) (D)
- Statement-1:**  $\text{NO}_2$  is paramagnetic at room temperature and diamagnetic at lower temperature.  
**Statement-2:** When electrons are odd, the molecule is paramagnetic or else diamagnetic.  
 (a) (A) (b) (B) (c) (C) (d) (D)
- Statement-1:**  $\text{SO}_2$  is bent.  
**Statement-2:**  $\text{SO}_2$  has two  $\pi$  bonds.  
 (a) (A) (b) (B) (c) (C) (d) (D)
- Statement-1:** Bond angle of  $\text{NH}_3$  is lower than in  $\text{CH}_4$ .  
**Statement-2:**  $\text{NH}_3$  has one lone pair and lone pair-bond pair repulsion is greater than bond pair-bond pair repulsion.  
 (a) (A) (b) (B) (c) (C) (d) (D)

5. **Statement-1:**  $\text{H}_2\text{O}$  is liquid while  $\text{H}_2\text{S}$  is gas.

**Statement-2:**  $\text{H}_2\text{O}$  has hydrogen bonding while  $\text{H}_2\text{S}$  does not.

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

**LINKED COMPREHENSION TYPE**

Ionic bond is defined as the electrostatic force of attraction holding the oppositely charged ions. Ionic compounds are mostly crystalline solids having high melting and boiling points, electrical conductivity in molten state, solubility in water etc. Covalent bond is defined as the force which binds atoms of same or different elements by mutual sharing of electrons in a covalent bond. Covalent compounds are solids, liquids or gases. They are low melting and boiling point compounds. They are more soluble in non-polar solvents.

1. The valence electrons not involved in formation of covalent bonds are called  
(a) non-bonding electrons                      (b) lone pairs  
(c) unshared pairs                              (d) all of these
2. The amount of energy released when one mole of ionic solid is formed by close packing of gaseous ion is called:  
(a) Ionisation energy                              (b) Solvation energy  
(c) Lattice energy                                  (d) Hydration energy
3. Which of the following gives a white precipitate with  $\text{AgNO}_3$ ?  
(a)  $\text{NaCl}$     (b)  $\text{CCl}_4$   
(c)  $\text{CO}_2$     (d)  $\text{CHCl}_3$

**EXERCISE – III (CROSSING THE HURDLES)****SUBJECTIVE PROBLEMS**

1.  $\text{MgCl}_2$  is linear but  $\text{SnCl}_2$  is angular. Explain.
2. Although  $\text{I}_3^-$  is known,  $\text{F}_3^-$  is not. Why?
3. Why are inert gases less reactive?
4. Predict the shapes of the following molecules using the VSEPR model :  $\text{BeCl}_2$ ,  $\text{SiCl}_4$ ,  $\text{AsF}_5$ ,  $\text{H}_2\text{S}$ ,  $\text{PH}_3$ .
5. Dichlorobenzene exists in three different isomers called ortho, meta and para. Which of these would have a non zero dipole moment and why?
6. Account for the fact that carbon-carbon bond lengths in ethane, ethene and ethyne are 154 pm, 134 pm and 120 pm, respectively.
7. On what factors does the polarity of a molecule depend?
8. Why is bond energy of P-Cl bond different in  $\text{PCl}_3$  and  $\text{PCl}_5$ ?
9. The hybridisation of central atom in  $\text{CH}_4$ ,  $\text{H}_2\text{O}$  and  $\text{NH}_3$  is  $\text{sp}^3$ . Why are the bond angles different in these three cases?
10.  $\text{BF}_3$  and graphite, both are coplanar having  $\text{sp}^2$  hybridization but later is a conductor. Explain.
11.  $\text{XeF}_2$  is linear inspite of the fact that Xe involves  $\text{sp}^3\text{d}$  hybridisation. Explain.
12. Give reason why  $\text{ClF}_3$  exists but  $\text{FCl}_3$  does not?
13.  $\text{NaCl}_a$  gives a white precipitate with  $\text{AgNO}_3$  solution but  $\text{CCl}_4$  or  $\text{CHCl}_3$  does not. Explain.
14. Explain why  $\text{N}_2$  has greater dissociation energy than  $\text{N}_2^+$ .
15.  $\text{BaSO}_4$  being an electrovalent compound does not pass into solution state in water. Explain.



## ANSWERS

### EXERCISE – I

#### COVANLECY

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.
Ans.	a	c	c	d	b	d	d	3	4

#### QUESTION – 10

Ques.	(1)	(2)	(3)	(4)	(5)	(6)	(7)
Ans.	1	1,3,5,7	2,4,6,8	3	3,5	1,3,5,7	2,4,6
Ques.	(8)	(9)	(10)	(11)	(12)	(13)	(14)
Ans.	1	2	3	2	4	1	4
Ques.	(15)						
Ans.	1,3,5,7						

#### OVERLAPPING OF ATOMIC ORBITALS

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	b	c	b	a, b, d	a, c	b	c, d	c	b
Ques.	11.	12.					13.	14.	15.	
Ans.	c	(a) – RT; (b) – PS; (c) – PS; (d) – QT					Zero	3	2	

#### HYBRIDISATION – I

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.
Ans.	a	c	d	d	a, b, c	a, c	5	5	6

#### QUESTION – 10

Ques.	(1)	(2)	(3)	(4)	(5)	(6)	(7)
-------	-----	-----	-----	-----	-----	-----	-----

Ans.	$sp^3$	$sp^3$	$sp$	$sp^3$	$sp^3$	$sp^2$	$sp^3d$
Ques.	(8)	(9)	(10)	(11)	(12)	(13)	(14)
Ans.	$sp^2$	$sp^2$	$sp$	$sp^2$	$sp^2$	$sp^3$	$sp^3$
Ques.	(15)	(16)	(17)	(18)	(19)	(20)	(21)
Ans.	$sp^3d$	$sp^3d^2$	$sp^3d$	$sp^3d$	$sp^3d^2$	$sp^3$	$sp^3d$
Ques.	(22)	(23)	(24)	(25)	(26)	(27)	(28)
Ans.	$sp^3d$	$sp^3d$	$sp^3d^2$	$sp^3$	$sp^3$	$sp^3d^2$	$sp$
Ques.	(29)	(30)	(31)	(32)	(33)	(34)	(35)
Ans.	$sp^3d^2$	$sp^3$	$sp^3$	$sp^3$	$sp^3$	$sp^3, sp^2$	$sp^3, sp^2$
Ques.	(36)	(37)	(38)	(39)	(40)		
Ans.	$sp^2$	$sp, sp^3$	$sp^3$	$sp^3d^2$	$sp^3$		

## HYBRIDISATION - II

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	a	b	b	b	c	b	a – QST, b – QRST, c – PQRST, d – ST	4	2	3

## VSEPR THEORY - I

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.
Ans.	a	c	c	b	b	a, b, c, d	b, d	6	1

## QUESTION – 10

Ques.	(1)	(2)	(3)	(4)	(5)	(6)	(7)
Ans.	Bent	Bent	Trigonal planar	Linear	Tetrahedral	Tetrahedral	Trigonal pyramidal
Ques.	(8)	(9)	(10)	(11)	(12)	(13)	(14)
Ans.	Trigonal pyramidal	Octahedral	Linear	Linear	Bent – T	See – saw	Trigonal bipyramidal

Ques.	(15)	(16)	(17)	(18)	(19)	(20)	(21)
Ans.	Tetrahedral	Square planar	Trigonal pyramidal	Bent – T	Bent	Bent	Linear
Ques.	(22)	(23)	(24)	(25)	(26)	(27)	(28)
Ans.	Bent	Square planar	See – saw	See – saw	Trigonal planar	Pentagonal planar	Square Pyramidal
Ques.	(29)	(30)	(31)	(32)	(33)	(34)	(35)
Ans.	Linear	Bent	Tetrahedral	Linear	Tetrahedral	Trigonal pyramidal	Bent
Ques.	(36)	(37)	(38)	(39)	(40)		
Ans.	Pentagonal bipyramidal	Bent	Linear	Bent	Linear		

### VSEPR THEORY - II

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	a	d	a, c	a, c, d	d	d	c	c	A – PQST; B – PS; C – PRT; D - PRT

### RESONANCE AND EQUIVALENT HYBRID ORBITALS

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.
Ans.	c	a	b	d	b, c, d	a, b, d	4	1	8

### QUESTION – 10

Ques.	(1)	(2)	(3)	(4)	(5)	(6)	(7)
Ans.	1.33	2	1.33	1.5	1.25	1.75	1.5
Ques.	(8)	(9)	(10)				
Ans.	1.67	1.5	1.5				

### BENT'S RULE

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	c	b	c	a, c, d	a, b, c, d	a, b, d	a, b, c, d	c	d	a

### BOND PARAMETERS

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	c	d	b	c	a	d	b	a	c
Ques.	11.	12.	13.	14.	15.	16.		17.		
Ans.	b	b	a, c	a, b	b, c	(a) – p, (b) – s, (c) – r, (d) – q		(a) – p, q, r, t (b) – s, t, (c) – p, s, t, (d) r, t		
Ques.	18.	19.	20.							
Ans.	0	6	5							

### DRAGO'S GENERALISATION

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	b	b	b	c	a	c	d	a	4	Zero

### BACK BONDING & BRIDGE BONDING

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	b	b	b	c, d	a, b, c, d	a, c, d	a, b, c, d	d	c	(a) – QRS; (b) – QRT; (c) – PRT; (d) – QRS
Ques.	11.			12.	13.					
Ans.	(a) P, (b) Q, S, (c) Q, S, T, (d) R, T,			2	4					

### QUESTION – 14 (P – Planar, NP – Non planar)

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	P	P	NP	NP	NP	P	NP	P	NP	NP
Ques.	11.	12.	13.	14.	15.	16.	17.	18.	19.	20.
Ans.	P	P	NP	P	NP	NP	P	P	P	P

### QUESTION – 15

15.
(i), (ii), (iii), (vi), (viii), (x)

### ODD ELECTRON SPECIES AND NON - EXISTING SPECIES

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.
Ans.	c	b	c	b	a, b, c, d	a, c, d	a, b, d	c	1
Ques.	10.								
Ans.	(i), (ii), (iii), (iv), (vi), (vii), (viii), (ix), (xi), (xiii)								

### DIPOLE MOMENT

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	b	a	a	b	a	a	a, b	a, c	a, b, c
Ques.	11.	12.	13.	14.						
Ans.	c, d	9	9	3						

### QUESTION – 15

Ques.	(i)	(ii)	(iii)	(iv)	(v)	(vi)	(vii)
Ans.	NP	P	NP	NP	P	NP	NP
Ques.	(viii)	(ix)	(x)	(xi)	(xii)	(xiii)	(xiv)
Ans.	NP	P	NP	NP	P	P	NP
Ques.	(xv)	(xvi)	(xvii)	(xviii)	(xix)	(xx)	(xxi)
Ans.	NP	NP	NP	P	NP	P	NP
Ques.	(xxii)	(xxiii)	(xxiv)	(xxv)	(xxvi)	(xxvii)	(xxviii)
Ans.	NP	P	P	NP	P	P	NP

Ques.	(xxix)	(xxx)
Ans.	NP	P

### MOLECULAR ORBITAL THEORY – I

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	a	b	d	a	b	c	a	a	a, c	a, c
Ques.	11.	12.	13.	14.						
Ans.	b, c, d	c	d	(a) – p, q, r, s, t; (b) – p, r, t; (c) – q, s, t; (d) – p, r, t						

### QUESTION – 15

Ques.	(1)	(2)	(3)	(4)	(5)	(6)	(7)
Ans.	0.5, p	0.5, p	3, D	1.5, p	1, D	2.5, p	2, D
Ques.	(8)	(9)	(10)	(11)	(12)	(13)	(14)
Ans.	1, p	2.5, p	1, D	3, D	2.5, p	0.5, p	0.5, p
Ques.	(15)	(16)	(17)	(18)	(19)	(20)	(21)
Ans.	3, D	1.5, p	0.5, p	2, p	1, D	0.5, p	0.5, p
Ques.	(22)	(23)	(24)	(25)	(26)	(27)	(28)
Ans.	0.5, p	2, p	3, D	2.5, p	3, D	3, D	2.5, p
Ques.	(29)	(30)					
Ans.	2, p	2, D					

### MOLECULAR ORBITAL THEORY – II

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	a	d	d	d	c	b	a, c, d	b, c	a, b, c, d	a, c
Ques.	11.	12.						13.	14.	15.
Ans.	c	A – P, Q, R, S; B – P, Q, R; C – Q, T; D – S						3	1	5

### INTERODUCTION OF IONIC BOND

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	c	b	a	a	d	d	b	c	9	1

### LATTICE ENTHALPY AND FAJAN'S RULE

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	c	b	b	c	c	d	a	b	b
Ques.	11.	12.	13.	14.	15.					
Ans.	a	a, b, d	a, c	4	2					

### INERT PAIR EFFECT

Ques.	1.	2.	3.	4.	5.	6.
Ans.	a	a	c	d	c	b
Ques.	7.	8.			9.	10.
Ans.	a, b, c, d	(A) – Q; (B) – PRS; (C) – P; (D) – Q			1	1

### THERMAL STABILITY & SOLUBILITY

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	d	a	a	a	a	a, b, c, d	a, b, c, d	0	4

### WEAK FORCES

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	d	c	c	b	d	c	b	c	6	1



## HYDROGEN BONDING

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.
Ans.	d	b	b	a	a, b, c	b	d	6	6
Ques.	10.	(i), (iv), (v), (vi), (viii) → form intermolecular H-bond							
Ans.		(ii), (iii), (ix) → form intermolecular H-bond (vii), (x) → Neither form intermolecular H-bond with their other molecule nor form intermolecular H-bond							

## SECTION -S

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	c	b	b	d	a	c	c	b	a	d
Ques.	11.	12.	13.	14.	15.	16.	17.	18.	19.	20.
Ans.	c, d	c, d	a, b	c, d	a, b	a	b	d	b	a
Ques.	21.	22.	23.				24.			
Ans.	b	c	(A) – Q; (B) – P; (C) – S; (D) – R				(A) – T; (B) – P; (C) – Q; (D) – R			
Ques.	25.									
Ans.	(A) – S; (B) – T; (C) – Q; (D) – P									

**EXERCISE – II**
**SINGLE CHOICE CORRECT**

<b>Ques.</b>	<b>1.</b>	<b>2.</b>	<b>3.</b>	<b>4.</b>	<b>5.</b>	<b>6.</b>	<b>7.</b>	<b>8.</b>	<b>9.</b>	<b>10.</b>
<b>Ans.</b>	a	b	c	c	a	b	c	a	a	a
<b>Ques.</b>	<b>11.</b>	<b>12.</b>	<b>13.</b>	<b>14.</b>	<b>15.</b>	<b>16.</b>	<b>17.</b>	<b>18.</b>	<b>19.</b>	<b>20.</b>
<b>Ans.</b>	c	d	d	b	b	a	b	c	c	d
<b>Ques.</b>	<b>21.</b>	<b>22.</b>	<b>23.</b>	<b>24.</b>	<b>25.</b>	<b>26.</b>	<b>27.</b>	<b>28.</b>	<b>29.</b>	<b>30.</b>
<b>Ans.</b>	a	c	c	c	d	c	c	a	c	d
<b>Ques.</b>	<b>31.</b>	<b>32.</b>	<b>33.</b>	<b>34.</b>	<b>35.</b>	<b>36.</b>	<b>37.</b>	<b>38.</b>	<b>39.</b>	<b>40.</b>
<b>Ans.</b>	b	c	b	d	c	b	d	a	b	b
<b>Ques.</b>	<b>41.</b>	<b>42.</b>	<b>43.</b>	<b>44.</b>	<b>45.</b>	<b>46.</b>	<b>47.</b>	<b>48.</b>	<b>49.</b>	<b>50.</b>
<b>Ans.</b>	c	d	a	b	b	c	d	b	b	d
<b>Ques.</b>	<b>51.</b>	<b>52.</b>	<b>53.</b>	<b>54.</b>	<b>55.</b>	<b>56.</b>	<b>57.</b>	<b>58.</b>	<b>59.</b>	<b>60.</b>
<b>Ans.</b>	b	c	c	c	b	d	b	a	b	d
<b>Ques.</b>	<b>61.</b>	<b>62.</b>	<b>63.</b>	<b>64.</b>	<b>65.</b>	<b>66.</b>	<b>67.</b>	<b>68.</b>	<b>69.</b>	<b>70.</b>
<b>Ans.</b>	c	a	a	c	c	c	c	c	c	a
<b>Ques.</b>	<b>71.</b>	<b>72.</b>	<b>73.</b>	<b>74.</b>	<b>75.</b>	<b>76.</b>	<b>77.</b>	<b>78.</b>	<b>79.</b>	<b>80.</b>
<b>Ans.</b>	b	a	c	d	c	b	c	c	c	a
<b>Ques.</b>	<b>81.</b>	<b>82.</b>	<b>83.</b>	<b>84.</b>	<b>85.</b>	<b>86.</b>	<b>87.</b>	<b>88.</b>	<b>89.</b>	<b>90.</b>
<b>Ans.</b>	b	d	a	a	d	b	b	b	a	c
<b>Ques.</b>	<b>91.</b>	<b>92.</b>	<b>93.</b>	<b>94.</b>	<b>95.</b>	<b>96.</b>	<b>97.</b>	<b>98.</b>	<b>99.</b>	<b>100.</b>
<b>Ans.</b>	b	b	c	d	c	b	d	d	c	c
<b>Ques.</b>	<b>101.</b>	<b>102.</b>	<b>103.</b>	<b>104.</b>	<b>105.</b>	<b>106.</b>	<b>107.</b>	<b>108.</b>	<b>109.</b>	<b>110.</b>
<b>Ans.</b>	c	c	b	d	a	c	d	b	a	a
<b>Ques.</b>	<b>111.</b>	<b>112.</b>	<b>113.</b>	<b>114.</b>	<b>115.</b>	<b>116.</b>	<b>117.</b>	<b>118.</b>	<b>119.</b>	<b>120.</b>
<b>Ans.</b>	c	b	a	a	d	a	c	b	d	a
<b>Ques.</b>	<b>121.</b>	<b>122.</b>	<b>123.</b>	<b>124.</b>	<b>125.</b>	<b>126.</b>	<b>127.</b>	<b>128.</b>	<b>129.</b>	<b>130.</b>
<b>Ans.</b>	b	b	b	a	b	d	c	b	a	c
<b>Ques.</b>	<b>131.</b>	<b>132.</b>	<b>133.</b>	<b>134.</b>						
<b>Ans.</b>	d	a	b	c						

### SINGLE CHOICE CORRECT

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	a	c	a	d	b	d	b	c	b	b
Ques.	11.	12.	13.	14.	15.	16.	17.	18.	19.	20.
Ans.	b	b	b	a	b	c	d	a	a	c

### ONE OR MORE THAN ONE CHOICE CORRECT

Ques.	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.
Ans.	a, d	d	c, d	a, c	b, d	c, d	a, b	c, d	b, c, d	a, b, d

### MATCH THE FOLLOWING

Ques.	1.
Ans.	I – (A) ; II – (B), (E) ; III – (B), (D), (E) ; IV – (C), (E)

### REASONING TYPE

Ques.	1.	2.	3.	4.	5.
Ans.	d	c	b	a	a

### LINKED COMPREHENSION TYPE

Ques.	1.	2.	3.
Ans.	d	c	a

**EXERCISE – III****SUBJECTIVE PROBLEMS**

1.  $\text{MgCl}_2$  is  $sp$  hybridized, whereas  $\text{SnCl}_2$  is  $sp^2$  hybridized.
2. Iodine can expand its octet due to presence of vacant d-orbitals while F cannot.
3. Inert gases have fully filled  $ns^2np^6$  configuration of valence shell.
4.  $\text{BeCl}_2$  : linear ;  $\text{SiCl}_4$  : tetrahedral ;  $\text{AsF}_5$  : trigonal bipyramidal ;  $\text{H}_2\text{S}$  : angular ;  $\text{PH}_3$  : pyramidal
5. Ortho and meta dichlorobenzene have got a non zero dipole moment because of unsymmetrical structure.
6. Ethane, ethene and ethyne contain single, double and triple bond respectively and bond length is inversely related to bond order.
7. The molecule contains polar bonds and must be non symmetrical in nature.
8. Phosphorus atom is  $sp^3$  hybridised in  $\text{PCl}_3$  and  $sp^3d$  hybridised in  $\text{PCl}_5$ .
9. Bond angles of  $\text{CH}_4$ ,  $\text{H}_2\text{O}$  and  $\text{NH}_3$  are different due to presence of one and two lone pairs on N in  $\text{NH}_3$  and O in  $\text{H}_2\text{O}$  respectively while  $\text{CH}_4$  has no lone pair.
10. Graphite (C) has one unhybridized p-orbital containing single electron. Due to this free electron, graphite is used as a conductor.
11. In  $\text{XeF}_2$  molecules, out of five  $sp^3d$  hybridised orbitals, the three equatorial orbitals have one lone pair each while the two axial orbitals have one bond pair each. This results in linear shape of  $\text{XeF}_2$ .
12. Chlorine can expand its octet due to the presence of vacant d-orbitals.
13.  $\text{NaCl}$  is an ionic compound while  $\text{CCl}_4$  and  $\text{CHCl}_3$  are covalent.  $\text{NaCl}$  gives  $\text{Na}^+$  and  $\text{Cl}^-$  in aqueous solution.
14. In  $\text{N}_2$  molecule, bond order is three. In  $\text{N}_2^+$  one electron is removed from bonding molecular orbital ( $\sigma 2p_x$ ). This reduces the bond order to 2.5. Since bond order is directly proportional to bond strength,  $\text{N}_2$  has higher dissociation energy than  $\text{N}_2^+$ .
15. The lattice energy of  $\text{BaSO}_4$  is more than the hydration energy.

**PROFICIENCY TEST– I**

The following 10 questions deal with the basic concepts of this section. Answer the following briefly. Go to the next section only if your score is greater than or equal to 8. Do not consult the study material while attempting the questions.

1. True/False. Every element tries to acquire the electronic configuration of the nearest noble gas.
2. True/False. Equal sharing of electrons lead to ionic bond while transfer of electrons lead to covalent bond.
3. True/False.  $\text{Na}_2\text{SO}_4$  contains only ionic and covalent bonds.
4. True/False. The dipole moment of  $\text{NH}_3$  is higher than that of  $\text{NF}_3$ .
5. True/False. In cyanide ion, the negative charge resides on nitrogen.
6. The percent ionic character using dipole moment is calculated by \_\_\_\_\_.
7.  $\text{AgI}$  is \_\_\_\_\_ covalent than  $\text{KI}$ .
8.  $\text{AlCl}_3$  has contraction/expansion of octet.
9. The dipole moment of  $\text{BCl}_3$  is \_\_\_\_\_ while that of  $\text{PCl}_3$  is \_\_\_\_\_.
10. Formation of ionic bond is favoured by \_\_\_\_\_.

Pinnacle

**ANSWERS TO PROFICIENCY TEST– I**

1. True
2. False
3. False
4. True
5. False
6.  $\left( \frac{\text{Observed dipole moment}}{\text{Calculated dipole moment}} \times 100 \right)$
7. More
8. Contraction
9. Zero, non-zero
10. Low ionization potential of metal, high electron affinity of non-metal, low dissociation enthalpy, low sublimation energy and high lattice energy of the crystal.



Pinnacle

**PROFICIENCY TEST– II**

The following 10 questions deal with the basic concepts of this section. Answer the following briefly. Go to the next section only if your score is greater than or equal to 8. Do not consult the study material while attempting the questions.

1. True/False. In  $\text{PCl}_5$ , all bonds are equivalent and have same bond length.
2. True/False. Sigma bonds are stronger than pi-bonds.
3. True/False. The hybridization of the central atom in  $\text{SO}_2$  is  $\text{sp}^2$ .
4. True/False.  $\text{H}_2\text{O}$  forms the strongest intermolecular hydrogen bond.
5. True/False. The bond strength of O–O bond is more in  $\text{O}_2^+$  than in  $\text{O}_2^-$ .
6. The shape of  $\text{SF}_4$  is \_\_\_\_\_.
7. A molecule of acetylene has \_\_\_\_\_ sigma and \_\_\_\_\_ pi-bonds.
8. The hybridization of Cl in  $\text{ClO}^-$  and  $\text{ClO}_4^-$  is \_\_\_\_\_ and \_\_\_\_\_ respectively.
9. o-hydroxy benzaldehyde is \_\_\_\_\_ volatile than p-hydroxy benzaldehyde.
10. When NO is converted to  $\text{NO}^+$ , bond length of N–O bond \_\_\_\_\_ while when CN is converted to  $\text{CN}^+$ , bond length of C–N bond \_\_\_\_\_.

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**ANSWERS TO PROFICIENCY TEST– II**

1. False
2. True
3. True
4. False
5. True
6. See-saw shaped or irregular tetrahedron.
7. 3, 2
8.  $sp^3$  in both
9. More
10. Decreases, increases



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