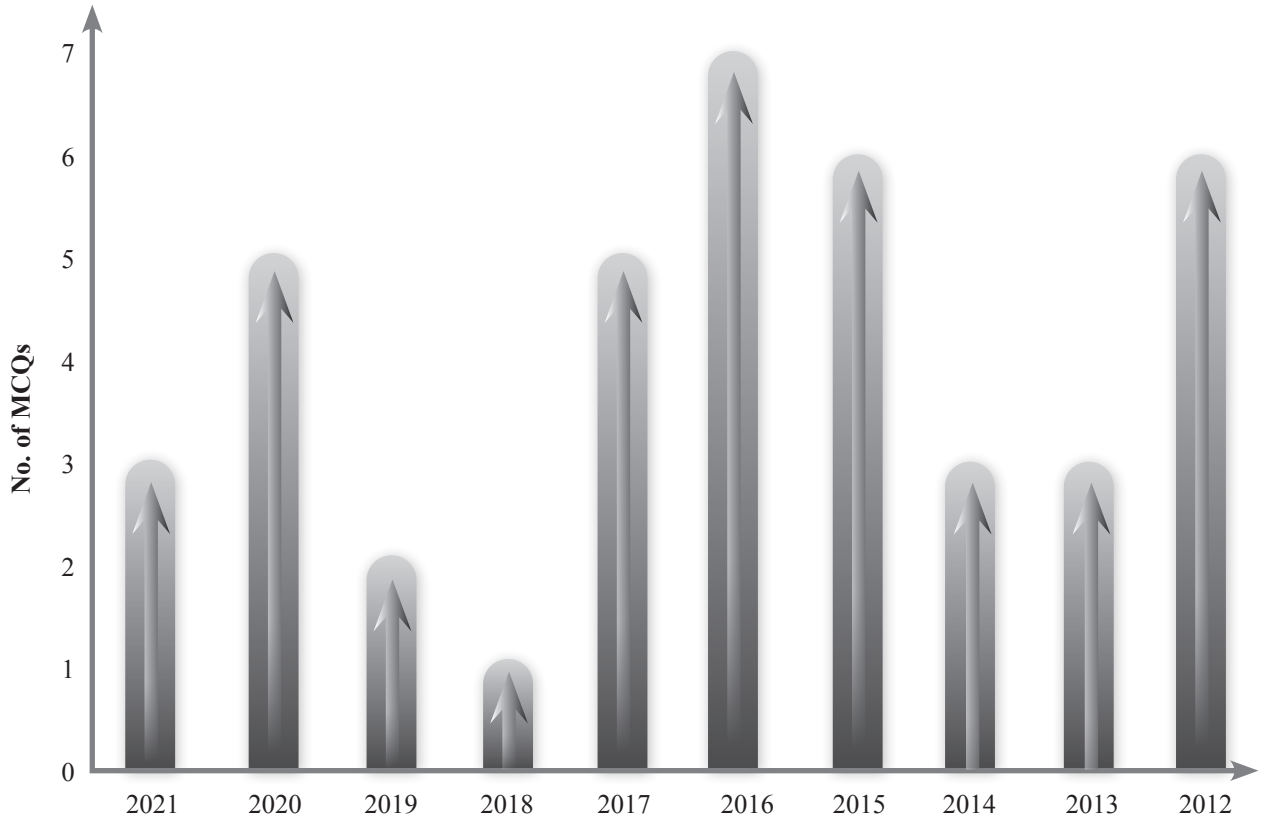
4 Chemical Bonding and Molecular Structure

Past Years NEET Trend



Investigation Report

TARGET EXAM	PREDICTED NO. OF MCQs	CRITICAL CONCEPTS
NEET	4-5	• Molecular orbital theory • Dipole moment, VSEPR Theory

Perfect Practice Plan

Topicwise Questions	Learning Plus	Multiconcept MCQs	NEET Past 10 Years Questions	Total MCQs
193	36	24	41	294

INTRODUCTION

An attractive force that acts between two or more particles (atoms, ions or molecules) to hold them together, is known as a chemical bond.

It is a union of two or more atoms to aquire stable inert gas configuration ns^2np^6.

In a Chemical Bond both attractive and repulsive forces exist in equilibrium.

Atoms are less stable and more energetic hence they form molecules by loosing some energy by participating in Chemical Bond.

$$H + H \rightarrow H_2 + 434.72kJ$$
$$Cl + Cl \rightarrow Cl_2 + 239.1kJ$$

Formation of bond is accompanied by decrease in potential energy. Therefore, Chemical Bond formation is always Exothermic.

To explain nature of chemical bond different theories are given, at first KOSSEL & LEWIS proposed the theory.

Kossel & Lewis approch chemical bonding: In early days, the ability of various elements to combine with one another was expressed in terms of their valency. The concept of valency was not based on any logical understanding. W. Kossel and G. N. Lewis were the first to provide some logical explanation of valency which is based on the inertness of noble gases.

This view later on, came to be known as Octet rule.

OCTET RULE
(Tendency to acquire noble gas configuration)

It has been observed that atoms of noble gases have little or no tendency to combine with each other or with atoms of other elements.

It means that these atoms must have a stable electronic configuration.

These elements (noble gases) have 8 electrons ($ns^2\ np^6$) except helium which has 2 electrons ($1s^2$) in their outer most shell.

Element Outer most shell	Configuration
Ne	$2s^22p^6$
Ar	$3s^2sp^6$
Kr	$4s^24p^6$
Xe	$5s^25p^6$
Rn	$6s^26p^6$

It is therefore concluded that ns^2np^6 configuration in the outer energy level constitutes a structure of maximum stability or minimum energy.

"Tendency of atoms to have eight electrons in their outermost shell is known as Lewis octet rule".To achieve inert gas configuration, atoms **lose**, **gain** or **share** electrons.

Limitations of the Octet Rule

It applies mainly to the second period elements of the periodic table. There are three types of exceptions to the octet rule:

1. The incomplete octet of the central atom: In some compounds, the number of electrons surrounding the central atom is less than eight. This is especially the case with elements having less than four valence electrons. Examples are LiCl, BeH_2 and BCl_3.

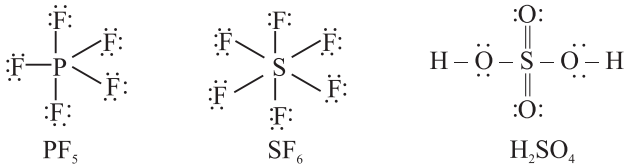


Li, Be and B have 1, 2 and 3 valence electrons only. Some other such compounds are $AlCl_3$ and BF_3.

2. Odd-electron molecules: In molecules with an odd number of electrons like nitric oxide. NO and nitrogen dioxide, NO_2, the octet rule is not satisfied for all the atoms



3. The expanded octet/ super octet / hypervalent compound: In some compounds, the number of electrons surrounding the central atom is more than eight due to availability of 3d orbitals (incomplete octet of the central atom).
Some of the examples of such compounds are: PF_5 SF_6, H_2SO_4 and a number of coordination compounds.



PF_5 — 10 electrons around the P atom

SF_6 — 12 electrons around the S atom

H_2SO_4 — 12 electrons around the S atom

FORMAL CHARGE

Formal charge is a factor based on a pure covalent bond formed by the sharing of electron pairs equally by neighbouring atoms. Formal charge may be regarded as the charge that an atom in a molecule would have if all the atoms had the same electronegativity. It may or may not approximate the real ionic charge. In case of a polyatomic ions, the net charge is possessed by the real ion as a whole and not by a particular atom. It is, however, feasible to assign a formal charge on an atom in a polyatomic molecule or ion.

The formal charge of an atom in a polyatomic molecule (or) ion may be defined as the difference between the number of valence electrons of that atom in an isolated or free state and the number electrons assigned to that atom in the lewis structure.

$$Q_f = [N_A - N_M] = [N_A - N_{LP} - 1/2N_{BP}]$$

Q_f = Formal charge

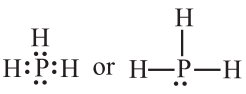
N_A = number of electrons in the valence shell in the free atom

N_M = number of electrons belonging to the atom in the molecule

N_{LP} = number of electrons in unshared pairs, i.e. number of electrons in lone pairs

N_{BP} = number of electrons in bond pairs, respectively.

Example: The Lewis dot formula of PH_3 is



Formal Charge of P

$$Q_f = [N_A - N_M] = [N_A - N_{LP} - 1/2\ N_{BP}]$$

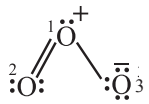
$\{5\text{-}2\text{-}1/2(6)\} = (5\text{-}5) = 0$

Formal Charge of H

$Q_1 = [N_A - N_M] = [N_A - N_{LP} - 1/2\ N_{BP}]$

$= (1\text{-}0\text{-}2/2)=0$

- Formal charges on oxygen atoms of ozone



- Formal charge of oxygen (1) = +1
- Formal charge of side oxygen atom(2) = 0
- Formal charge of side oxygen atom (3) = –1

The formal charge on an atom may or may not be same. It changes with the structural environment of the atom in the molecule. In resonating structures electronic environment changes. Hence the formal charge also may change.

Example: Resonating structures of N_2O

Example: $\overset{1}{:\ddot{N}} = \overset{2}{N} = \overset{3}{\ddot{O}:} \leftrightarrow :\overset{4}{N} \equiv \overset{5}{N} \rightarrow \overset{6}{\ddot{O}}:$

$$Q_{f_1} = 5 - 4 - \frac{1}{2}(4) = -1$$

$$Q_{f_2} = 5 - 0 - \frac{1}{2}(8) = +1$$

$$Q_{f_3} = 6 - 4 - \frac{1}{2}(4) = 0$$

$$Q_{f_4} = 5 - 2 - \frac{1}{2}(6) = 0$$

$$Q_{f_5} = 5 - 0 - \frac{1}{2}(8) = +1$$

$$Q_{f_6} = 6 - 6 - \frac{1}{2}(2) = -1$$

The formal charges do not represent real ionic charges.

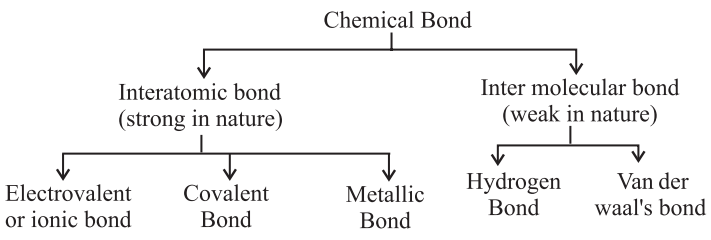
Formal charges represent a tendency to build up positive (or) negative charges.

The lowest energy structure is the one with the smallest formal charges on the atoms.

> ✍ **KEY NOTE**
>
> The most stable structure is the one which has the smallest formal charge on the atoms or zero formal charge on the atom.

CLASSIFICATION OF CHEMICAL BONDS

On the basis of electronic valency theory and structure, chemical bonds can be classified as follows:



Electronegativity of Elements		Changes occuring in the valence electrons	Nature of the bond formed
Low	High	Transfer of e^-	Ionic bond
High	High	Sharing of e^-	Covalent bond
Low	Low	Sea of e^- molecular orbital bond	Metallic bond

Bond		Bond Energy per mole
Ionic bond	———	700 – 4000 kJ
Covalent and coordinate bond	———	200 – 500 kJ
Hydrogen bonding	———	10 – 40 kJ
Van der–waal's	———	2 – 10 kJ

ELECTROVALENT OR IONIC BOND

The strong electrostatic forces of attraction between two oppositely charged ions which are formed due to transfer of electrons from one atom to another is called **Ionic Bond** (or) Electrovalent Bond.

Electrovalent bond is not possible between similar atoms. This type of bonding requires two atoms of different nature. One atom should have the tendency to loose electrons i.e. electropositive in nature and the other atom should have the tendency to accept electrons i.e. electronegative in nature.

Formation of Ionic bond is a redox process, because one atom undergoes oxidation and other one undergoes reduction.

$Mg \rightarrow Mg^{+2} + 2e^-$

$2F + 2e^- \rightarrow 2F^-$

$Mg^{+2} + 2F^- \rightarrow MgF_2$ or $Mg^{+2}(F^-)_2$

Example: IA and VII A group elements form strong ionic compound.

Na	+	Cl	→	Na^+	+	Cl^-
2, 8, 1		2, 8, 7		2, 8 (Ne) (configuration)		2, 8, 8 (Ar) (configuration)

Force of attraction is equal in all direction so ionic bond is non–directional.

A definite three dimensional structure are called crystal lattice.

Energy released during the formation of one mole crystal lattice is called lattice energy.

e.g. $(6.023 \times 10^{23})Na^+_{(g)} + (6.023 \times 10^{23})\ Cl^-_{(g)} \rightarrow$

$(6.23 \times 10^{23})\ NaCl(s) + 788\ KJ\ mol^{-1}$ (Lattice energy).

Ionic compounds do not have a molecular formula, they only have empirical formula.

> ✍ **KEY NOTE**
>
> - More the distance between two elements in the periodic table more will be the ionic character of the bond.
> - Total number of electrons lost or gained is called electrovalency.

Factors Favouring Formation of Ionic Bonds

1. Ionisation energy (I.E.): Amount of energy required to remove an electron from the outermost orbit of an isolated gaseous atom to form positive ion or cation is called ionzation energy [energy is absorbed so it is an endothermic process]

$M + I.E. \rightarrow M^+ + e^-$

Less Ionisation energy ⇒ Greater tendency to form cation.

Example : $Na^+ > Mg^{+2} > Al^{+3}$, $Cs^+ > Rb^+ > K^+ > Na^+ > Li^+$ } Cation formation tendency

2. Electron affinity: Amount of energy released when an electron is added to an isolated gaseous atom to form negative ion or anion is called electron affinity [energy is released so it is an exothermic process]

$$X + e^- \rightarrow X^- + EA$$

High electron affinity ⇒ Greater tendency to form anions

$Cl^- > F^- > Br^- > I^-$, $F^- > O^{-2} > N^{-3}$ } Anion formation tendency

LATTICE ENTHALPY

The Lattice Enthalpy of an ionic solid is defined as the energy required to completely separate one mole of a solid ionic compound into gaseous constituent ions or energy released during the formation of one mole crystal from its gaseous constituent ions. For example, the lattice enthalpy of NaCI is 788 kJ mol^{-1}. This means that 788 kJ of energy is required to separate one mole of solid NaCI into one mole of Na^{+1} (g) and one mole of Cl^-(g) to an infinite distance.

The solid crystal being 3D; it is not possible to calculate lattice enthalpy directly from the interaction of forces of attraction and repulsion only.

Factors associated with the crystal geometry have to be included.

$Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$

$\Delta H = \Delta H_{lattice} \Rightarrow$ –ve (always)

$Mg^{2+}(g) + 2Cl^-(g) \rightarrow MgCl_2(s)$

Factors affecting L.E.

(*i*) Lattice energy (L.E.) $\propto \frac{1}{r}$ $r = r^+ + r^- =$ interionic distance

(*ii*) L.E. $\propto Z^+, Z^-$

$Z^+ \Rightarrow$ charge on cation in terms of electronic charge

$Z^- \Rightarrow$ charge on anion in terms of electronic charge

(*iii*) Size of cation ↑ L.E. ↑

	Lattice energy of	Factor
(*a*)	NaCl > KCl	(size)
(*b*)	NaCl < MgO	(size, charge)
(*c*)	NaCl < $MgCl_2$	(size)

In size and charge, charge will dominate

$Na_2O > NaF$

$NaCl < Na_2S$

(*d*) Al_2O_3 Na_2O MgO

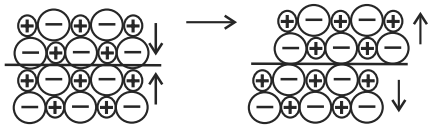
$Al_2O_3 > MgO > Na_2O$

> ✍ **KEY NOTE**
> Born - Haber cycle is used to calculate lattice energy.

Characteristics of Ionic Compounds

(i) Physical state:

- Electrovalent compounds are generally crystalline, hard & brittle in nature.
- These compounds are generally made from ions which are arranged in a regular way as a lattice structure.
- Thus electrovalent compounds exist as three dimensional solid aggregates.
- Normally each ion is surrounded by a number of oppositely charged ions and this number is called co-ordination number.



Attraction Repulsion

{Same charged ions comes near So they repel each other}

(ii) Boiling point and melting point: High boiling point and melting points are due to strong electrostatic force of attraction.

(iii) Electrical conductivity: It depends on ionic mobility. In solid state there are no free ions so they are bad conductors of electricity.

In fused state or aqueous solution free ions are present so they are good conductors of electricity.

Solid state < Fused state < Aqueous solution (Conductivity order)

(iv) Ionic reactions: Ionic compounds show ionic reactions & covalent compounds show molecular reactions.

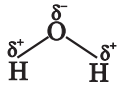
Ionic reactions are fast reactions.

Example: When NaCl is added to $AgNO_3$ solution, white ppt of AgCl is formed at once.

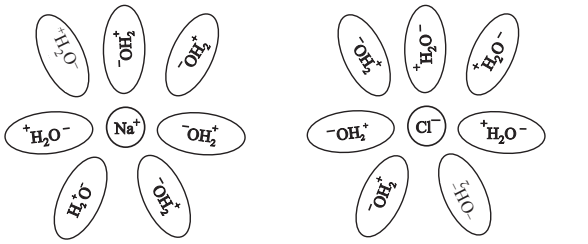
$$Ag^+NO_3^- + Na^+Cl^- \longrightarrow Na^+NO_3^- + \underset{\text{white ppt}}{AgCl\downarrow}$$

(v) Solubility: *Ionic compounds are soluble in polar solvent like H_2O, HF etc.*

To explain solubility of ionic compound consider an example of NaCl in water.

H_2O is polar solvent it can be represented as 

- The Na^+ ions gets associated with partially negatively charged 'O' of water.
- And Cl^- ions get associated with partially positively charged 'H' of water.

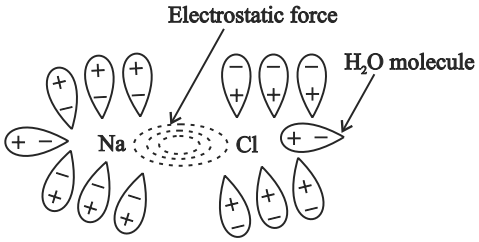


Oxygen atom of H_2O gives its electron to $\overset{\oplus}{Na}$ H atom of H_2O gain electron from $Cl^{\ominus}$

Thus charge on Na^+ and Cl^- decreases and electrostatic force also decreases which leads to free ion

- Here, attraction force between H_2O and Na^+ or Cl^- (Hydration energy) > Electrostatic force between ions in lattice (Lattice energy)
- Energy released due to interaction between polar solvent molecule and ions of solute is called solvation energy. If water used as solvent it is called hydration energy.
- Condition of the solubility of ionic compound in water is (Hydration energy > Lattice energy)

Solvation or Hydration



Whenever any compound generally ionic or polar covalent is dissolved in a polar solvent or in water then, different ions of the compound will get separated and will get surrounded by polar solvent molecules. This process is known as solvation or hydration. Energy released in this process is known as solvation energy or hydration energy

The ionic compound will be soluble only if solvation energy (H.E.) is more than the lattice energy

Factors affecting solvation energy or hydration energy

$$S.E \propto Z^+ Z^- \propto \left(\frac{1}{r+}+\frac{1}{r-}\right)$$

$$\propto \left(1-\frac{1}{\epsilon_r}\right) \text{ (nature of solvent) where } \epsilon_r \text{ is dielectric constant.}$$

Greater the polarity, greater will be ϵ_r

$\epsilon_r \uparrow \Rightarrow 1/\epsilon_r \downarrow$, $1-(1/\epsilon_r)\uparrow \Rightarrow$ S.E. $\uparrow$

	H_2O	CH_3OH	C_2H_5OH	C_6H_6
ϵ_r	81	34	27	3

Applications of hydration energy

- Size of the hydrated ions: Greater the hydration of the ion greater will be its hydrated radii.
  $Li^+(aq) > Na^+(aq)$
- Mobility of the ion: more is the hydration smaller will be the mobility of the ions.
  $Li^+(aq) < Na^+(aq) < K^+(aq) < Rb^+(aq) < Cs^+(aq)$.
- Electrical conductance: It is related to mobility so follows the same order.

COVALENT CHARACTER IN IONIC COMPOUNDS (FAJAN'S RULE)

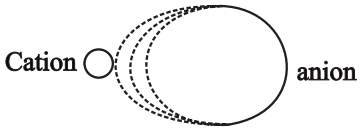
No Bond is 100% ionic in nature. It has some percentage of covalent character which is explained on the basis of Fajan's rule.

In a binary compound AB, if the electronegativity difference between the elements A and B is equal to 1.7, the compound AB is 50% ionic.

Greater than 1.7, the compound AB is an ionic compound less than 1.7, the compound AB is a covalent compound. (Exceptionally HF is covalent compound even though electronegativity diffrence is 1.7)

Polarising Power

When anion and cation approach each other, the valence shell of anion is pulled towards cation nucleus and thus shape of anion is deformed. This phenomenon of deformation of anion by a cation is known as polarisation and the ability of cation to polarize a nearby anion is called as polarizing power of cation.



Polarisability: The tendency of anion to get distorted or polarised by the cation is called its polarisability

Polarising relates to cation:

Polarising power $\propto$ charge on cation $\propto \frac{1}{size}$

Polarisability relates to anion

Polarisability $\propto$ charge $\propto$ Size

Polarising power increases covalent character increases

Ex: In case of NaCl, $MgCl_2$, $AlCl_3$ the polarisation increases $Na^+ < Mg^{2+} < Al^{3+}$

In case of $AlF_3 < AlCl_3 < AlBr_3 < AlI_3$

$F^- < Cl^- < Br^- < I^-$

Covalent character increases with increase in size of halide ion.

Fajan pointed out that greater is the polarization of anion in a molecule, more is covalent character in it.

More distortion of anion, more will be polarisation then covalent character increases.

Fajan gives some rules which govern the covalent character in the ionic compounds, which are as follows:

(*i*) **Size of cation :** Smaller is the cation more is its polarizing power and thus more will be the polarisation of anion. Hence more will be covalent character in compound.

Size of cation $\propto$ 1 / polarisation.

e.g. $BeCl_2$ $MgCl_2$ $CaCl_2$ $SrCl_2$ $BaCl_2$

Size of cation increases $\rightarrow$ Polarisation decreases $\rightarrow$ Covalent character decreases

(*ii*) **Size of anion :** Larger is the anion, greater is its polarisability and, therefore, more will be the polarisation. Thus more will be covalent character in compound.

Size of anion ∝ polarisation

$$e.g. \xrightarrow[\begin{array}{l}-\text{ Size of anion increases}\\ -\text{ Polarisation increases}\\ -\text{ Covalent character increases}\end{array}]{LiF \quad LiCl \quad LiBr \quad LiI}$$

(*iii*) **Charge on cation :** Higher is the oxidation state of cation, more will be the deformation of anion and thus, more will be covalent character in compound.

Charge on cation α polarisation.

$$e.g. \xrightarrow[\begin{array}{l}-\text{ Charge of cation increases}\\ -\text{ Polarisation increases}\\ -\text{ Covalent character increases}\end{array}]{\begin{array}{c}NaCl \quad MgCl_2 \quad AlCl_3\\ Na^+ \quad Mg^{2+} \quad Al^{3+}\end{array}}$$

(iv) **Charge on anion :** Higher is the charge on anion more will be the polarisation of anion and thus more will be covalent character in the compound.

Charge on anion ∝ polarisation.

$$e.g. \xrightarrow[\begin{array}{l}-\text{ Charge on anion increases}\\ -\text{ Polarisation increases}\\ -\text{ Covalent character increases}\end{array}]{\begin{array}{c}AlF_3 \quad Al_2O_3 \quad AlN\\ F^- \quad O^{2-} \quad N^{3-}\end{array}}$$

(v) **Pseudo inert gas configuration of cation :** Cation having pseudo inert gas configuration has more polarizing power than the cation that has inert gas configuration. Thus NaCl having inert gas configuration will be more ionic whereas CuCl having pseudo inert gas configuration will be more covalent in nature.

$Cu^+ =$ [Ne] $3s^2\ 3p^6\ 3d^{10}$ $\quad$ $Na^+ = 1s^2\ 2s^2\ 2p^6$

$18e^-$ $\quad$ $8e^-$

Pseudo inert gas configuration (poor shielding of d-electrons) $\quad$ Inert gas configuration (more shielding of s and p electrons)

Application & Exceptions of Fajan's Rules

(*i*) Ag_2S is less soluble than Ag_2O in H_2O because Ag_2 S is more covalent due to bigger S^{2-} ion.

(*ii*) $Fe(OH)_3$ is less soluble than $Fe(OH)_2$ in water because Fe^{3+} is smaller than Fe^{2+} and thus charge is more. Therefore, $Fe(OH)_3$ is more covalent than $Fe(OH)_2$.

(*iii*) The colour of some compounds can be explained on the basis of polarisation of their bigger negative ions.

(*iv*) Variation of melting point

[melting point of covalent compound< melting point of ionic compound]

$$BeCl_2,\ MgCl_2,\ CaCl_2,\ SrCl_2,\ BaCl_2 \longrightarrow$$

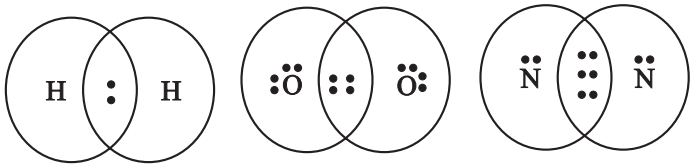
Ionic character increases, melting point increases ; since size of cations increases & size of anion is constant.

$$CaF_2,\ CaCl_2,\ CaBr_2,\ CaI_2 \longrightarrow$$

Covalent character increases, melting point decreases ; since size of anions increases & size of cation is constant.

COVALENT BOND

A covalent bond may be defined as the bond formed by mutual sharing of electrons between the participating atoms (which are short of electrons) of comparable electronegativity.



H_2 molecule H—H $\quad$ O_2 O = O $\quad$ N_2 N ≡ N

Covalency

(*i*) It is defined as the number of covalent bonds formed by the atom of the element with other atoms.

(*ii*) The usual covalency of an element except hydrogen is equal to 8 minus the number of group to which the element belongs.

Ex: In NH_3 the valency of Nitrogen is 3

In CO_2 the valency of carbon is 4

Lewis Structure and Covalent Bond

(*i*) Structures in which valence electrons are represented by dots are called Lewis structures.

(*ii*) All atoms in the formulae will have eight electrons in its valence shell. H atom is an exception.

(*iii*) Lewis dot formulae shows only the number of valence electrons, the number and kinds of bonds, but does not depict the three dimensional shapes of molecules and polyatomic ions.

(*iv*) Lewis formulaes are based on the fact that the representative elements achieve a noble gas configuration in most of their compounds i.e. 8 electrons in their outermost shell (except for H_2, Li^+ and Be^{2+} ions which have only 2 electrons)

No. of valence electrons	Example	Lewis structure
1	Hydrogen/Group IA	H·/Li·
2.	Helium/Group IIA	He:/Mg:
3.	Group IIIA	·B·
4.	Group IV A	·C·
5.	Group V A	·N·
6.	Group VI A	:O·

Factors Favouring the Formation of Covalent Bond

(1) Small size of cation

(2) Large size of anion

(3) Large charge on cation and Anion

(4) Generally a covalent bond is formed between 2 atoms of similar E.N. values

High ionisation enthalpy: Atoms having high ionisation enthalpies do not have a tendency to lose electrons to form cations. Such elements prefer to form covalent bonds. Eg: Cl-Cl

Comparable electronegativity: Atoms with equal or nearly equal electronegativities tend to share equally a pair of electrons with opposite spins. This leads to a better overlap of their atomic orbitals thus resulting in the formation of a covalent bond.

Note: Sidgwick concept of maximum covalency

The maximum covalency of an element actually depends on the period of the periodic table to which it belongs.

Period	Elementt	Max. Covalency
1	H	2
2	Li to F	4
3	Na to Cl	6
4	K to Br	6
5 6	Rb onwards and rest of the elements	8

✍ KEY NOTE

This rule explains the formation of PCl_5 and SF_6. This also explains why Nitrogen does not form NF_5 or NCl_5. Because Nitrogen belongs to second period and the maximum covalency of Nitrogen is 4.

Characteristics of Covalent Compounds

(i) Physical state: Under normal temperature and pressure, they exist as gases or liquids of low boiling points.

This is due to the fact that very weak forces of attraction (Van der waal's forces) exist between the molecules due to which molecules are far from each other.

If their molecular masses are high they exist as soft solids e.g. sulphur, phosphorus and iodine.

(ii) Crystal structures: Covalent solid – In this type of structure every atom is bonded to four other atoms by single covalent bonds resulting in the formation of a giant structure e.g. SiC, AlN and diamond. These crystals are very hard and possess high melting point.

Diamond is sp^3 hybridized carbon atom and it forms a tetrahedral structure.

Graphite is sp^2 hybridized carbon atoms and it forms hexagonal layers which can slide over each other due to weak Van der waal's forces of attraction. Distance between C–C atom is 1.42 Å distance between layers is 3.6 Å. **Graphite is more stable than diamond due to high value of change in enthalpy.**

Molecular solids: They are formed when one atom combines with another by a covalent bond and then the molecule combines with another similar molecule with the help of Van der waal's force of attraction or hydrogen bond.

Example : CH_4(Solid), dry ice, ice.

(iii) Melting and boiling points: With the exception of a few which have giant three dimensional structures such as diamond, carborundum (SiC), Silica (SiO_2), others have relatively low melting and boiling points.

This is due to the presence of weak attractive forces between the molecules.

(iv) Electrical conductivity: In general, covalent substances are bad conductors of electricity. Since they do not contain charged particles or free electrons.

Substances which have polar character like HCl in a solution, can conduct electricity.

Graphite can conduct electricity since electrons can pass from one layer to other.

Some show conductivity due to self ionisation, example: Liq.NH_3, H_2O, etc.

$$NH_3 + NH_3 \rightarrow NH_4^+ + NH_2^-$$

$$H_2O + H_2O \rightarrow H_3O^+ + OH^-$$

(v) Chemical reactions: Covalent substances give molecular reactions. Reaction rate is usually low because it involves two steps (i) breaking of covalent bonds of the reactants (ii) establishing new bonds. While in ionic reactions there is only regrouping of ions.

(vi) Solubility: Non polar compounds are soluble in non polar solvents. Non polar solvents are CCl_4, benzene, CS_2 etc.

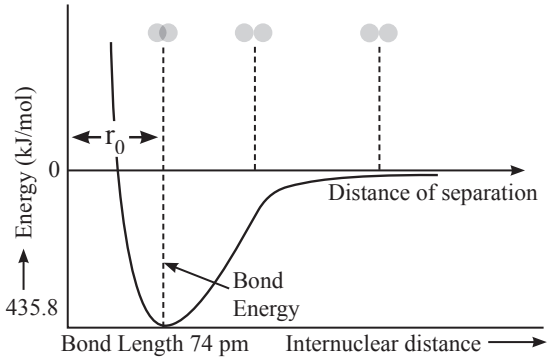
Polar compounds are soluble in polar solvents (i.e. ΔE.N. is 0.9 to 1.8). Polar solvents are H_2O, $CHCl_3$, alcohol etc.

(vii) Isomerism: Covalent bonds are rigid and directional. On account of this there is a possibility of different arrangement of atoms in space. Covalent compounds can thus shows isomerism (structural and spatial).

MODERN CONCEPT OF COVALENT BOND (VBT)

Enthalpy Diagram for H_2 Molecule

When two hydrogen atoms are at an infinite distance from each other, there is no interaction between them and therefore, enthalpy of the system is assumed to be zero in this state (stage-A). As the two atoms start coming closer to each other, the potential enthalpy continues to decrease (stage B). Ultimately a stage is reached when the enthalpy of the system becomes minimum and hydrogen atoms are said to be bonded together to form a stable H_2 molecule (stage C).

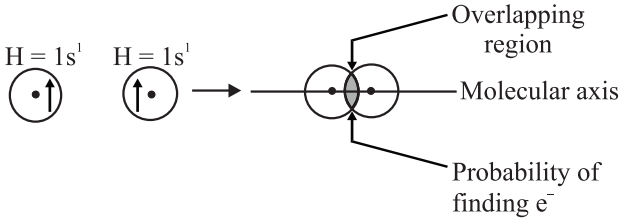


The internuclear distance r_0 between two hydrogen atoms at this stage is referred to as bond length. In the case of the hydrogen molecule, the bond length is 74 pm. It should be noted that two hydrogen atoms can not be brought at a distance lesser than r_o (i.e., 74 pm) because the potential enthalpy of the system increases and curve shows an upward trend (dotted lines) and molecule becomes unstable.

VBT cannot explain paramagnetic behavior of O_2.

Main Points of Valence Bond Theory

(*i*) A covalent bond is formed by partial overlapping of two atomic orbitals



(*ii*) More is the extent of overlapping between the two atomic orbital, stronger will be bond.



[Principal Quantum no. same, n = 2]

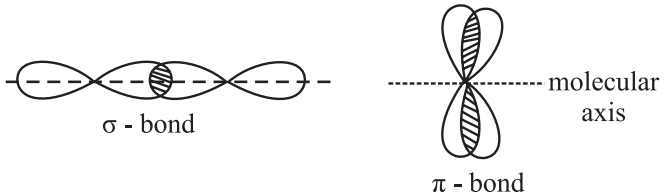
∵ s orbital are spherical in nature so they are least diffused hence it will provide less area for overlapping.

(*iii*) Orbitals which are undergoing overlapping must be such that

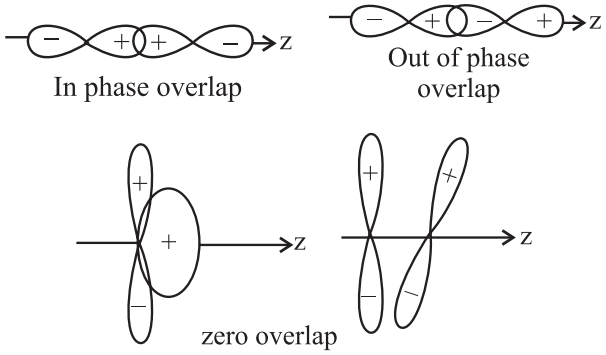
(*a*) Each orbital should have one electron with opposite spin (for formation of covalent bond).

(*b*) One orbital have pair of electron and the other orbital have no electron (for formation of co-ordinate bond).

(*iv*) If the overlapping is along the molecular axis then bond will be sigma (σ) & in the perpendicular direction, it will be pi(π) bond.



(*v*) If two atomic orbitals having same sign (+ or –) overlap with each other, it is called in phase overlapping (constructive). If both orbitals having opposite sign, called out of phase overlap (destructive). And if both orbitals are out of phase due to the direction of approach then zero overlap.



Types of Overlapping and Nature of Covalent Bonds

- Linear overlapping of atomic orbitals results in the formation of sigma bond (σ).
- A covalent bond formed by the sidewise overlap of atomic orbitals of atoms already bonded through a 'σ' bond and in which the electron cloud is present on either side of the inter nuclear axis is known as a pi bond (π).
- Lateral overlapping of atomic orbitals results in the formation of pi bond (π).
- 'π' bond is formed only after the formation of 'σ' bond.
- Any type of orbitals can involve in 'σ' bond formation.
- Only 'p' or 'd' orbitals can involve in 'π' bond formation.
- Single bond is equal to one 'σ' bond.
- Double bond is a combination of one 'σ' bond and one 'π' bond.
- Triple bond is a combination of one 'σ' bond and two 'π' bonds.

Strength of Sigma and pi Bonds

Strength of the bonds follows the order

triple bond > double bond > single bond.

Strength of the bonds follows the order : $\sigma_{p-p} > \sigma_{s-p} > \sigma_{s-s} > \pi$

> ✍ **KEY NOTE**
>
> 'σ' bond between '1s' orbitals is exceptionally stronger

Difference in σ and π Bonds

	Sigma (σ) bond	Pi (π) bond
1.	It results from the end to end overlapping of two s orbitals or two p–orbitals or one s and one p-orbital	It result from the sidewise (lateral) overlapping of two p orbitals.
2.	Its bonded orbital consists of a single electron cloud symmetrical about internuclear axis	Its bonded orbital consists of two electron clouds one above and the other below the plane of symmetry
3.	Strong	Weak
4.	Bond energy 80 kcals	65 kcals
5.	More stable	Less Stable
6.	Less reactive	More reactive

	Sigma (σ) bond	Pi (π) bond
7.	Can independently exists	Always exists along with a σ bond hybridsation
8.	Hybridization depends upon σ bond	doesn't depend on π bond
9.	The groups or atoms can undergo bond rotation about single sigma (σ) bonds	Due to resistance to rotation around the π bond the groups attached to it are not free to rotate.
10.	The σ electrons are referred as localised	In π bond the electrons are held less firmly and thus can be easily dissociated or polarised by an external charge and hence the π e^- are referred as mobile electrons.
11.	Shape of the molecule is determined by the σ bonds present in the molecule	π bonding does not affect the shape of the molecule

TRAIN YOUR BRAIN

Q. Identify the correct statement :

(*a*) Single N–N bond is stronger than single P–P bond

(*b*) Single N–N bond is weaker than single P–P bond

(*c*) N ≡ N is weaker than P ≡ P

(*d*) None of these

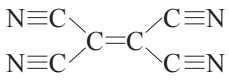
Ans. (*b*) Due to small size of nitrogen, the lp–lp repulsion is more than that in P. Hence statement B is correct

Q. How many sigma and pi bonds are present in tetracyanoethylene?

(*a*) Nine σ and nine π (*b*) Five π and nine σ

(*c*) Nine σ and seven π (*d*) Eight σ and eight π

Ans. (*a*)



VALENCE SHELL ELECTRON PAIR REPULSION (VSEPR) THEORY

Postulates

The shape of a molecule depends upon the number of valence shell electron pairs (bonded or non-bonded) around the central atom.

Pairs of electrons in the valence shell repel one another since their electron clouds are negatively charged.

These pairs of electrons tend to occupy such positions in space that minimise repulsion and thus maximise distance between them.

The valence shell is taken as a sphere with the electron pairs localising on the spherical surface at maximum distance from one another.

A multiple bond is treated as if it is a single electron pair and the two or three electron pairs of a multiple bond are treated as a single super pair.

Where two or more resonance structures can represent a molecule, the VSEPR model is applicable to any such structure.

The repulsion between electron pairs on central atom follows the order

l.p-l.p>l.p-b.p>b.p-b.p.

In VSEPR theory the number of electron pairs in single bond or double bond or triple bond or dative bond is counted as only one pair because all the electron pairs in the same bond are oriented in the same direction.

The magnitude of repulsions between bond pair of electrons depends on the electronegativity difference between the central atom and other atom.

Identification of hybridisation and shape of a molecule or ion having single central atom.

Rules:

Step-1: Calculate valence electrons in central atom.

Step-2: If it is cation then substract charge from central atom

Step-3: If it is anion then add charge to the central atom.

Step-4: Bond pairs = No.of atoms attached to central atom.

Step-5: Lone pairs =

$$(\text{Valence } e^- \text{ on central atom}) - \left(\frac{\textit{No. of electrons given by bonded atoms}}{2}\right)$$

Step-6: B.P+L.P = Electron Pairs

2	3	4	5	6	7
↓	↓	↓	↓	↓	↓
sp	sp^2	sp^3	sp^3d	sp^3d^2	sp^3d^3

✍ KEY NOTE

If bonded atoms are H, F then each one contribute one electron for bond formation. If bonded atoms are O, S then each one contribute two electrons for bond formation.

Shape (molecular geometry) of Some Simple Molecules/Ions

No. of B.Ps	No of L.Ps	Formula	Molecular shape	Example
2	0	AB_2	Linear	$BeCl_2$, BeF_2, C_2H_2
3	0	AB_3	Trigonal Planar	BF_3, BCl_3
4	0	AB_4	Tetrahedral	CH_4, NH_4^+, CCl_4
5	0	AB_5	Trigonal bipyramidal	PCl_5
6	0	AB_6	Octahedral	SF_6
2	1	AB_2E	Bent	SO_2, O_3
3	1	AB_3E	Trigonal Pyramidal	NH_3
2	2	AB_2E_2	Bent	H_2O

4	1	AB_4E	See Saw	SF_4
3	2	AB_3E_2	T-Shape	ClF_3
5	1	AB_5E	Square Pyramidal	BrF_5
4	2	AB_4E_2	Square Planer	XeF_4

A = Central atom in the compound.

B = atom linked to the central atom

E = Lone pairs of electron

TRAIN YOUR BRAIN

Q. The interhalogen compound that cannot exist is

(*a*) IBr_3 (*b*) ICl_7 (*c*) IF_4 (*d*) BrF_5

Ans. (c) I does not have its valency 4. It has valency 1, 3, 5 & 7

Q. How many number of species do not exist :-

PCl_5, PH_5, SF_6, SH_6, PBr_6^-, OF_4

(*a*) 2 (*b*) 3

(*c*) 4 (*d*) 5

Ans. (*c*) PH_5, SH_6 → not exist due to non expandable covalency

OF_4 → O-atom does not have vacant d-orbital.

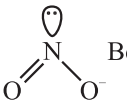
PBr_6^- → steric repulsion

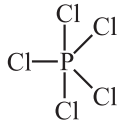
TRAIN YOUR BRAIN

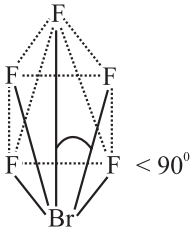
Q. Which of the following pairs of species have identical shapes?

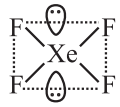
(*a*) NO_2^+ and NO_2^- (*b*) PCl_5 and BrF_5

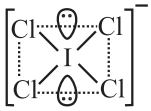
(*c*) XeF_4 and ICl_4^- (*d*) $TeCl_4$ and XeO_4

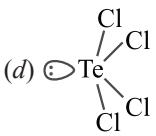
Ans. (c) (*a*) $\overset{\text{linear}}{O=\overset{+}{N}=O}$  Bent or V-shaped

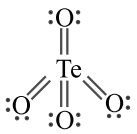
(*b*)  Trigonal bipyramidal

 Square pyramidal or Distorted octahedral

(*c*)  Square planar

 Square planar

(*d*)  See-Saw

 Tetrahedral

HYBRIDISATION

- The concept of hybridisation was introduced by Pauling
- The intermixing of atomic orbitals of almost same energy and their redistribution into an equal number of identical orbitals is known as hybridisation
- The orbitals of one and the same atom only involve in hybridisation.
- The no.of hybrid orbitals formed is numerically equal to no.of orbitals participating in hybridisation.
- The hybrid orbitals symmetrically arranged around the nucleus such that they have maximum stability.
- The order of repulsions will be in order of lp-lp> lp-bp>bp-bp repulsions.
- The orbitals involving in the hybridisation have different shapes but almost same energy.
- The orbitals formed in hybridisation process are called hybrid orbitals.
- The hybrid orbitals have same shape and same energy.
- The angle between any two hybrid orbitals in an atom is generally same.
- Electron filling in hybrid orbitals obeys Hunds rule and Pauli's rule.
- The hybrid orbitals involve only in σ bond formation. They never involve in 'π' bond formation.
- The concept of hybridisation was introduced to explain the shapes of molecules, bond angles and bond lengths in the molecules.
- A half filled or completely filled or even vacant orbital can participate in Hybridization.
- The hybrid orbitals are more effective in forming stronger bonds that leads to the formation of more stable molecules.

✍ KEY NOTE

Same atom undergoes different types of hybridisation under different situations.

More the directional bond, greater is the bond strength

$sp^3 - sp^3 > sp^2 - sp^2 > sp - sp > p - p > s - p > s - s$

sp-hybridisation

- This is also known as diagonal hybridisation or linear hybridisation.
- One s- orbital combines with one p- orbital to give two identical orbitals called sp - hybrid orbitals.
- The angle between the two sp-hybrid orbitals in an atom is 180°
- sp- hybrid orbital will have $\frac{1}{2}$ s- character or 50% s- character. $\frac{1}{2}$ p- character or 50% p- character.
- The hybridisation of central atom in a linear molecule is generally sp.

sp^2-hybridisation

- This is also known as trigonal hybridisation.
- One s- orbital combines with two p-orbitals to give three identical orbitals called sp^2-hybrid orbitals.
- The angle between any two sp^2-hybrid orbitals in an atom is 120°.
- sp^2 hybrid orbital will have $\frac{1}{3}$rd s-character or 33.3% s-character and $\frac{2}{3}$rd p- character or 66.7% p-character.
- The hybridisation of central atom in a molecule having trigonal planar shape is sp^2.

sp^3- hybridisation

- This is also known as tetrahedral hybridisation or tetragonal hybridisation.
- One s- orbital combines with three p-orbitals to give four identical orbitals called sp^3 hybrid orbitals.
- The angle between any two sp^3-hybrid orbitals in an atom is 109°28'.
- sp^3-hybrid orbital will have $\frac{1}{4}$ th s-character or 25% s-character. $\frac{3}{4}$ th p-character or 75% p-character.
- The hybridisation of central atom in a molecule having tetrahedral shape is sp^3.

dsp^2 hybridisation

- The orbitals involved in hybridisation are $d_{x^2-y^2}$, s and two p-orbitals having square planar geometry, bond angle between any two dsp^2 orbitals is 90°.

sp^3d-hybridisation

- One s-orbital, three p-orbitals and one d-orbital combines and gives five identical orbitals called sp^3d-hybrid orbitals.
- The angle between two sp^3d hybrid orbitals in an atom is 90° and 120°.
- sp^3d hybrid orbital will have $\frac{1}{5}$th s-character or 20% s-character, $\frac{3}{5}$th p-character or 60% p-character, $\frac{1}{5}$th d-character or 20% d-character.
- The hybridisation of central atom in a molecule having trigonal bipyramid shape is sp^3d.

  Example for sp^3d hybridisation is PCl_5
- $3d_{z^2}$ orbital is involved in sp^3d hybridisation.
- The axial bonds (2.19Å) have been found to be larger than equatorial bonds (2.04Å) since axial P-Cl bonds experience greater repulsion from equatorial P-Cl bonds.

  No. of planar atoms in PCl_5 is 4.

  PCl_5 is highly reactive and in the solid state exist as $[PCl_4]^+$ and $[PCl_6]^-$ ions.

sp^3d^2- hybridisation

- One s-orbital , three p-orbitals and two d- orbitals combines to give six identical orbitals called sp^3d^2 hybrid orbitals.
- The angle between any two sp^3d^2 hybrid orbitals is 90° and 180°.
- sp^3d^2 hybrid orbital will have $\frac{1}{6}$th s-character or 16.7% s-character, $\frac{1}{2}$ p – character & $\frac{1}{3}$ d – character

  Example for sp^3d^2 hybridisation is SF_6
- $3d_{z^2}$ and $3d_{x^2-y^2}$ orbitals are involved in sp^3d^2 hybridisation.

Assigning the Type of Hybridisation

(I) Total molecular valency method

Step 1: Total number of valence electron in the molecule is calculated by adding the individual valency of all constituent atoms.

Step 2: If total electrons are greater than eight, then divided by eight and quotient is equal to hybridised orbitals.

e.g:.PCl_5

Total valence electrons = 5+(5 × 7)=40

40 / 8=5=no.of hybridised orbitals → sp^3d

If remainder comes, that remainder should be divided by 2, initial quotient comes then sum of quotient and final quotient is equal to number of hybridised orbitals, e.g., $XeOF_2$

Total valence electrons = 8 + 6 + (2 x 7) = 28

28 / 8 Quotient=3 ; Remainder = 4

Remainder is divided by 2

4 / 2 Quotient = 2

Hence, sum of Quotient = 3+2 = 5 = no. of hybridised orbitals → sp^3d

Step 3: If total valence electrons are either eight or less is divided by 2,what Quotient appears,should be considered as number of hybridised orbitals, e.g., H_2O

Total valence electrons = 2 + 6 = 8

8 / 2 = 4 = sp^3

(II) Total number of Hybridised orbitals(For neutral molecule)

= 1/2(V + M)

V = no. of valence electrons

M = no. of monovalent atom

Total number of Hybridised orbitals(For cationic species)

= 1/2(V + M – C)

Total number of Hybridised orbitals(For anionic species)

= 1/2(V + M + A)

C = +ve charge. A = -ve charge.

Hybridisation in Carbon Compounds

- The carbon atom in carbon compounds is tetravalent (i.e., carbon atoms form four covalent bonds).
- The sp^3 carbon atom involves in four single bonds, 

- The sp^2 carbon atom involves in a double bond and two single bonds, $=\overset{|}{C}-$
- The sp carbon atom involves in two double bonds or one triple bond and one single bond $= C =$ or $\equiv C -$

Steric number	Types of Hybridisation	Geometry
2	sp	Linear
3	sp^2	Trigonal planar
4	sp^3	Tetrahedral
5	sp^3d	Trigonal bipyramidal
6	sp^3d^2	Octahedral
7	sp^3d^3	Pentagonal bipyramidal

TRAIN YOUR BRAIN

Q. In which of the following central atom is unhybridised?

(*a*) $S(CH_3)_2$ (*b*) SO_2

(*c*) SiH_4 (*d*) PCl_3

Ans. (*a*) As per Drago's rule

MOLECULAR ORBITAL THEORY

It was proposed by Hund and Mulliken to explain paramagnetic nature of oxygen molecule.

The electrons in a molecule are present in various molecular orbitals as the electrons of atoms are present in the various atomic orbitals.

Molecular orbitals are formed by linear combination (LCAO) of atomic orbitals of comparable energies and proper symmetry. This means that 1s orbital can combine with another 1s orbital but not with 2s orbital because the energy of 2s orbitals is higher than that of 1s orbital. This is true only for homonuclear diatomic molecules.

By convention z-axis is taken as the molecular axis, $2p_z$ orbital of one atom can combine with $2p_z$ orbital of the other atom but not with the $2p_x$ or $2p_y$ orbitals because of their symmetry.

Molecular orbital is polycentric i.e. an electron in molecular orbital is under the influence of two or more nuclei.

The number of molecular orbitals formed is equal to the number of combining atomic orbitals. When two atomic orbitals combine, two molecular orbitals are formed. One is known as bonding molecular orbital while the other is called anti bonding orbital.

It may be noted that the bonding molecular orbital is stabilized to the same extent as the anti bonding molecular orbital is destablised i.e. energy lost by BMO is equal to energy gained by ABMO but total energy of two MO equal to sum of energy of two atomic orbitals.

Order of energies of various orbitals is bonding orbitals < non bonding orbitals < anti bonding oritals.

The molecular orbitals like atomic orbitals are filled in accordance with the Aufbau principle obeying the Pauli's exclusion principle and the Hunds rule.

The molecular orbitals are named as σ, π, δ etc.

According to wave mechanics, the atomic orbitals can be expressed by wave functions (ψ), which represent the amplitude of waves. These are solutions of schrodinger wave equation. However it is difficult to solve the above exactly for many electron systems. To over come this problem scientists have used an approximate method known as LCAO (Linear Combination of Atomic Orbitals) method.

The atomic orbitals of two hydrogen atoms A and B may be represented by wave functions ψ_A and ψ_B .

Mathematically, formation of molecular orbitals may be described as

$\psi_{MO} = \psi_A \pm \psi_B$

$\sigma = \psi_A + \psi_B$ (Bonding molecular orbital)

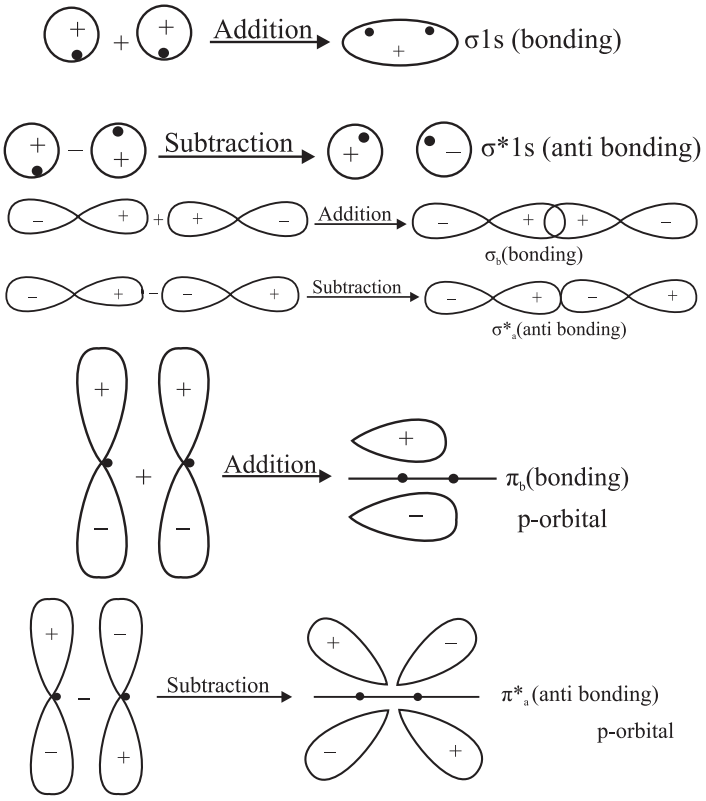
$\sigma^* = \psi_A - \psi_B$ (Antibonding molecular orbital)

Quantitatively, the formation of molecular orbitals can be understood in terms of the constructive or destructive interference of the electron waves of the combining atoms.

During the formation of BMO, the two electron waves of the bonding atoms reinforce each other due to constructive interference.

During the formation of ABMO, the electron waves cancel each other due to destructive interference.

As a result,the electron density in a bonding molecular orbital is located between the nuclei of the bonded atoms but in case of antibonding molecular orbital most of the electron density is located away from the space between nuclei(nodal plane).



The increasing order of relative energies of M.O having less than or **equal to 14 electrons.**

$\sigma 1s < \sigma^*1s < \sigma 2s < \sigma^*2s < \pi 2p_x = \pi 2p_y < \sigma 2p_z < \pi^*2p_x = \pi^*2p_y < \sigma^*2p_z$

for more than 14 electrons

$\sigma 1s < \sigma^*1s < \sigma 2s < \sigma^*2s < \sigma 2p_z < [\pi 2p_x = \pi 2p_y] < [\pi^*2p_x = \pi^*2p_y] < \sigma^*2p_z$

Difference between σ and π MO's

	σ-molecular orbital		π- molecualr orbital
1.	Formed by the end on overlap along the internuclear axis	1.	Formed by the sidewise overlap perpendicular to internuclear axis
2.	Overlapped region is very large	2.	Overlapped region is small
3.	Rotation about the internuclear axis is symmetrical	3.	Rotation about the internuclear axis is unsymmetrial
4.	Strong bonds are favoured	4.	Weak bonds are favoured

✍ KEY NOTE

If all the electrons in the molecule or ion are paired it is diamagnetic in nature.If it is unpaired that is paramagnetic.

Magnetic moment $\mu = \sqrt{n(n+2)}\ B.M$

Stability of Molecules

If N_b is the number of electrons occupying bonding orbitals and N_a the number of electrons occupying antibonding orbitals, then

A positive bond order is $N_b > N_a$ means a stable molecule. While a negative ($N_b < N_a$) or zero ($N_b = N_a$) bond order means an unstable molecule

Electronic configuration / Bond order of simple diatomic molecules

The electronic configuration and the bond order in case of simple diatomic molecules can be obtained by filling the molecular orbitals by applying Aufbau principle and Hund's rule etc.

Bond Order

The relative stability of a molecule can be determined on the basis of bond order. It is defined as the number of covalent bonds in a molecule. It is equal to one half of the difference between the number of electrons in the bonding and antibonding molecular orbitals.

Bond order = $\frac{1}{2}$ [Number of bonding electrons – Number of antibonding electrons]

or, $= \frac{N_b - N_a}{2}$

Bond order ∝ Bond energy

∝ stability of a molecule

$\propto \frac{1}{bond\ length}$

Bond order of $N_2 = \frac{10-4}{3} = 3$

Bond order of $N_2^+ = \frac{9-4}{2} = 2.5$

Bond order of $N_2^{2+} = \frac{8-4}{2} = 2.0$

Bond order of $N_2^- = \frac{10-5}{2} = 2.5$

Bond order of $N_2^{2-} = \frac{10-6}{2} = 2$

As the bond order increases the stability of molecule increase. However, some of the bond orders are identical. In such case, a molecule or ion with more number of electrons in their antibonding orbitals is less stable. Hence correct order is

$N_2^{2-} < N_2^{2+} < N_2^- < N_2^+ < N_2$.

The bond orders of 1,2 or 3 correspond to single, double or triple bond. But bond order may be fractional in some cases.

The magnetic properties of molecules can also be ascertained

Bonding in some diatomic molecules and ions

Hydrogen molecule (H_2)

Total number of electrons = 2, filling in molecular orbitals we have

$\sigma_{1s}^2 < \sigma_{1s}^{*0}$

Bond order = $\frac{(N_b - N_a)}{2} = \frac{2-0}{2} = 1$

Hence there is a single bond between two hydrogen atoms and due to absence of unpaired electrons it is **diamagnetic**.

Helium molecule (He_2)

The total number of electrons = 4 and filling in molecular orbitals we have

$\sigma_{1s}^2 < \sigma_{1s}^{*2}$

Bond order = $\frac{(N_b - N_a)}{2} = \frac{2-2}{2} = 0$

Hence, He_2 molecule can not exist.

Nitrogen molecule (N_2)

The total number of electrons =14 and filling in molecular orbitals we have

$\sigma_{1s}^2\ \sigma_{1s}^{*2}\ \sigma_{2s}^2\ \sigma_{2s}^{*2} \left\{ \begin{matrix} \pi 2p_x^2 \\ \pi 2p_y^2 \end{matrix} \right\} \sigma_{2p_z}^2$

Bond order = $\frac{(N_b - N_a)}{2} = \frac{10-4}{2} = 3$

It is **diamagnetic**.

Oxygen molecule (O_2)

Total number of electrons =16 and electronic configuration is

$$\sigma_{1s}^2\ \sigma_{1s}^{*2}\ \sigma_{2s}^2\ \sigma_{2s}^{*2}\ \sigma_{2p_z}^2 \begin{Bmatrix}\pi 2p_x^2\\ \pi 2p_y^2\end{Bmatrix}\begin{Bmatrix}\pi^* 2p_x^1\\ \pi^* 2p_y^1\end{Bmatrix}\sigma_{2p_z}^*$$

$$\text{Bond order} = \frac{(N_b - N_a)}{2} = \frac{10-6}{2} = 2$$

As shown by electronic configuration the O_2 molecule contains two unpaired electrons, hence it is **paramagnetic** in nature.

O^+_2 ion

Total number of electrons (16 - 1) = 15,

Electronic configuration

$$\sigma_{1s}^2\ \sigma_{1s}^{*2}\ \sigma_{2s}^2\ \sigma_{2s}^{*2}\ \sigma_{2p_z}^2 \begin{Bmatrix}\pi 2p_x^2\\ \pi 2p_y^2\end{Bmatrix}\begin{Bmatrix}\pi^* 2p_x^1\\ \pi^* 2p_y\end{Bmatrix}\sigma_{2p_z}^*$$

$$\text{Bond order} = \frac{10-5}{2} = 2.5$$

It is **paramagnetic.**

O_2^- Super oxide ion

Total number of electrons (16 +1) = 17. Electronic fonguration

$$\sigma_{1s}^2\ \sigma_{1s}^{*2}\ \sigma_{2s}^2\ \sigma_{2s}^{*2}\ \sigma_{2p_z}^2 \begin{Bmatrix}\pi 2p_x^2\\ \pi 2p_y^2\end{Bmatrix}\begin{Bmatrix}\pi^* 2p_x^2\\ \pi^* 2p_y^1\end{Bmatrix}\sigma_{2p_z}^*$$

$$\text{Bond order} = \frac{(N_b - N_a)}{2} = \frac{10-7}{2} = 1.5$$

It is **Paramagnetic.**

Peroxide ion (O_2^{2-})

Total number of electrons (16 + 2) = 18. The electronic configuration is

$$\sigma_{1s}^2\ \sigma_{1s}^{*2}\ \sigma_{2s}^2\ \sigma_{2s}^{*2}\ \sigma_{2p_z}^2 \begin{Bmatrix}\pi 2p_x^2\\ \pi 2p_y^2\end{Bmatrix}\begin{Bmatrix}\pi^* 2p_x^2\\ \pi^* 2p_y^2\end{Bmatrix}\sigma_{2p_z}^*$$

$$\text{Bond order} = \frac{10-8}{2} = 1$$

It is **diamagnetic.**

Bond dissociation energy order

$N_2 > N_2^+ = N_2^- > N_2^{2-}$

Bond length order

$N_2^{2-} > N_2^- = N_2^+ > N_2$

Bond dissociation energy order

$O_2^+ > O_2 > O_2^- > O_2^{2-}$

Bond length order

$O_2^{2-} > O_2^- > O_2 > O_2^+$

Boron Molecule (B_2)

No of electrons = 10

The electronic configuration is $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_x^1 = \pi 2p_y^1$

It has 2 paired electrons. Hence, **paramagnetic.**

Basic Idea of Metallic Bonding

In a metal crystal, all atoms are identical. So, these cannot be linked by electrovalent bonds as all have same E.N..

Hence, in metal crystal, the atoms are bonded with each other with a special type of bonding known as Metallic bonding.

Since the metals are electropositive, their ionisation evergies are low. So they readily lose their valence e^-s and convert into positive ions. These e^-s move from place to place through the empty valence orbitals and are shared simultaneously among all of them.

Essential Conditions for Metallic Bonding

The metal atoms should have low I.E.

There should be sufficient number of vacant orbitals in the valence shell.

Thus, the bonding which holds the metal atoms together on account of the force of attraction between metal ions & the mobile e^-s is called metallic bonding.

The strength of metallic bond is directly proportional to the no. of valence e^-s and the charge present on the nucleus. It provides an answer to the fact that alkali metals are soft and have low M.P. & B. P.. While Transition metals are hard and have high M.P. & B.P.

TRAIN YOUR BRAIN

Q. How many number of molecules have bond dissociation energy greater than that of I_2 ?

F_2, Cl_2, Br_2, N_2, O_2

(*a*) 2 (*b*) 3

(*c*) 4 (*d*) 5

Ans. (d) $Cl_2 > Br_2 > F_2 > I_2$

$N \equiv N$, $O = O$ there for B.D.E is higher than I_2

DIPOLE MOMENT

When a covalent bond is formed between two similar atoms (For ex H_2, F_2, O_2) the shared pair of electrons situated exactly between two atoms, the bond so formed is called non-polar covalent bond.

If a covalent bond is formed between two dissimilar atoms (ex H-F) the shared pair gets displaced more towards Fluorine since Fluorine is more E.N. than Hydrogen, such bond is called polar covalent bond.

Polarity of a polar molecule is expressed in terms of dipole moment (μ).

Dipole moment (μ)

= charge on the pole × distance between the poles.

Dipole moment is usually expressed in Debye units (D)

$1\ D = 3.33564 \times 10^{-30}$ Cm (S.I) where

C = coulomb and m = meter

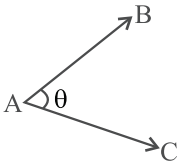
$= 10^{-18}$ esu. Cm (CGS)

Dipole moment is vector quantity and is represented by crossed arrow ⟷ . The cross is on positive end and arrow head is on negative end. For ex $\overset{+}{H}-\overset{}{F}$

In polyatomic molecules the dipole moment not only depends upon the individual dipole moments of bonds known as bond dipoles but also on the spatial arrangements of various bonds in the molecule. In such case, the dipole moment of a molecule is the vector sum of the bond dipoles of various bonds.

Calculation of Resultant Bond Moment

Let AB & AC are two polar bonds inclined at angle θ, their dipole moments are μ_1 & μ_2. The resultant dipole moment may be calculated using $\mu_R = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos\theta}$



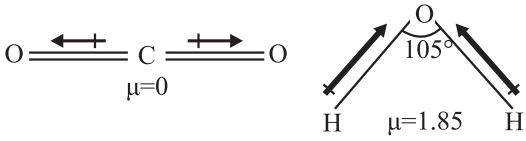
Type of Molecules	Example	Dipole Moment, μ(D)	Geometry
AB	HF	1.78	Linear
	HCl	1.07	Linear
	HBr	0.79	Linear
	HI	0.38	Linear
	H_2	0	Linear
AB_2	H_2O	1.85	Bent
	H_2S	0.95	Bent
	CO_2	0	Linear
AB_3	NH_3	1.47	Trigonal-pyramidal
	NF_3	0.23	Trigonal-pyramidal
	BF_3	0	Trigonal-planar
AB_4	CH_3Cl	1.87	Tetrahedral
	$CHCl_3$	1.04	Tetrahedral
	CCl_4	0	Tetrahedral

Applications of Dipole Moment

Dipole moment is helpful in predicting the geometry of the molecule.

Dipole moment helps in determining the polarity.

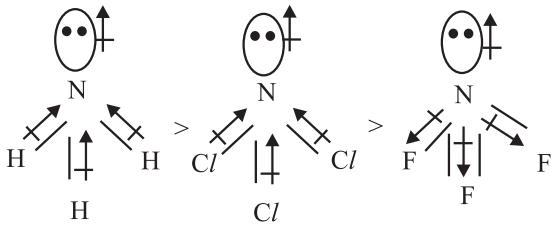
Dipole moment can distinguish between symmetrical and non symmetrical molecules e.g. CO_2 has 0 diploe moment as it is symmetrical whereas H_2O has a dipole moment of 1.85 D.



Net dipole moment of water = 1.85 D

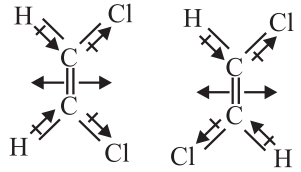
$= 1.85 \times 3.336 \times 10^{-30}$ Cm

$= 6.17 \times 10^{-30}$ Cm



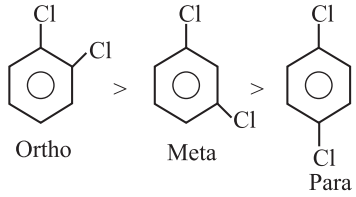
In $\ddot{N}F_3$ the bond pairs are moving opp. to lone pair and in NH_3 the bond pairs moving towards the lone pair.

Cis and trans isomers can be distinguished by dipole moments, usually cis isomers have higher dipole moment and hence higher polarity e.g.



Cis ($\mu \neq 0$) trans ($\mu = 0$)

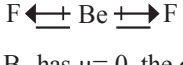
Dipole moment is greatest for ortho isomer, zero for para isomer which is less than that of ortho or meta isomer.



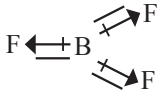
$\mu = 2.53$ $\mu = 1.48$ $\mu = 0$

Hybridisation can be determined by dipole moment for e.g.

(*i*) If a molecule AB_2 has $\mu = 0$, the σ orbitals used by A (Z < 21) must be sp hybridised e.g. BeF_2



(*ii*) If a molecule AB_3 has $\mu = 0$, the σ orbitals used by A (Z < 21) must be sp^2 hybridised e.g. BF_3



Note:

1. Dipole moment ∝ electronegativity difference
   HF > HCl > HBr > HI
2. Dipole moment ∝ No. of lone pair of electrons.
   $HF > H_2O > NH_3$
3. For homo atomic molecules and molecules having normal shapes such as linear, trigonal, tetrahedral posses zero dipole moment.
   Ex: Cl_2, N_2, O_2, CO_2
4. Molecules having distorted shapes like angular, pyramidal shows dipole moment
   Ex: H_2O, SO_2

Some important Points about Dipole Moment

A polyatomic molecule having polar covalent bonds but zero dipole moment indicates the symmetrical structure of the molecule. e.g. B–F bonds are polar in BF_3 but BF_3 has $\mu = 0$ due to its symmetrical geometry.

If molecule have $\mu = 0$, then it should be linear or having symmetrical geometry.

e.g. linear – CO_2, CS_2, $BeCl_2$ (g) ; symmetrical geometry – BF_3, CH_4 , PCl_5 , SF_6 , IF_7 , XeF_4.

If molecule has $\mu \neq 0$ then it should be angular or having unsymmetrical geometry.

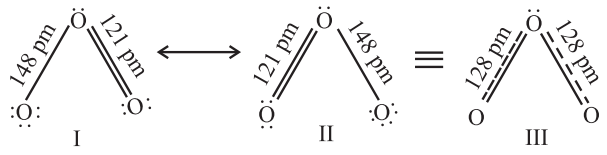
e.g. $SnCl_2$, $PbCl_2$, SO_2 – angular molecular geometry.

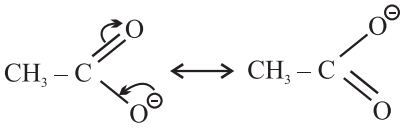
NH_3, H_2O, NF_3, SF_4, H_2S – unsymmetrical molecular geometry.

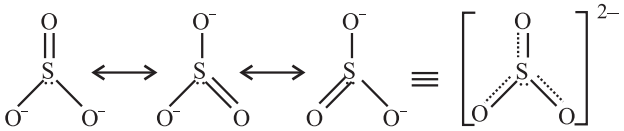
RESONANCE

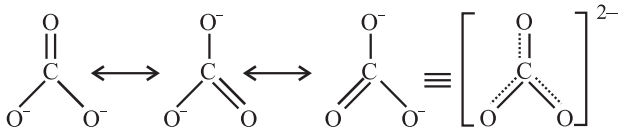
Representation of molecule in more than one structure and none of them explains all the properties of the molecule singly is called Resonance.

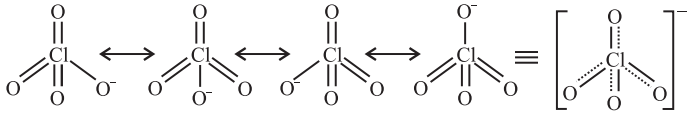
Example:





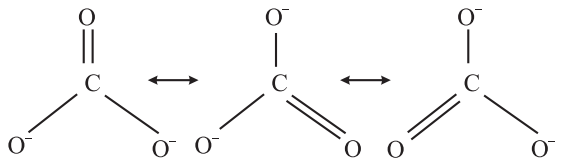






Bond order =

$$\frac{\text{Total No. of bonds formed between two atoms in all structures}}{\text{Total No. of resonating structures}}$$



$$\text{Bond order} = \frac{2+1+1}{3} = 1.33$$

✍ KEY NOTE

- Resonance energy = Actual bond energy – Energy of most stable resonating structure.
- Stability of molecule ∝ resonance energy.
- More is the number of covalent bonds in molecule more will be its resonance energy.
- Resonance energy ∝ number of resonating structures.

BOND LENGTH

The equilibrium distance between the nuclei of two bonded atoms in a molecule is called bond length.

Bond length is measured by spectroscopic, x-ray diffraction, and electron -diffraction techniques.

Covalent radius of an atom is equal to half the internuclear distance between two identical atoms that are held together by covalent bond.

Covalent radius of $A = \frac{d_{A-A}}{2}$

Covalent radius of an atom in heterodiatomic molecule (A-B type) when the E.N. difference between two atoms is large is calculated by using pauling empirical equation.

$$d_{A-B} = r_A + r_B + C(X_A - X_B)$$

Where X_A and X_B= E.N. of atoms A and B the value of C depends on the type of atoms involved in covalent bond.

For bonds involving atoms of II nd period, C = 0.09

For bonds between Si, P and S bonded to more E.N atoms not belonging to first period, C = 0.06.

The bond length between a set of atoms in different molecules is always same.

O-H bond length in H_2O, H_2O_2, C_2H_5OH is same i.e. 0.97Å

As the p-character of hybrid orbital changes the bond length changes.

C-H bond length in C_2H_6, C_2H_4 and C_2H_2 are 1.09Å, 1.08Å and 1.06 Å respectively.

As number of bonds between two atoms increases bond length decreases.

C – C;	C = C;	C ≡ C
1.54 Å;	1.33 Å;	1.20 Å

As the size of bonded atom increases bond length increases.

HF	HCl	HBr	HI
92pm	127pm	141pm	160pm

✍ KEY NOTE

It is effected by resonance and hyperconjugation.

BOND ENTHALPY

It is the amount of energy required to break one mole of bonds of a particular type between two atoms in a gaseous state.

The unit of bond enthalpy is kJ/mole

$$H_{2(g)} \rightarrow 2H_{(g)}, \Delta H^{\circ}_{d} = 435.8 \text{ kJ/mole}$$

Heterolytic fission requires more bond energy than homolytic fission.

The number of bonds between two atoms increases as bond enthalpy increases.

C-C	C=C	C ≡ C
341	610	828 kJ/mole

As the number of lone pairs on a bonded atom increases, the bond enthalpy decreases.

$$-C-C- > -\ddot{N}-\ddot{N}- > -\ddot{O}-\ddot{O}-$$

In case of H_2O molecule, the enthalpy needed to break two O-H bonds is not the same due to changed chemical environment. Therefore mean or average bond enthalpy is used.

$$H_2O_{(g)} \rightarrow H^{+}_{(g)} + {}^{-}OH_{(g)}, \ \Delta_a H^{\Theta}_1 = 502 \text{ kJ mol}^{-1}$$

$${}^{-}OH_{(g)} \rightarrow H^{+}_{(g)} + O^{2-}_{(g)}, \Delta_a H^{\Theta}_2 = 427 \text{ kJ mol}^{-1}$$

Average bond enthalpy of O-H bond in water

$$= \frac{502 + 427}{2} = 464.5 \text{ kJ}$$

✍ KEY NOTE

- The magnitude of bond energy depends upon size of the atoms forming the bond.
- Shorter the bond length higher the bond energy. Bond energy values are less for homolytic fission when compared to heterolytic fission.

BOND ANGLE

- Angle between the orbitals containing bonding pair of electrons around the central atom in a molecule or complex ion is called bond angle.
- Bond angle is determined by spectroscopic methods.
- As bond angle increases stability increases.

Bond angle is effected by the presence of lone pair of electrons on central atom

Ex: $CH_4 \rightarrow 109°28' \rightarrow [0 \ \text{L.P}]$

$NH_3 \rightarrow 107° \rightarrow [1 \ \text{L.P}]$

$H_2O \rightarrow 104.5° \rightarrow [2 \ \text{L.P}]$

Bond angle depends on state of hybridisation.

If the electronegativity of central atom decreases bond angle decreases.

In case, the central atom remains same bond angle increases with the decrease in electronegativity of surrounding atoms.

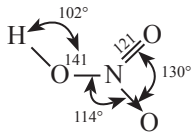
% s-character increases bond angle increases.

TRAIN YOUR BRAIN

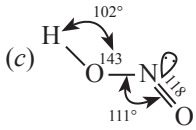
Q. The ONO angle is maximum in :

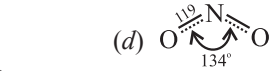
(*a*) HNO_3 (*b*) NO_2^+

(*c*) HNO_2 (*d*) NO_2

Ans. (b) (*a*) 

(*b*) $\underset{115pm}{O} = \overset{+}{N} = O$ Bond angle is 180º because of sp hybridisation of nitrogen.

(*c*) 

(*d*) 

BOND ORDER

- Bond order indicates number of bonds between two atoms in a molecule.
- Greater the bond order more is the strength of the bond.
- With increase in bond order, bond enthalpy increases and bond length decreases.
- Bond order 3 is being one of the highest for a diatomic molecule.
- Isoelectronic species have same bond order.

  Eg,. N_2, CO,NO^+
- It is observed that with the increase in bond order, bond energy increases and bond length decreases.

HYDROGEN BOND

Nitrogen, oxygen and fluorine are the highly electronegative elements. When they are tied to a hydrogen atom to form covalent bond, the electrons of the covalent bond are shifted towards the more electronegative atom. This partially positive charged hydrogen atom forms a bond with the other more electronegative atom. This bond is called as **hydrogen bond** and is weaker than covalent bond.

Hydrogen bonding is said to be formed when slightly acidic hydrogen attached to a strongly electronegative atom such as F, N and O is held with weak electrostatic forces by the lone pair of electrons of the electronegative atom.

Hydrogen bond is represented by a dotted line (– – –)

For example, in HF molecule, the hydrogen bond exists between hydrogen atom of one molecule and fluorine atom of another molecule as given below :

$$- - - H^{\delta+} - F^{\delta-} - - - - H^{\delta+} - F^{\delta-} - - - - H^{\delta+} - F^{\delta-}$$

Hydrogen bond is weaker than covalent bond and stronger than van der waal force of attraction.

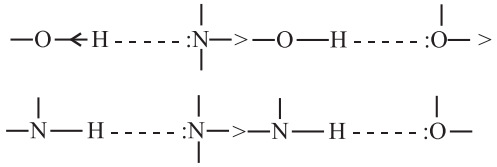
Hydrogen bond present in the same molecule is known as **intra molecular hydrogen bond.**

Due to intramolecular hydrogen bonding, ring formation or chelation occurs.

Substances having intramolecular hydrogen bonds are less water soluble and steam volatile and have low boiling points.

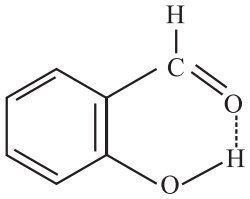
Hydrogen bond formed between different polar molecules is known as **inter molecular hydrogen bond.**

Order of H-bond strength



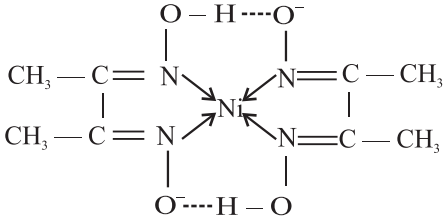
Types of Hydrogen Bonds

(A) Intramolecular H-Bonding: This type of H-bonding occurs when polar H and electronegative atom are present in the same molecule.



o-hydroxy benzaldehyde

It has lower boiling point (i.e. more volatile) than its para-derivative (where association of molecules takes place using intermolecular H-bonding) because it exists as discrete molecules.



Extra stability of the complex is because of intramolecular hydrogen bonding in addition to the chelating effect.

Necessary conditions for the formation of intramolecular hydrogen-bonding

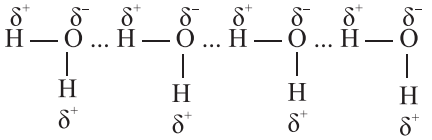
(*a*) The ring formed as a result of hydrogen bonding should be planar.

(*b*) 5- or 6- membered ring should be formed.

(*c*) Interacting atoms should be placed in such a way that there is minimum strain during the ring closure.

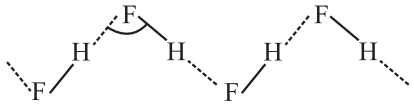
(B) Intermolecular H-Bonding: It is formed between two different molecules of the same or different compounds.

In water molecules: Due to polar nature of H_2O there is association of water molecules giving a liquid state of abnormally high boiling point.

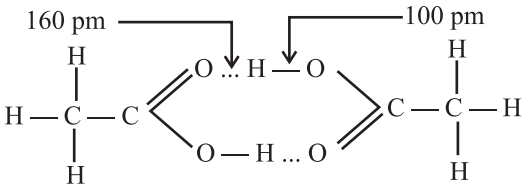


There is also similar H-bonding in alcohol (R—OH) ammonia (NH_3) and phenol (C_6H_5OH) molecules.

The hydrogen bonds in HF link the F atom of one molecule with the H-atom of another molecule, thus forming a zig-zag chain $(HF)_n$ in both the solid and also in the liquid.



Carboxylic acid dimerises in gaseous state due to H-bonding



Alcohol is said to be highly soluble in water due to crossed intermolecular H-bonding (between H_2O and R—OH molecules).

However, isomeric ether is less soluble in water due to less polar nature of ether.

HCO_3^- ions exist as dimer in $KHCO_3$

In $Na^+HCO_3^-$, the HCO_3^- ions are linked to an infinite chain through intermolecular H-bonding.

✍ KEY NOTE

- The strongest H-bonds are formed by F atoms. Deuterium is more electropositive than H, therefore it also forms stronger bonds.

Topicwise Questions

KOSSEL LEWIS THEORY & LEWIS DOT STRUCTURE

1. Metal 'M' forms a peroxide of the type MO_2. Valency of the metal with respect to oxygen
(*a*) 0 (*b*) 1 (*c*) 2 (*d*) 4

2. Valency of the metal atom with respect to oxygen is maximum in
(*a*) Mn_2O_7 (*b*) OsO_4
(*c*) MnO_2 (*d*) CrO_3

3. Variable valency is a property of
(*a*) Alkali metals (*b*) Transition metals
(*c*) Alkaline earth metals (*d*) Inert gases

4. The molecule that deviates from octet rule is
(*a*) NaCl (*b*) $BeCl_2$
(*c*) MgO (*d*) NH_3

5. Which of the following molecule deviates from octet rule with respect to central atom?
(*a*) PCl_3 (*b*) H_2S (*c*) NH_3 (*d*) XeF_4

6. Expanded octet can be observed in the valency shell of the central atom in
(*a*) NH_3 (*b*) CH_4 (*c*) PCl_5 (*d*) $BeCl_2$

7. How many number of species are hypovalent ?
$BeCl_2$, BCl_3, CCl_4, PCl_3, O_2, H_2
(*a*) 2 (*b*) 3 (*c*) 4 (*d*) 5

8. Which of the following species will be the strongest Lewis acid ?
(*a*) Fe^0 (*b*) Fe^{3+} (*c*) Fe^{2+} (*d*) Fe^+

9. Which of following molecule/specie is having maximum number of lone pairs in Lewis - dot structure ?
(*a*) BH_4^- (*b*) BF_4^-
(*c*) CN^- (*d*) $COCl_2$

10. Which of the following species follows the octet rule for all of its atoms ?
(*a*) BeH_2-dimer (*b*) TeF_6
(*c*) XeF_6 (*d*) C_2F_6

11. What is the formal charge on O-atom in carbonate ion.
(*a*) –1 (*b*) 0
(*c*) +1 (*d*) – 1 & 0 both

IONIC OR ELECTROVALENT/COVALENT BOND

12. Molecule having maximum number of covalent bonds is
(*a*) NH_4OH (*b*) NH_4Cl
(*c*) $CO(NH_2)_2$ (*d*) CH_3OH

13. Number of bonded electrons in ethane molecule are
(*a*) 7 (*b*) 12
(*c*) 10 (*d*) 14

14. In an ionic compound A^+X^-, the degree of covalent bonding is greater when
(*a*) A^+ and X^- ions are small
(*b*) A^+ is small and X^- is large
(*c*) A^+ and X^- ions are approximately of the same size
(*d*) X^- is small and A^+ is large

15. Which of the following has highest ionic character?
(*a*) $MgCl_2$ (*b*) $CaCl_2$
(*c*) $BaCl_2$ (*d*) $BeCl_2$

16. In which of the following molecule, the central atom has three lone pairs of electrons ?
(*a*) Ammonia (*b*) Xenon difluoride
(*c*) Chlorine trifluoride (*d*) Hydrogen sulphide

17. Number of lone pairs of electrons in 9 gms. of water are [N = Avogadro Number]
(*a*) 2N (*b*) N / 2
(*c*) N (*d*) N / 4

18. The number of electron pairs involved in the formation of hydrogen cyanide molecule are
(*a*) Two (*b*) Eight
(*c*) Three (*d*) Four

19. The number of π-bonds in the structure given below are: $(NC)_2\ C = C\ (CN)_2$
(*a*) 9 (*b*) 7 (*c*) 5 (*d*) 2

20. The molecule ML_x is planar with six pair of electrons around M in the valence shell. The value of x is
(*a*) 6 (*b*) 2 (*c*) 4 (*d*) 3

21. The electronegativities of two elements are 0.7 and 3.0. The bond formed between them would be
(*a*) Ionic (*b*) Covalent
(*c*) Co-ordinate covalent (*d*) Metallic

22. Highly ionic compound is formed by the combination of elements belonging to
(*a*) I A group and VII A gruop
(*b*) II A group VI A group
(*c*) III A group and V A group
(*d*) 'O' group and VII A group

23. An electrovalent compound is made up of
(*a*) Electrically charged particles
(*b*) Neutral molecules
(*c*) Neutral atoms
(*d*) Electrically charged atom or group of atoms

24. Potassium forms a highly ionic compound when it combines with
(*a*) Chlorine (*b*) Fluorine
(*c*) Bromine (*d*) Iodine

25. The number of sp^2- sigma bonds in benzene are:
(*a*) 3 (*b*) 6
(*c*) 12 (*d*) None of these

26. The strength of the bonds 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

27. Most ionic compound among the following is
(*a*) Sodium fluoride (*b*) Sodium Chloride
(*c*) Sodium bromide (*d*) Sodium iodide

28. Which of the following has least polarity in bond?
(*a*) H – F (*b*) H – Cl
(*c*) H – O (*d*) H – S

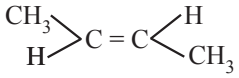
29. Which contains both polar and non-polar bonds ?
(*a*) NH_4Cl (*b*) HCN
(*c*) H_2O_2 (*d*) CH_4

30. Which of the following is non-polar?
(*a*) H_2S (*b*) NaCl
(*c*) Cl_2 (*d*) H_2SO_4

31. Which of the following unit conversion of dipole moment is correct?
(*a*) 1D = 3.3356 Cm
(*b*) 1D= 3.3365×10^{-30} Cm
(*c*) 1 C m = 3.3356 D
(*d*) 1 C m = 3.3356×10^{-30} D

32. The compound which does not contain ionic bond is:
(*a*) NaOH (*b*) HCl
(*c*) K_2S (*d*) LiH

33. Which of the following chloride has considerable covalent character?
(*a*) LiCl (*b*) NaCl
(*c*) KCl (*d*) CsCl

34. Which of the following hydrocarbons has the lowest dipole moment?
(*a*)  (*b*) $CH_3C \equiv CCH_3$
(*c*) $CH_3CH_2C \equiv CH$ (*d*) $CH_2 = CH - C \equiv CH$

BOND PARAMETERS

35. Which of the following is correct order of bond length ?
(*a*) $BF_4^- < BF_3$ (*b*) $NO_2^+ < NO_2^-$
(*c*) $CCl_4 < CF_4$ (*d*) $^+CH_3 > CH_4$

36. Indicate the nature of bonding in diamond.
(*a*) Ionic (*b*) Covalent
(*c*) Molecular (*d*) Metallic

37. In which of the following species, the angle around the central atom is exactly equal to 109°28′?
(*a*) SF_4 (*b*) NH_3
(*c*) NH_4^+ (*d*) None of these

38. Which of the following species has bond angle less than 98°?
(*a*) H_2O (*b*) NH_3
(*c*) CH_4 (*d*) PH_3

39. Find out the relation between (adjacent angle) $\angle\widehat{FClF}$ and $\angle\widehat{FBrF}$ bond angle in ClF_3 and BrF_3 molecule respectively.
(*a*) $\widehat{FClF} = \widehat{FBrF}$ (*b*) $\widehat{FClF} > \widehat{FBrF}$
(*c*) $\widehat{FClF} < \widehat{FBrF}$ (*d*) Can't predicted

40. The number of P–P–P angle in P_4 structure is:
(*a*) 6 (*b*) 9
(*c*) 12 (*d*) 10

41. The bond angle and hybridization in ether (CH_3OCH_3) is :
(*a*) 106°51', sp^3 (*b*) 104°31', sp^3
(*c*) > 109° 28' sp^3 (*d*) None of these

42. The bond angle in H_2S is:
(*a*) > NH_3 (*b*) Same as in $BeCl_2$
(*c*) > H_2Se, < H_2O (*d*) Same as in CH_4

43. The smallest bond angle around the central atom will be present in:
(*a*) H_2O (*b*) BeF_2
(*c*) CH_4 (*d*) NH_3

44. The bond length of the S–O bond is maximum in which of the following compound ?
$SOBr_2$, $SOCl_2$, SOF_2
(*a*) $SOCl_2$ (*b*) $SOBr_2$
(*c*) SOF_2 (*d*) All have same length

45. Which of the angle is most influenced by the presence of lone pair in SF_4 ?
(*a*) Equatorial angle
(*b*) Axial angle
(*c*) Angle between axial and equatorial
(*d*) None

46. Correct order of bond angle in CH_2F_2:
(*a*) $H\hat{C}H > H\hat{C}F > F\hat{C}F$ (*b*) $F\hat{C}F > H\hat{C}F > H\hat{C}H$
(*c*) $H\hat{C}H > F\hat{C}F > H\hat{C}F$ (*d*) $H\hat{C}F > F\hat{C}F > H\hat{C}H$

47. Which of the following molecule has shortest C-H bond ?
(*a*) Ethene (*b*) Ethane
(*c*) Ethyne (*d*) Methane

48. The correct sequence of decrease in the bond angles of the following hydrides is:
(*a*) $NH_3 > PH_3 > AsH_3 > SbH_3$
(*b*) $NH_3 > AsH_3 > PH_3 > SbH_3$
(*c*) $SbH_3 > AsH_3 > PH_3 > NH_3$
(*d*) $PH_3 > NH_3 > AsH_3 > SbH_3$

49. Find the maximum number of identical angle in CH_2F_2 molecule.
(*a*) 2 (*b*) 4
(*c*) 6 (*d*) 8

50. In which of the following bond angle is maximum:
(*a*) NH_3 (*b*) NH_4^+
(*c*) PCl_3 (*d*) SCl_2

51. Which of the following H-bonds is expected to have maximum strength?
(*a*) H-O----H (*b*) H-N----H
(*c*) H-S-----H (*d*) All have same strength

52. Which one of the following compounds has the smallest bond angle in its molecule?
(*a*) OH_2 (*b*) SH_2 (*c*) NH_3 (*d*) SO_2

53. In O_3, there are:
(*a*) 2σ, 1π bond (*b*) 1σ, 2π bonds
(*c*) 2σ, 2π bonds (*d*) 2σ, 1π bond

54. Which one of the following has the smallest bond angle?
(*a*) NH_3 (*b*) BeF_2
(*c*) H_2O (*d*) CH_4

55. The correct order of bond angles (smallest first) in , H_2S, NH_3, BF_3 and SiH_4 is:
(*a*) $H_2S < NH_3 < SiH_4 < BF_3$
(*b*) $NH_3 < H_2S < SiH_4 < BF_3$
(*c*) $H_2S < SiH_4 < NH_3 < BF_3$
(*d*) $H_2S < NH_3 < BF_3 < SiH_4$

56. Which one of the following bonds has the highest average bond energy?
(*a*) S = O (*b*) C ≡ C
(*c*) C ≡ N (*d*) N ≡ N

57. The correct decreasing order of bond angles is:
(*a*) $ClF_3 > PF_3 > NF_3 > BF_3$
(*b*) $BF_3 > PF_3 > NF_3 > ClF_3$
(*c*) $BF_3 > NF_3 > PF_3 > ClF_3$
(*d*) $BF_3 > ClF_3 > PF_3 > NF_3$

58. Bond angle in water is:
(*a*) 120° (*b*) 109.5°
(*c*) 107° (*d*) 104.5°

59. The correct order in which the O–O bond length increases in the following :
(*a*) $H_2O_2 < O_2 < O_3$ (*b*) $O_3 < H_2O_2 < O_2$
(*c*) $O_2 < H_2O_2 < O_3$ (*d*) $O_2 < O_3 < H_2O_2$

VSEPR THEORY

60. Cyanogen, $(CN)_2$, has a ___ shape/structure:
(*a*) Linear (*b*) Zig-zag
(*c*) V-shape (*d*) Cyclic

61. The shape of covalent molecule AX_3 is:
(*a*) Triangular (*b*) T-shape
(*c*) Pyramidal (*d*) Any of the above

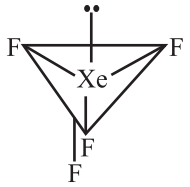
62. Of the three molecules XeF_4, SF_4, SiF_4, one which have tetrahedral structure is ?
(*a*) All the three (*b*) Only SiF_4
(*c*) Both SF_4 & XeF_4 (*d*) Only SF_4 & XeF_4

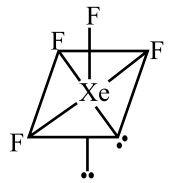
63. The species/molecule is having same shape and different hybridisation:
(*a*) XeF_2, CO_2 (*b*) I_3^-, H_3O^+
(*c*) OCl^-, CO (*d*) SO_2, OCl_2

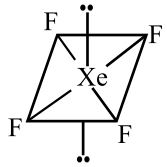
64. Which of the following tetra atomic species is non-polar and planar?
(*a*) BrF_4^- (*b*) NH_4^+
(*c*) ClF_3 (*d*) BF_3

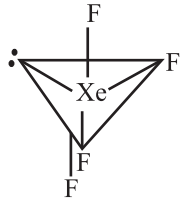
65. Which of the possible molecule / species is having maximum values for dipole moment. (where "A" is the central atom)?
(*a*) AX_3 (having one lone pair on central atom)
(*b*) AX_4 (Tetrahedral)
(*c*) AX_4Y (having no lone pair on central atom)
(*d*) Can't be predicted

66. Which is the right structure of XeF_4 ?

(*a*) 

(*b*) 

(*c*) 

(*d*) 

67. The shape of SO_4^{2-} ion is:
(*a*) Square planar (*b*) Square Pyramidal
(*c*) Tetrahedral (*d*) None of these

68. Which of the following species given below have shape similar to $XeOF_4$?
(*a*) XeO_3 (*b*) IOF_4^+
(*c*) PCl_5 (*d*) XeF_5^+

69. The total number of lone pairs of electrons in NO_3^- :
(*a*) 1 (*b*) 2
(*c*) 6 (*d*) 8

70. The central atom in IF_7 has seven pairs of valency electrons. The shape of the molecule as per VSEPR theory is:
(*a*) Trigonal bipyramid
(*b*) Hexagonal pyramid
(*c*) Pentagonal bipyramid
(*d*) Square bipyramid

71. The correct order of decreasing polarity is:
(*a*) $HF > SO_2 > H_2O > NH_3$
(*b*) $HF > H_2O > SO_2 > NH_3$
(*c*) $HF > NH_3 > SO_2 > H_2O$
(*d*) $H_2O > NH_3 > SO_2 > HF$

72. Which of the following is a pyramidal species?
(*a*) CO_2 (*b*) H_3O^+
(*c*) $SnCl_2$ (*d*) NH_4^+

73. In CCl_4 the four valencies of carbon are directed towards the corners of a
(*a*) Cube (*b*) Hexagon
(*c*) Prism (*d*) Tetrahedron

74. The shape of CO_2 molecule is similar to:
(*a*) H_2O (*b*) BeF_2
(*c*) SO_2 (*d*) None of these

75. Out of $CHCl_3$, CH_4 and SF_4, the molecules having regular geometry are:
(*a*) $CHCl_3$ only (*b*) $CHCl_3$ and SF_4
(*c*) CH_4 only (*d*) CH_4 and SF_4

76. In OF_2 molecule, the total number of bond pairs and lone pairs of electrons are respectively:
(*a*) 2, 6 (*b*) 2, 8
(*c*) 2, 10 (*d*) 2, 9

77. Which one of the following has the regular tetrahedral structure?
(*a*) XeF_4 (*b*) $[Ni(CN)_4]^{2-}$
(*c*) BF^-_4 (*d*) SF_4

78. The shape of formaldehyde molecule as per the VSEPR theory is:
(*a*) Linear (*b*) Planar triangle
(*c*) Pyramid (*d*) Tetrahedron

79. Hybrid orbital with least s - character is:
(*a*) sp^3d (*b*) sp^2
(*c*) sp^3 (*d*) sp

80. The number of hybrid orbitals in a molecule of decane are:
(*a*) 36 (*b*) 40
(*c*) 38 (*d*) 8

81. The molecule which contains $\sigma_{sp^3-sp^3}$ and σ_{sp^3-p} bonds in it is:
(*a*) CH_3CH_3
(*b*) CH_3CHO
(*c*) CH_3CH_2Cl
(*d*) $CHCl_3$

82. Bond angle between two hybrid orbitals is 107^o, s - orbital character of hybrid orbital is nearly
(*a*) 50 % (*b*) 33.33%
(*c*) 16.6 % (*d*) 25 %

83. Carbon atoms in $C_2(CN)_4$ are:
(*a*) sp hybridized
(*b*) sp^2 Hybridized
(*c*) sp and sp^2 hybridized
(*d*) sp,sp^2 and sp^3 hybridized

84. The interparticle forces in liquid hydrogen are.
(*a*) H-bonds
(*b*) Vander waal's forces
(*c*) Covalent bonds
(*d*) None of these

85. The number of sigma and Pi bonds in a molecule of cyanogen are:
(*a*) 4, 3 (*b*) 3, 4
(*c*) 5, 2 (*d*) 3, 5

VALENCE BOND THEORY & HYBRIDISATION

86. Which of the following molecule is having shortest bond length of C–O bond ?
(*a*) CH_3OH (*b*) H_2CO
(*c*) CO (*d*) Na_2CO_3

87. The ratio of σ and π bonds in benzene is:
(*a*) 2 (*b*) 6
(*c*) 4 (*d*) 8

88. A hybrid orbital formed from s and p- orbital can contribute to
(*a*) σ- bond only (*b*) π-bond only
(*c*) Either σ or π-bond (*d*) Cannot be predicted

89. The hybrid state of C in CS_2 should be:
(*a*) sp^2 (*b*) sp
(*c*) sp^3 (*d*) Not specific

90. Indicate the wrong statement according to Valence bond theory:
(*a*) A sigma bond is stronger than π-bond
(*b*) p-orbitals always have only sidewise overlapping
(*c*) s-orbitals never form π-bonds
(*d*) There can be only one sigma bond between two atoms

91. The strongest P–O bond is found in the molecule
(*a*) F_3PO (*b*) Cl_3PO
(*c*) Br_3PO (*d*) $(CH_3)_3PO$

92. In pent-3-en-1-yne the terminal carbon-atoms have following hybridisation:
(*a*) sp & sp^2 (*b*) sp^2 & sp^3
(*c*) sp^2 & sp (*d*) sp & sp^3

93. During the complete combustion of methane CH_4, what change in hybridisation does the carbon atom undergo?
(*a*) sp^3 to sp (*b*) sp^3 to sp^2
(*c*) sp^2 to sp (*d*) sp^2 to sp^3

94. Which orbital is used by oxygen atom to form a sigma bond with other oxygen atom in O_2 molecule?
(*a*) Pure p-orbital (*b*) sp^2 - hybrid orbital
(*c*) sp^3 - hybrid orbital (*d*) sp - hybrid orbital

95. Which of the following compounds have the same no. of lone pairs with their central atom ?
[I] XeF^-_5 [II] BrF_3 [III] XeF_2
[IV] H_3S^+ [V] Triplet Methylene
(*a*) IV and V (*b*) I and III
(*c*) I and II (*d*) II, IV and V

96. The correct order of hybridization of the central atom in the following species :
NH_3, XeO_2F_2, SeF_4, NO_2^+
(*a*) sp^3, sp^3, sp^3d, sp (*b*) sp^3, sp^3d, sp^3d, sp
(*c*) sp^3, sp^3d^2, sp^3d, sp^2 (*d*) sp^2, sp^3d, sp^3d^2, sp

97. The hybridisation of the central atom will change when
(*a*) NH_3 combines with H^+
(*b*) H_3BO_3 combines with OH^-
(*c*) NH_3 forms NH^-_2
(*d*) H_2O combines with H^+

98. What is the state of hybridisation of anionic part of solid N_2O_5 ?
(*a*) sp (*b*) sp^2
(*c*) sp^3 (*d*) Not applicable

99. Which one of the following pairs is isostructural?
(*a*) [BCl_3 and $BrCl_3$] (*b*) [NH_3 and NO_3^-]
(*c*) [NF_3 and BF_3] (*d*) [BF_4^- and NH_4^+]

100. The hybridisation of C-atoms in tetracyanomethane is
(*a*) sp, sp^2 (*b*) sp^3, sp
(*c*) sp^3, sp^3 (*d*) sp^3 , sp^2

101. $BF_3 + F^- \rightarrow BF_4^-$
What is the hybridiation state of B in BF_3 and BF_4^- ?
(*a*) sp^2, sp^3 (*b*) sp^3, sp^3
(*c*) sp^2, sp^2 (*d*) sp^3, sp^3d

102. Choose the molecules in which hybridisation occurs in the ground state ?
(*i*) BCl_3 (*ii*) NH_3
(*iii*)PCl_3 (*iv*) BeF_2
The correct answer is
(*a*) i, ii, iv (*b*) i, ii, iii
(*c*) ii, iii (*d*) iii, iv

103. The geometry of electron pairs around I in IF_5 is:
(*a*) Octahedral (*b*) Trigonal Bipyramidal
(*c*) Square Pyramidal (*d*) Pentagonal Planar

104. sp^3d hybridization is considered to be a combination of two hybridization. They are
(*a*) $p^3 + sd$ (*b*) $sp^2 + pd$
(*c*) $spd + p^2$ (*d*) None of these

105. All the four sigma bonds in perchlorate ion are formed by which orbital?
(*a*) sp^3d^3 (*b*) sp^3d^2
(*c*) sp^3d (*d*) sp^3

106. Which one of the following conversions involve change in both hybridisation and shape?
(*a*) $CH_4 \rightarrow C_2H_6$ (*b*) $NH_3 \rightarrow NH^+_4$
(*c*) $BF_3 \rightarrow BF_4^-$ (*d*) $H_2O \rightarrow H_3O^+$

107. Which of the following penta atomic species is tetrahedral and non polar ?
(*a*) CH_3F (*b*) $\overset{+}{N}H_2F_2$
(*c*) $[Ni(CO)_4]$ (*d*) BF_4^-

108. The orbital overlapping is maximum in
(*a*) Cl_2 (*b*) HI
(*c*) HCl (*d*) HBr

109. Linear combination of two hybridised orbitals belonging to two atoms and each having one electron leads to a
(*a*) Sigma bond
(*b*) Double bond
(*c*) Co-ordinate covalent bond
(*d*) pi bond

110. pi bond is formed by the overlapping of
(*a*) p - p orbitals along their axis
(*b*) s - p orbitals along the axis of p - orbital
(*c*) p - p orbitals perpendicular to their axis
(*d*) s - s orbitals

111. Which of the following is not correct ?
(*a*) A sigma bond is weaker than pi bond
(*b*) A sigma bond is stronger than pi bond
(*c*) A double bond is stronger than a single bond
(*d*) A double bond between two atoms is shorter than a single bond between the same atoms.

112. The s - character in the hybrid orbital of the central atom, present in a molecule having the shape of an octahedron is
(*a*) 25 % (*b*) 75 %
(*c*) 40 % (*d*) 16.66 %

113. The C-H bond in propane is
(*a*) $\sigma_{sp\text{ - }s}$ (*b*) $\sigma_{sp^2\text{–}s}$ (*c*) $\sigma_{p\text{ - }s}$ (*d*) $\sigma_{sp^3\text{–}s}$

114. The hybridisation of P in phosphate ion (PO_4^{3-}) is the same as:
(*a*) I in ICl_4^- (*b*) S in SO_3
(*c*) N in NO_3^- (*d*) S in SO_3^{2-}

115. Number of hybrid orbitals present in a molecule of propene are
(*a*) 12 (*b*) 10 (*c*) 9 (*d*) 8

116. Molecule obtained by sp^3d^2 hybridisation has bond angle of
(*a*) 90^o (*b*) $109^o28'$
(*c*) 120^o (*d*) 72^o

117. Hybrid orbital having maximum p - character is
(*a*) sp^3d (*b*) sp^3
(*c*) sp (*d*) sp^3d^2

118. Hybridisation involves
(*a*) Addition of an electron pair
(*b*) Combination and redistribution of atomic orbitals
(*c*) Removal of an electron pair
(*d*) Separation of orbitals

119. On hybridisation of one ‘s’ and one ‘p’ orbitals we get
(*a*) Two mutually perpendicular orbitals
(*b*) Two orbitals at 180^o
(*c*) Four orbitals directed tetrahedrally
(*d*) Three orbitals in a plane

120. As the s - character in a hybrid orbital increases, the bond angle
(*a*) Increases (*b*) Decreases
(*c*) Does not change (*d*) Becomes zero

121. sp^3d hybridisation results in
(*a*) A square planar molecule
(*b*) An octahedron molecule
(*c*) A trigonal bipyramid molecule
(*d*) A tetrahedron molecule

122. s - p overlapping is present in

(*a*) Br_2 (*b*) H_2

(*c*) O_2 (*d*) HF

123. Valence bond theory of Pauling and Slater accounts for the following characteristic of covalent bond

(*a*) Directional (*b*) ionic

(*c*) Strength (*d*) hybrid

124. Shape of molecule is decided by

(*a*) Sigma bonds (*b*) pi bonds

(*c*) Both sigma and pi bonds (*d*) Neither sigma nor pi bonds

125. The strength of bonds by 2s-2s, 2p-2p, 2s-2p, 1s-1s overlap has the order:

(*a*) 1s-1s > 2p-2p > 2s-2p > 2s-2s

(*b*) 2p-2p > 2s-2p > 2s-2s > 1s-1s

(*c*) 2s-2s > 1s-1s > 2s-2p > 2p-2p

(*d*) 2s-2p > 2s-2s > 2p-2p > 1s-1s

126. The hybridization of atomic orbitals of nitrogen in NO_2^+ , NO_3^- and NH_4^+ are

(*a*) sp^2, sp^3 and sp^2 respectively

(*b*) sp, sp^2 and sp^3 respectively

(*c*) sp^2, sp and sp^3 respectively

(*d*) sp^2, sp^3 and sp respectively

127. According to valence bond theory the predicted bond angle for H_2O

(*a*) 90° (*b*) 109°28′

(*c*) $107^\circ 18'$ (*d*) $104^\circ 28'$

128. Which one of the following is a correct set?

(*a*) H_2O,sp^3,angular (*b*) H_2O,sp^2, linear

(*c*) NH_4^+,dsp^2, square planar (*d*) CH_4,dsp^2, tetrahedral

129. Which one of the following is the correct set with reference to molecular formula, hybridisation of central atom and shape of the molecule?

(*a*) CO_2,sp^2, bent (*b*) H_2O, sp^2, bent

(*c*) $BeCl_2$, sp, linear (*d*) H_2O, sp^3, linear

130. What is the hybridisation state of the central atom in the conjugate base of NH_4^+ ion?

(*a*) sp (*b*) sp^3

(*c*) sp^2 (*d*) dsp^2

131. For which hybridization, there are two unequal bond angles

(*a*) sp^3 (*b*) sp^2

(*c*) sp (*d*) sp^3d

132. The shape of CH_4; SO_4^{-2}; PO_4^{-3} is

(*a*) Trigonal planar

(*c*) Angular

(*c*) Tetrahedral

(*d*) Trigonal Bipyramidal

133. The compound in which carbon uses sp^3 hybridisation for bond formation

(*a*) H-COOH (*b*) $(NH_2)_2C{=}O$

(*c*) H-CHO (*d*) CH_3CH_2OH

134. On catalytic hydrogenation, ethylene gives ethane during this reaction

(*a*) Hybridization of carbon atoms changes from sp^2 to sp^3

(*b*) bond angle decreases from120^0 to 109.5^0

(*c*) C-C bond length increases from1.34 A^0 to 1.54A^0

(*d*) All of these.

135. The ratio of pure orbitals to hybridized orbitals in ethylene is

(*a*) 2 : 3 (*b*) 3 : 1

(*c*) 1 : 1 (*d*) 1 : 3

MOLECULAR ORBITAL THEORY

136. Which of the following species has lowest ionization potential?

(*a*) O_2^{2-} (*b*) O_2

(*c*) O_2^+ (*d*) O_2^-

137. Among KO_2, AlO_2^- , BaO_2 & NO_2^+ unpaired electron is present in:

(*a*) NO_2^+ and BaO_2 (*b*) KO_2 and AlO_2^-

(*c*) KO_2 only (*d*) BaO_2 only

138. What is correct sequence of bond order ?

(*a*) $O_2^+ > O_2^- > O_2$ (*b*) $O_2^+ > O_2 > O_2^-$

(*c*) $O_2 > O_2^- > O_2^+$ (*d*) $O_2^- > O_2^+ > O_2$

139. Which statement is correct about O_2^+ ?

(*a*) Paramagnetic and bond order < O_2

(*b*) Paramagnetic and bond order > O_2

(*c*) Diamagnetic and bond order < O_2

(*d*) Diamagnetic and bond order > O_2

140. In which of the following ionization processes, the bond order has increased and the magnetic behaviour has changed?

(*a*) $C_2 \rightarrow C_2^+$ (*b*) $NO \rightarrow NO^+$

(*c*) $O_2 \rightarrow O_2^+$ (*d*) $N_2 \rightarrow N_2^+$

141. The molecular electronic configuration of H_2^- ion is

(*a*) $(\sigma 1s)^2$ (*b*) $(\sigma 1s)^2(\sigma^* 1s)^2$

(*c*) $(\sigma 1s)^2(\sigma^* 1s)^1$ (*d*) $(\sigma 1s)^3$

142. Which one of the following molecules contains no π-bond?

(*a*) CO_2 (*b*) H_2O

(*c*) SO_2 (*d*) NO_2

143. Which of the following molecules/ions does not contain unpaired electrons?

(*a*) O_2^{2-} (*b*) B_2

(*c*) N_2^- (*d*) O_2

144. Anti- bonding molecular orbital is formed by

(*a*) Addition of wave functions of atomic orbitals

(*b*) Substraction of wave functions of atomic orbitals

(*c*) Multiplication of wave functions of atomic orbitals

(*d*) Finding the arithmetic mean

145. Which one of the following species does not exist under normal conditions?

(*a*) Li_2 (*b*) Be_2^+

(*c*) Be_2 (*d*) B_2

146. Oxygen molecule is
(*a*) Diamagnetic with no unpaired electrons
(*b*) Diamagnetic with two unpaired electrons
(*c*) Paramagnetic with two unpaired electrons
(*d*) Paramagnetic with no unpaired electrons

147. The species having bond order different from that in CO is
(*a*) NO^- (*b*) NO^+
(*c*) CN^- (*d*) N_2

148. Which one of the following has the strong O-O bond?
(*a*) O_2^+ (*b*) O_2
(*c*) O_2^- (*d*) O_2^{2-}

149. In which of the following pairs, the two species have identical bond order?
(*a*) N_2^- , O_2^{2-} (*b*) N_2^- , O_2^-
(*c*) N_2^- , O_2^+ (*d*) O_2^+ , N_2^{2-}

150. How may bonds does B_2 have?
(*a*) 0 (*b*) 1
(*c*) 2 (*d*) 3

151. According to MOT, C_2 molecule has
(*a*) 1σ & 1π bond (*b*) Only 2π-bond
(*c*) Only 2σ-bond (*d*) 1σ and 2π bond

152. Which of the following is diamagnetic?
(*a*) H_2^+ (*b*) O_2
(*c*) Li_2 (*d*) He_2^+

153. Molecular orbital theory was given by:
(*a*) Kossel (*b*) Moseley
(*c*) Mulliken (*d*) Werner

154. Molecular orbital electronic configuration for 'X' anion is- KK $\sigma^*2s^2\ \sigma^*2s^2\ \pi\ 2p_x^2\ \pi\ 2p_y^2\ \sigma 2p_z^2\ \pi^*\ 2p_x^1$, the anion 'X' is
(*a*) N_2^- (*b*) O_2^-
(*c*) N_2^{2-} (*d*) O_2^{2-}

155. The number of nodal planes present in σ^*2s anti- bonding orbital is
(*a*) 1 (*b*) 2
(*c*) 0 (*d*) 3

156. Which of the following is paramagnetic with bond order 0.5?
(*a*) F_2 (*b*) H_2^+
(*c*) N_2 (*d*) O_2^-

157. Which of the following pairs have identical bond order?
(*a*) N_2, O_2^{2+} (*b*) N_2, O_2^-
(*c*) N_2^-, O_2 (*d*) O^{2+}, N_2

158. The bond order of O_2^- is:
(*a*) 0.5 (*b*) 1.5
(*c*) 3.5 (*d*) 2.5

159. The molecule having bond order 3 is
(*a*) H_2 (*b*) N_2
(*c*) O_2 (*d*) He^{2+}

HYDROGEN BONDING

160. Which compound has electrovalent, covalent, coordinate as well as hydrogen bond ?
(*a*) NH_4Cl (*b*) HNO_3
(*c*) $CuSO_4.5H_2O$ (*d*) $BeCl_2.2H_2O$

161. Which of the following is the correct order of strength of H-bonding in the given compound ?
(*a*) $HF < NH_3$ (*b*) $H_2O > H_2O_2$
(*c*) $H_2O_2 > H_2O$ (*d*) $NH_3 > H_2O$

162. Strongest hydrogen bond is present in
(*a*) Ammonia (*b*) Water
(*c*) Hydrogen fluoride (*d*) Ethyl alcohol

163. Intermolecular hydrogen bond is not present in
(*a*) Ammonia (*b*) Water
(*c*) Hydrofluoric acid (*d*) Salicylaldehyde

164. Which of the following exists as a liquid at room temperature due to the formation of associated molecules only?
(*a*) Benzene (*b*) Ammonia
(*c*) Bromine (*d*) Carbon disulphide

165. Which of the following species has intramolecular H-bonds?
(*a*) Phenol (*b*) o-Nitrophenol
(*c*) p-Nitrophenol (*d*) Nitroethane

166. Bond energy of covalent O - H bond in water is
(*a*) Greater than bond energy of hydrogen bond
(*b*) Equal to bond energy of hydrogen bond
(*c*) Less than bond energy of hydrogen bond
(*d*) Half of the bond energy of hydrogen bond

167. Water is a liquid while hydrogen sulphide is a gas because
(*a*) Water has higher molecular weight
(*b*) Hydrogen sulphide is a weak acid
(*c*) Sulphur has high electronegativity than oxygen
(*d*) Water molecules associate through hydrogen bonding

168. Which of the following is a normal liquid ?
(*a*) NH_3 (*b*) H_2O
(*c*) HF (*d*) Br_2

169. Among the following, the boiling point is high for
(*a*) Ethyl alcohol (*b*) Dimethyl ether
(*c*) Acetone (*d*) Chloroform

170. Strongest hydrogen bond is present in
(*a*) CH_3OH (*b*) $CHCl_3$
(*c*) $(CH_3)_2CHOH$ (*d*) $(CH_3)_3COH$

171. Which of the following is miscible with water ?
(*a*) CS_2 (*b*) C_2H_5OH
(*c*) CCl_4 (*d*) $CHCl_3$

172. The coupling between base units of DNA is through
(*a*) Hydrogen bonding (*b*) Electrostatic bonding
(*c*) Covalent bonding (*d*) Vander Waal's forces

173. Which of the following is steam volatile ?
(*a*) Phenol (*b*) o - Nitrophenol
(*c*) m - Nitrophenol (*d*) p - Nitrophenol

174. Which of the following compounds would show evidence of the strongest hydrogen bonding ?

(*a*) Propan-l-ol (*b*) Propan-2-ol

(*c*) Propan-l,2-diol (*d*) Propan-l,2,3-triol

175. The compound having hydrogen bonds in it is

(*a*) NH_3 (*b*) H_2S

(*c*) HCl (*d*) PH_3

176. Strength of hydrogen bond is intermediate between

(*a*) Vander Waal forces and covalent bond

(*b*) Ionic bond and covalent bond

(*c*) Ionic bond and metallic bond

(*d*) Resonance

177. NH_3 has a much higher boiling point than PH_3 because

(*a*) NH_3 has a larger molecular weight

(*b*) NH_3 undergoes umbrella inversion

(*c*) NH_3 contains hydrogen bonds

(*d*) NH_3 contains ionic bonds where as PH_3 contains covalent bonds

178. The maximum number of molecules that one water molecule can hold through hydrogen bonding is

(*a*) 2 (*b*) 4

(*c*) 6 (*d*) 8

179. The high density of water compared to ice is due to

(*a*) H-bonding interactions

(*b*) Dipole-dipole interactions

(*c*) Dipole-induced dipole interactions

(*d*) Induced dipole-induced dipole interactions

180. The force responsible for the union of two ice blocks as a single block is

(*a*) Vanderwaals force

(*b*) Hydrogen bonds

(*c*) dipole interaction

(*d*) vanderwaals repulsion

181. The Electronegative elements between which the H-bond is formed principally in aqueous ammonia

(*a*) O, O (*b*) N,N

(*c*) O, N (*d*) Both O,O & N,N

182. Which of the following form hydrogen bond ?

(*a*) CH_3F (*b*) CH_3Cl

(*c*) CH_3Br (*d*) None

183. Chelation is observed in

(*a*) P-nitrophenol (*b*) O-nitrophenol

(*c*) Benzaldehyde (*d*) HF

184. H-bond is not present in:

(*a*) Water

(*b*) Glycerol

(*c*) Hydrogen fluoride

(*d*) Hydrogen sulphide

185. Which of the following has strongest hydrogen bonding?

(*a*) Ethyl amine (*b*) Ammonia

(*c*) Ethyl alcohol (*d*) Diethyl Ether

186. The high boiling point of water is due to-

(*a*) Co-ordinate bonding

(*b*) Covalent bonding

(*c*) Electrostatic force of attraction

(*d*) Hydrogen Bonding

187. When two ice cubes are pressed over each other, they unite to form one cube. Which of the following forces is responsible to hold them together?

(*a*) Dipole forces

(*b*) Vander waal's forces

(*c*) Covalent forces

(*d*) Hydrogen bond forces

188. Hydrogen bonding is maximum in:

(*a*) Ethyl chloride (*b*) Triethyl amine

(*c*) Ethanol (*d*) Diethyl ether

189. What is the dominant intermolecular force or bond that must be overcome in converting liquid CH_3OH to a gas?

(*a*) Hydrogen bonding

(*b*) Dipole-dipole force

(*c*) Covalent bond

(*d*) London dispersion force

190. Which of the following is least volatile?

(*a*) H_2O (*b*) H_2S

(*c*) H_2Se (*d*) H_2Te

191. The hydrogen bonding is maximum strong in:

(*a*) S-H-----S (*b*) N-H-----O

(*c*) S-H-----O (*d*) F-H------O

192. The comparatively high boiling point of HF is due to

(*a*) High reactivity of fluorine

(*b*) Small size of hydrogen atom

(*c*) Formation of hydrogen bonds and consequent association

(*d*) High IE of fluorine

193. Intramolecular H-bonding is present in:

(*a*) Meta- nitrophenol (*b*) Salicyaldehyde

(*c*) HCl (*d*) Benzophenone

Learning Plus

1. The correct order of hybridization of the central atom in the following species NH_3, PCl_5 and BCl_3 is

(*a*) sp^3d, sp^2 and sp^3 (*b*) sp^3 , sp^3d, sp^2

(*c*) sp^2, sp^3, sp^3d (*d*) sp^3, sp^2, sp^3d

2. How many molecule which has sp^3d hybridisation but planar as well as non-polar in structure?

$[ICl_2]^-$, BrF_5, SOF_4, SF_6, $XeOF_4$, ClF_3, I_3^-, $[I(CN)_2]^-$, PCl_5

(*a*) 2 (*b*) 4

(*c*) 6 (*d*) 8

3. 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*) CO_2 (*b*) HI

(*c*) H_2O (*d*) SO_2

4. Choose the INCORRECT options.

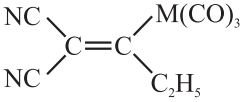
(*a*) All $d_{N\text{-}O}$ bond lengths are equal in nitrate ion.

(*b*) All $d_{C\text{-}O}$ bond lengths are equal in sodium carbonate.

(*c*) All $d_{C\text{-}O}$ bond lengths are equal in acetic acid.

(*d*) All $d_{B\text{-}O}$ bond length are equal in boric acid.

5. In a compound



The number of sigma and pi bonds respectively are

(*a*) 19, 11 (*b*) 19, 5

(*c*) 13, 11 (*d*) 7, 3

6. Hydrogen bonds are formed in many compounds e.g., H_2O, HF, NH_3. The boiling point of such compounds depends to a extent on the strength of hydrogen bond and the number of hydrogen bonds. The correct decreasing order of the boiling points above compounds is:

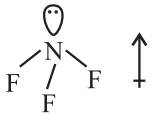
(*a*) $HF > H_2O > NH_3$ (*b*) $H_2O > HF > NH_3$

(*c*) $NH_3 > HF > H_2O$ (*d*) $NH_3 > H_2O > HF$

7. Choose the correct statement:

(*a*) NH_3 is having bond angle of 109°28'.

(*b*) The direction of the dipole moment of NF_3 is as shown in the diagram



(*c*) sp^2 hybrid orbital is consisting of 66.67 % '*p*' character.

(*d*) None of these

8. Among the triatomic molecules/ions $BeCl_2$, XeF_2, ICl_2^+, I_3^+, $I(CN)_2^-$

How many are sp^3d hybridised and linear choose the correct pairs ?

(*a*) $BeCl_2$ & XeF_2 (*b*) XeF_2 & $I(CN)_2^-$

(*c*) XeF_2 & ICl_2^+ (*d*) $BeCl_2$ & I_3^+

9. In PO_4^{3-} ion the average formal charge on the oxygen atom of P–O bond is:

(*a*) + 1 (*b*) –1

(*c*) – 0.75 (*d*) + 0.75

10. In NO_3^- ion, the number of bond pairs and lone pairs of electrons on nitrogen atom are:

(*a*) 2, 2 (*b*) 3, 1

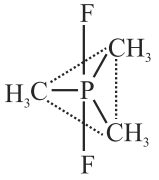
(*c*) 1, 3 (*d*) 4, 0

11. If the equatorial plane is x–y plane in sp^3d hybridisation then the orbital used in pd hybridisation are -

(*a*) p_z and d_{z^2} (*b*) p_x and d_{xy}

(*c*) p_y and d_{yz} (*d*) None of these

12. How many maximum number of atoms are present in one plane in $PF_2(CH_3)_3$ molecule, only consider those geometry which is non-polar in nature?



(*a*) 2 (*b*) 4 (*c*) 7 (*d*) 8

13. Which of the following statement is true for $IO_2F_2^-$?

(*a*) The electrons are located at the corners of a trigonal bipyramidal but one of the equatorial pairs is unshared.

(*b*) It has sp^3d hybridisation and is T–shaped.

(*c*) Its structure is analogous to SF_4.

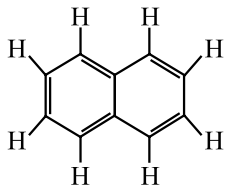
(*d*) (*a*) and (*c*) both

14. Which of the following species has tetrahedral geometry?

(*a*) BH_4^- (*b*) NH_2^-

(*c*) CO_3^{2-} (*d*) H_3O^+

15. Number of π bonds and σ bonds in the following structure is:



(*a*) 6, 19 (*b*) 4, 20 (*c*) 5, 19 (*d*) 5, 20

16. In NO_2^+, N undergoes

(*a*) sp^3 hybridisation (*b*) sp^2 hybridisation

(*c*) sp hybridisation (*d*) sp^2d hybridisation

17. Which of the following should have pyramidal shape ?

(*a*) $[ClOF_2]^+$ (*b*) ICl_3

(*c*) $[BrICl]^-$ (*d*) All of these

18. Among the triatomic molecules/ion $BeCl_2$, NO_2^+, ICl_2^-, XeF_2 , I_3^-, OCl_2 total number of linear molecule(s)/ion (s) are :

(*a*) 2 (*b*) 3

(*c*) 4 (*d*) 5

19. Which molecule/ion out of the following does not contain unpaired electrons?

(*a*) N_2^+ (*b*) O_2 (*c*) O_2^{2-} (*d*) B_2

20. In which of the following molecule/ion all the bonds are not equal?

(*a*) XeF_4 (*b*) BF_4^- (*c*) C_2H_4 (*d*) SiF_4

21. The geometry of ammonia molecule can be best described as:

(*a*) Nitrogen at one vertex of a regular tetrahedron, the other three vertices being occupied by the three hydrogens

(*b*) Nitrogen at the centre of the tetrahedron, three of the vertices being occupied by three hydrogens

(*c*) Nitrogen at the centre of an equilateral triangle, three corners being occupied by three hydrogens

(*d*) Nitrogen at the junction of a T, three open ends being occupied by three hydrogens

22. Which molecular geometry is least likely to result from a trigonal bypyramidal electron geometry?

(*a*) Trigonal planar (*b*) see–saw

(*c*) Linear (*d*) T–shaped

23. Which of the following has bond angle greater than 120°?

(*a*) OCl_2 (*b*) $O(CH_3)_2$

(*c*) $O(SiH_3)_2$ (*d*) NO_2^-

24. The correct order of bond angle is :

(*a*) $H_2S < NH_3 < BF_3 < CH_4$

(*b*) $NH_3 < H_2S < CH_4 < BF_3$

(*c*) $H_2S < NH_3 < CH_4 < BF_3$

(*d*) $H_2S < CH_4 < NH_3 < BF_3$

25. Consider the following molecules ;

H_2O	H_2S	H_2Se	H_2Te
I	II	III	IV

Arrange these molecules in increasing order of bond angles.

(*a*) I < II < III < IV (*b*) IV < III < II < I

(*c*) I < II < IV < III (*d*) II < IV < III < I

26. If the electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$, the four electrons involved in chemical bond formation will be:

(*a*) $3p^6$ (*b*) $3p^6, 4s^2$

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

27. Which of the following angle corresponds to sp^2 hybridization?

(*a*) 90° (*b*) 120°

(*c*) 180° (*d*) 109°

28. The bond angle in PH_3 is :

(*a*) Much lesser than NH_3

(*b*) Equal to that in NH_3

(*c*) Much greater than in NH_3

(*d*) Slightly more than in NH_3

29. In which of the following sets, all the species are diamagnetic in nature?

(*a*) $[O_2^{2-}, O_2^{2+}, N_2^{2-}]$ (*b*) $[C_2, O_2^{2-}, N_2^{2+}]$

(*c*) $[O_2^{2-}, O_2^{+}, O_2^{2-}]$ (*d*) $[N_2^{+}, O_2^{2+}, B_2]$

30. Which of the following order of energies of molecular orbitals of N_2 is correct?

(*a*) $(\pi 2p_y) < (\sigma 2p_z) < (\pi^* 2p_x) \approx (\pi^* 2p_y)$

(*b*) $(\pi 2p_y) > (\sigma 2p_z) > (\pi^* 2p_x) \approx (\pi^* 2p_y)$

(*c*) $(\pi 2p_y) < (\sigma 2p_z) > (\pi^* 2p_x) \approx (\pi^* 2p_y)$

(*d*) $(\pi 2p_y) > (\sigma 2p_z) < (\pi^* 2p_x) \approx (\pi^* 2p_y)$

31. Which of the following statement is not correct from the view point molecular orbital theory?

(*a*) Be_2 is not a stable molecule

(*b*) He_2 is not stable but He_2^+ is expected to exist.

(*c*) Bond strength of N_2 is maximum amongst the homonuclear diatomic molecules belonging to the second period.

(*d*) The order of energies of molecular orbitals in N_2 molecule is $\sigma 1s^2\, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 = \pi 2p_y^2$

32. Among KO_2, Al_2O_3 and BaO_2, the unpaired electron is present in:

(*a*) KO_2 only (*b*) KO_2, BaO_2

(*c*) Only Al_2O_3 (*d*) Only BaO_2

33. The fluoride whose value of dipole moment is not equal to zero, is:

(*a*) XeF_4 (*b*) CF_4

(*c*) SF_4 (*d*) PF_5

34. In which of the following, d-orbitals are not used by central atom in hybridisation?

(*a*) PF_5 (*b*) PCl_5

(*c*) PBr_5 (gaseous) (*d*) None of these

35. BrO_3^- is isostructural with which of the following ?

(*a*) XeO_3 (*b*) XeF_3

(*c*) XeF_4 (*d*) XeO_2

36. In which of the following compounds B–F length is shortest ?

(*a*) BF_4^- (*b*) $BF_3 \rightarrow NH_3$

(*c*) BF_3 (*d*) $BF_3 \rightarrow N(CH_3)_3$

Multiconcept MCQs

1. Assuming the bond direction to be z-axis, which of the overlapping of atomic orbitals of two atom (A) and (B) will result in bonding?

(I) s-orbital of A and p_x orbital of B

(II) s-orbital of A and p_z orbital of B

(III) p_y-orbital of A and p_z orbital of B

(IV) s-orbitals of both (A) and (B)

(*a*) I and IV (*b*) I and II (*c*) III and IV (*d*) II and IV

2. Choose the INCORRECT property for the following molecules:

(I) CH_2F_2 ; (II) CHF_3 ; (III) CH_3F

(*a*) C–F bond length order : $CH_3F > CH_2F_2 > CHF_3$

(*b*) C–H bond length order : $CH_3F > CH_2F_2 > CHF_3$

(*c*) Shape of the all species is perfectly tetrahedral

(*d*) Dipole moment is non zero for given compounds.

3. Identify the correct match.

(*i*) XeF_2	(A) Central atom has sp^3 hybridisation and bent geometry.
(*ii*) N_3^-	(B) Central atom has sp^3d^2 hybridisation and octahedral.
(*iii*) PCl_6^-(PCl_5 (s) anion)	(C) Central atom has sp hybridisation and linear geometry.
(*iv*) ICl_2^+ (I_2Cl_6 (ℓ) cation)	(D) Central atom has sp^3d hybridisation and linear geometry.

(*a*) (*i* – A), (*ii* – B), (*iii* – C), (*iv* – D)

(*b*) (*i* – D), (*ii* – B), (*iii* – D), (*iv* – C)

(*c*) (*i* – B), (*ii* – C), (*iii* – A), (*iv* – D)

(*d*) (*i* – D), (*ii* – C), (*iii* – B), (*iv* – A)

4. Which is not correctly matched?

(*a*) XeO_3 ; Trigonal bipyramidal

(*b*) ClF_3 ; bent T-shape

(*c*) $XeOF_4$; Square pyramidal

(*d*) XeF_2 ; Linear shape

5. Which of the following compounds have the same no. of lone pairs with their central atom ?

[I] XeF_5^- ; [II] BrF_3 ; [III] XeF_2 ; [IV] H_3S^+ [V] : CH_2

(*a*) IV and III

(*b*) I and III

(*c*) I and II

(*d*) II, IV and V

6. Choose the correct order of bond angle.

	Column-I		Column-II
A	NO_2^+	**1.**	180°
B	NO_2	**2.**	134°
C	NO_2^-	**3.**	120°
D	NO_3^-	**4.**	115°
		5.	109°

	A	B	C	D		A	B	C	D
(*a*)	5	4	3	2	(*b*)	5	2	4	3
(*c*)	1	2	4	3	(*d*)	1	4	3	2

7. N_2 and O_2 are converted to monocations N_2^+ and O_2^+ respectively, which is wrong statement ?

(*a*) In N_2^+, the N–N bond weakens

(*b*) In O_2^+, the O–O bond order increases

(*c*) In O_2^+, the paramagnetism decreases

(*d*) N_2^+ becomes diamagnetic

8. 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 – bp > bp – bp

(*ii*) In general, as the number of lone pair of electrons on central atom increases, value of bond angle from normal 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.

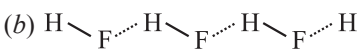
(*iv*) The structures of Xenon-fluorides and Xenon-oxyfluorides could not be explained on the basis of VSEPR theory.

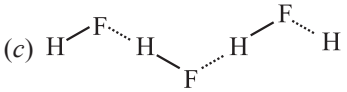
(*a*) TTTF (*b*) TFTF

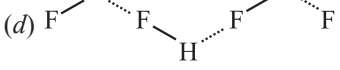
(*c*) TFTT (*d*) TFFF

9. The *H* bond in solid *HF* can be best represented as:

(*a*) H–F...H–F...H–F

(*b*) 

(*c*) 

(*d*) 

10. The experimental value of the dipole moment of HCl is 1.03 D. The length of the H – Cl bond is 1.278Å. The percentage of ionic character in HCl is:

(*a*) 43 (*b*) 21

(*c*) 17 (*d*) 7

11. Which of the following pair of molecules have the same shape but difference in polarity (polar or non-polar)?

(*a*) H_2O and NH_3 (*b*) $SnCl_2$ and SO_2

(*c*) CO_2 and N_2O (*d*) SO_2 and SO_3

12. Hydrogen bonding plays a central role in the following phenomenon :

(*a*) Ice floats in water

(*b*) Higher Lewis basicity of primary amines than tertiary amines in aqueous solutions.

(*c*) Formic acid is more acidic than acetic acid

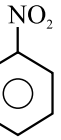
(*d*) Dimerisation of acetic acid in benzene

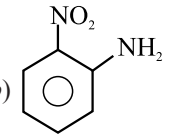
13. Among the following pairs, which two species are not isostructural?

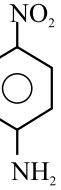
(*a*) PF_6^- and SF_6 (*b*) SiF_4 and SF_4

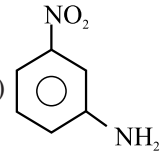
(*c*) IO_3^- and XeO_3 (*d*) BH_4^-, NH_4^+

14. Which compound exhibits maximum dipole moment among the following?

(*a*) 

(*b*) 

(*c*) 

(*d*) 

15. Which of the following is a wrong order with respect to the property mentioned against each ?

(*a*) $O_2^{2-} > O_2 > O_2^+$ [Paramagnetic moment]

(*b*) Bond Angle : $H_2O > H_2S > H_2Se$

(*c*) $H_2 > H_2^+ > He_2^+$ [stability]

(*d*) $NO_2^+ > NO_2 > NO_2^-$ [bond angle]

16. O_2F_2 is an unstable yellow orange solid and H_2O_2 is a colourless liquid, both have O–O bond and O–O bond length in H_2O_2 and O_2F_2 respectively is -

(*a*) 1.22 Å, 1.48 Å

(*b*) 1.48 Å, 1.22 Å

(*c*) 1.22 Å, 1.22 Å

(*d*) 1.48 Å, 1.48 Å

17. Which bond angle θ would result in maximum dipole moment for the triatomic molecule XY_2 shown below ?

(*a*) θ = 90° (*b*) θ = 120°

(*c*) θ = 150° (*d*) θ = 180°

18. Among the following select the correct statements:

I. PH_6 does not exist

II. pπ-dπ bond is present in SO_2

III. SeF_4 and CH_4 have the same shape

IV. I_3^+ has a bent shape

(*a*) II, III (*b*) I, II

(*c*) I, IV (*d*) I, II, IV

19. Match List I (Molecules)with List II (Bond order) and select the correct answer using the codes.

List – I	List – II
I. Li_2	A. 3
II. N_2	B. 1.5
III. Be_2	C. 0
IV. O_2	D. 0
	E. 2

Codes

(*a*) I - B, II - C, III - A, IV - E (*b*) I - C, II - A, III - D, IV - E

(*c*) I - D, II - A, III - E, IV - C (*d*) I - C, II - B, III - E, IV - A

20. The Atomic numbers of three elements A,B and C are a, a+1 and a+2. C is an alkali metal. In a compound of A and C, the nature of bonding is

(*a*) Coordinate

(*b*) Covalent

(*c*) Ionic

(*d*) Metallic

21. The electronegativity difference between N and F is greater than N and H, yet the dipole moment of NH_3 (1.5D) is greater than that of NF_3 (0.2D). This is because:

(*a*) In NH_3 as well as NF_3, the orbital dipole and bond dipole are in opposite direction

(*b*) In NH_3, the orbital dipole and bond dipole are in the opposite direction, where as in NF_3 these are in same direction

(*c*) In NH_3 , as well as in NF_3 the orbital dipole and bond dipole are in same direction.

(*d*) In NH_3, the orbital dipole and bond dipole are in same direction where as in NF_3 these are in opposite direction.

22. Which of the following statements are correct?

(I) The hybridisation found in cation of solid PCl_5 is sp^3.

(II) In AB_2L_2 type, the BAB bond angle is always greater than the normal tetrahedral bond angle.

(III) In ClO_3^-, NH_3 and XeO_3, the hybridisation and the number of lone pairs on the central atoms are same.

(IV) In discrete P_4 molecule, there are six P—P bonds and four lone pairs of electrons.

(*a*) I, II and III only

(*b*) I, III and IV only

(*c*) III and IV only

(*d*) All are correct

23. Which of the following statements are correct? (As the s-character of a hybrid orbital decreases).

(I) The bond angle decreases

(II) The bond strength increases

(III) The bond length increases.

(*a*) (I) and (III) (*b*) (II) and (III)

(*c*) (I) and (II) (*d*) All are correct

24. The correct order of increasing C - O bond length of CO, CO_3^{2-}, CO_2 is

(*a*) $CO_3^{2-} < CO_2 < CO$

(*b*) $CO_2 < CO_3^{2-} < CO$

(*c*) $CO < CO_3^{2-} < CO_2$

(*d*) $CO < CO < CO^{2-}$

NEET Past 10 Years Questions

1. BF_3 is planar and electron deficient compound. Hybridization and number of electrons around the **central atom**, respectively are: **(2021)**

(*a*) sp^3 and 6 (*b*) sp^2 and 6

(*c*) sp^2 and 8 (*d*) sp^3 and 4

2. Match List-I with List-II. **(2021)**

List-I		List-II	
(A)	PCl_5	(*i*)	Square pyramidal
(B)	SF_6	(*ii*)	Trigonal planar
(C)	BrF_5	(*iii*)	Octahedral
(D)	BF_3	(*iv*)	Trigonal bipyramidal

Choose the correct answer from the options given below.

(*a*) A-ii B-iii C-iv D-i

(*b*) A-iii B-i C-iv D-ii

(*c*) A-iv B-iii C-ii D-i

(*d*) A-iv B-iii C-i D-ii

3. Which of the following molecules is non-polar in nature? **(2021)**

(*a*) CH_2O (*b*) $SbCl_5$

(*c*) NO_2 (*d*) $POCl_3$

4. Which of the following set of molecules will have zero dipole moment? **(2020)**

(*a*) Boron trifluoride, hydrogen fluoride, carbon dioxide, 1,3-dichlorobenzene

(*b*) Nitrogen trifluoride, beryllium difluoride, water 1, 3-dichlorobenzzene

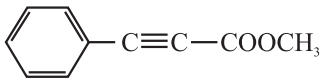
(*c*) Boron trifluoride, berylium difluoride, carbon dioxide, 1,4-dichlorobenzene

(*d*) Ammonia, beryllium difluoride, water, 1,4-dichlorobenzene

5. Identify a molecule which does not exist. **(2020)**

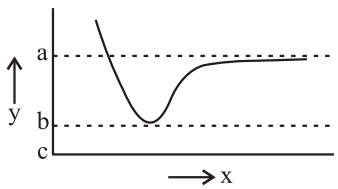
(*a*) Li_2 (*b*) C_2 (*c*) O_2 (*d*) He_2

6. How many (i) sp^2 hybridised carbon atoms and (ii) π bonds are present in the following compound? **(2020 Covid Re-NEET)**



(*a*) 8, 6 (*b*) 7, 6 (*c*) 8, 5 (*d*) 7, 5

7. The potential energy (y) curve for H_2 formation as a function of internuclear distance (x) of the H atoms is shown below.



The bond energy of H_2 is **(2020 Covid Re-NEET)**

(*a*) $\frac{(c-a)}{2}$ (*b*) $\frac{(b-a)}{2}$

(*c*) (c – a) (*d*) (b – a)

8. Identify the wrongly matched pair. **(2020 Covid Re-NEET)**

	Molecule	**Shape or geometry of molecule**
(*a*)	SF_6	Octahedral
(*b*)	$BeCl_2$	Linear
(*c*)	NH_3	Trigonal pyramidal
(*d*)	PCl_5	Trigonal planar

9. Which of the following diatomic molecular species has only π bonds according to Molecular Orbital Theory? **(2019)**

(*a*) O_2 (*b*) N_2

(*c*) C_2 (*d*) Be_2

10. Identify the incorrect statement related to PCl_5 from the following: **(2019)**

(*a*) Three equatorial P-Cl bonds make an angle of 120° with each other

(*b*) Two axial P-Cl bonds make an angle of 180° with each other

(*c*) Axial P-Cl bonds are longer than equatorial P-Cl bonds

(*d*) PCl_5 molecule is non-reactive

11. Consider the following species: **(2018)**

CN^+, CN^-, NO and CN

Which one of these will have the highest bond order?

(*a*) NO (*b*) CN^- (*c*) CN (*d*) CN^+

12. Which one of the following pair of species have the same bond order? **(2017-Delhi)**

(*a*) N_2, O_2^- (*b*) CO, NO

(*c*) O_2, NO^+ (*d*) CN^-, CO

13. The species, having bond angles of 120° is **(2017-Delhi)**

(*a*) BCl_3 (*b*) PH_3

(*c*) ClF_3 (*d*) NCl_3

14. Which one of the following ions is not tetrahedral in shape? **(2017-Gujarat)**

(*a*) $[NiCl_4]^{2-}$ (*b*) NH_4^+

(*c*) BF_4^- (*d*) $[Cu(NH_3)_4]^{2+}$

15. Which of the following pair of species is not iso-structural? **(2017-Gujarat)**

(*a*) BrO_3^-, XeO_3 (*b*) ICl_4^-, XeF_4

(*c*) ClO_3^-, CO_3^{2-} (*d*) IBr_2^-, XeF_2

16. Which of the following hydrides has the largest bond angle? **(2017-Gujarat)**

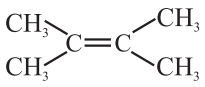
(*a*) H_2Se (*b*) H_2S (*c*) H_2Te (*d*) H_2O

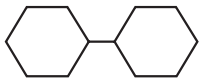
17. Which one of the following compounds shows the presence of intramolecular hydrogen bond? **(2016 - II)**

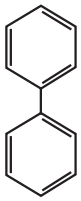
(*a*) Cellulose (*b*) Concentrated acetic acid

(*c*) H_2O_2 (*d*) HCN

18. In which of the following molecules, all atoms are coplanar? **(2016 - II)**

(*a*) 

(*b*) 

(*c*) 

(*d*) 

19. Among the following which one is a wrong statement? **(2016 - II)**

(*a*) SeF_4 and CH_4 have same shape

(*b*) I_3^+ has bent geometry

(*c*) PH_5 and $BiCl_5$ do not exist

(*d*) $p\pi$ - $d\pi$ bonds are present in SO_2

20. The hybridisations of atomic orbitals of nitrogen in NO_2^+, NO_3^- and NH_4^+ respectively are: **(2016 - II)**

(*a*) *sp*, sp^3 and sp^2 (*b*) sp^2, sp^3 and *sp*

(*c*) *sp*, sp^2 and sp^3 (*d*) sp^2, *sp* and sp^3

21. Which of the following pairs of ions is isoelectronic and isostructural? **(2016 - II)**

(*a*) CO_3^{2-} , NO_3^- (*b*) ClO_3^- , CO_3^{2-}

(*c*) SO_3^{2-} , NO_3^- (*d*) ClO_3^- , SO_3^{2-}

22. Consider the molecules CH_4, NH_3 and H_2O. Which of the given statement is false? **(2016 - I)**

(*a*) The H─C─H bond angle in CH_4 is larger than the H─N─H bond angle in NH_3

(*b*) The H─C─H bond angle in CH_4, the H─N─H bond angle in NH_3, and the H─O─H bond angle in H_2O are all greater than 90º.

(*c*) Then H─O─H bond angle in H_2O is larger than the H─C─H bond angle in CH_4

(*d*) The H─O─H bond angle in H_2O is smaller than the H─N─H bond angle in NH_3

23. Predict the correct order among the following: **(2016 - I)**

(*a*) Lone pair ─ bond pair > bond pair ─ bond pair > lone pair ─ lone pair

(*b*) Lone pair ─ lone pair > lone pair ─ bond pair > bond pair ─ bond pair

(*c*) Lone pair ─ lone pair > bond pair ─ bond pair > lone pair ─ bond pair

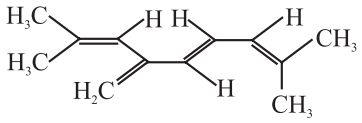
(*d*) Bond pair ─ bond pair > lone pair ─ bond pair > lone pair ─ lone pair

24. Decreasing order of stability of O_2 , O_2^- , O_2^+ and O_2^{2-} is: **(2015 RE)**

(*a*) $O_2^- > O_2^{2-} > O_2^+ > O_2$ (*b*) $O_2^+ > O_2 > O_2^- > O_2^{2-}$

(*c*) $O_2^{2-} > O_2^- > O_2 > O_2^+$ (*d*) $O_2 > O_2^+ > O_2^{2-} > O_2^-$

25. The total number of π-bond electrons in the following structure is: **(2015)**



(*a*) 8 (*b*) 12 (*c*) 16 (*d*) 4

26. Which of the following options represents the correct bond order? **(2015)**

(*a*) $O_2^- < O_2 < O_2^+$ (*b*) $O_2^- > O_2 < O_2^+$

(*c*) $O_2^- < O_2 > O_2^+$ (*d*) $O_2^- > O_2 > O_2^+$

27. The correct bond order in the following species is: **(2015)**

(*a*) $O_2^{2+} < O_2^- < O_2^+$ (*b*) $O_2^{2+} < O_2^- < O_2^+$

(*c*) $O_2^- < O_2^+ < O_2^{2+}$ (*d*) $O_2^{2+} < O_2^+ < O_2^-$

28. Which of the following pairs of ions are isoelectronic and isostructural? **(2015)**

(*a*) SO_3^{2-} , NO_3^- (*b*) ClO_3^- , SO_3^{2-}

(*c*) CO_3^{2-} , SO_3^{2-} (*d*) ClO_3^- , CO_3^{2-}

29. Maximum bond angle at nitrogen is present in which of the following? **(2015)**

(*a*) NO_2^+ (*b*) NO_3^- (*c*) NO_2 (*d*) NO_2^-

30. Which of the following molecules has the maximum dipole moment? **(2014)**

(*a*) CH_4 (*b*) NH_3 (*c*) NF_3 (*d*) CO_2

31. Which one of the following species has planar triangular shape? **(2014)**

(*a*) NO_3^- (*b*) NO_2^-

(*c*) CO_2 (*d*) N_3

32. Which of the following organic compounds has same hybridisation as its combustion product (CO_2)? **(2014)**

(*a*) Ethyne (*b*) Ethene

(*c*) Ethanol (*d*) Ethane

33. Which of the following is a polar molecule? **(2013)**

(*a*) BF_3 (*b*) SF_4

(*c*) SiF_4 (*d*) XeF_4

34. Which of the following is electron-deficient? **(2013)**

(*a*) $(CH_3)_2$ (*b*) $(SiH_3)_2$

(*c*) $(BH_3)_2$ (*d*) PH_3

35. Which of the following is paramagnetic? **(2013)**

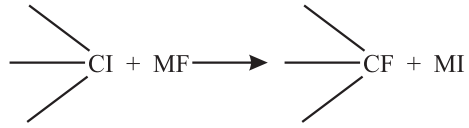
(*a*) CO (*b*) O_2^- (*c*) CN^- (*d*) NO^+

36. Four diatomic species are listed below. Identify the correct order in which the bond order is increasing in them: **(2012 Mains)**

(*a*) $He_2^+ < O_2^- < NO < C_2^{2-}$ (*b*) $NO < O_2^- < C_2^{2-} < He_2^+$

(*c*) $O_2^- < NO < C_2^{2-} < He_2^+$ (*d*) $C_2^{2-} < He_2^+ < O_2^- < NO$

37. In the replacement reaction



The reaction will be most favorable if M happens to be:

(*2012 Mains*)

(*a*) Li (*b*) Na (*c*) K (*d*) Rb

38. The pair of species with the same bond order is: **(*2012 Pre*)**

(*a*) N_2 , O_2 (*b*) O_2^{2-}, B_2

(*c*) O_2^+ , NO^+ (*d*) NO , CO

39. Bond order of 1.5 is shown by: **(*2012 Pre*)**

(*a*) O_2 (*b*) O_2^+

(*c*) O_2^- (*d*) O_2^{2-}

40. Which one of the following pairs is iso-structural (i.e., having the same shape and hybridisation)? **(*2012 Pre*)**

(*a*) [BF_4^- and NH_4^+] (*b*) [BCl_3 and $BaCl_3$]

(*c*) [NH_3 and NO_3^-] (*d*) [NF_3 and BF_3]

41. During change of O_2 to O_2^- ion, the electron adds on which one of the following orbitals? **(*2012 Mains*)**

(*a*) π^* orbital (*b*) π orbital

(*c*) σ^* orbital (*d*) σ orbital

ANSWER KEY

Topicwise Questions

1. (*c*) **2.** (*b*) **3.** (*b*) **4.** (*b*) **5.** (*d*) **6.** (*c*) **7.** (*b*) **8.** (*b*) **9.** (*b*) **10.** (*d*)
11. (*d*) **12.** (*c*) **13.** (*d*) **14.** (*b*) **15.** (*c*) **16.** (*b*) **17.** (*c*) **18.** (*d*) **19.** (*a*) **20.** (*c*)
21. (*a*) **22.** (*a*) **23.** (*d*) **24.** (*b*) **25.** (*c*) **26.** (*c*) **27.** (*a*) **28.** (*d*) **29.** (*c*) **30.** (*c*)
31. (*b*) **32.** (*b*) **33.** (*a*) **34.** (*b*) **35.** (*b*) **36.** (*b*) **37.** (*c*) **38.** (*d*) **39.** (*b*) **40.** (*c*)
41. (*c*) **42.** (*c*) **43.** (*a*) **44.** (*b*) **45.** (*a*) **46.** (*a*) **47.** (*c*) **48.** (*a*) **49.** (*b*) **50.** (*b*)
51. (*a*) **52.** (*b*) **53.** (*d*) **54.** (*c*) **55.** (*a*) **56.** (*d*) **57.** (*c*) **58.** (*d*) **59.** (*d*) **60.** (*a*)
61. (*d*) **62.** (*b*) **63.** (*b*) **64.** (*d*) **65.** (*d*) **66.** (*c*) **67.** (*c*) **68.** (*d*) **69.** (*d*) **70.** (*c*)
71. (*b*) **72.** (*b*) **73.** (*d*) **74.** (*b*) **75.** (*c*) **76.** (*b*) **77.** (*c*) **78.** (*b*) **79.** (*a*) **80.** (*b*)
81. (*c*) **82.** (*d*) **83.** (*c*) **84.** (*b*) **85.** (*b*) **86.** (*c*) **87.** (*c*) **88.** (*a*) **89.** (*b*) **90.** (*b*)
91. (*a*) **92.** (*d*) **93.** (*a*) **94.** (*b*) **95.** (*c*) **96.** (*b*) **97.** (*b*) **98.** (*b*) **99.** (*d*) **100.** (*b*)
101. (*a*) **102.** (*c*) **103.** (*c*) **104.** (*b*) **105.** (*d*) **106.** (*c*) **107.** (*d*) **108.** (*a*) **109.** (*a*) **110.** (*c*)
111. (*a*) **112.** (*d*) **113.** (*d*) **114.** (*d*) **115.** (*b*) **116.** (*a*) **117.** (*b*) **118.** (*b*) **119.** (*b*) **120.** (*a*)
121. (*c*) **122.** (*d*) **123.** (*a*) **124.** (*a*) **125.** (*a*) **126.** (*b*) **127.** (*d*) **128.** (*a*) **129.** (*c*) **130.** (*b*)
131. (*d*) **132.** (*c*) **133.** (*d*) **134.** (*d*) **135.** (*c*) **136.** (*d*) **137.** (*c*) **138.** (*b*) **139.** (*b*) **140.** (*b*)
141. (*c*) **142.** (*b*) **143.** (*a*) **144.** (*b*) **145.** (*c*) **146.** (*c*) **147.** (*a*) **148.** (*a*) **149.** (*c*) **150.** (*b*)
151. (*b*) **152.** (*c*) **153.** (*c*) **154.** (*a*) **155.** (*a*) **156.** (*b*) **157.** (*a*) **158.** (*b*) **159.** (*b*) **160.** (*c*)
161. (*b*) **162.** (*c*) **163.** (*d*) **164.** (*b*) **165.** (*b*) **166.** (*a*) **167.** (*d*) **168.** (*d*) **169.** (*a*) **170.** (*a*)
171. (*b*) **172.** (*a*) **173.** (*b*) **174.** (*d*) **175.** (*a*) **176.** (*a*) **177.** (*c*) **178.** (*b*) **179.** (*a*) **180.** (*b*)
181. (*c*) **182.** (*d*) **183.** (*b*) **184.** (*d*) **185.** (*c*) **186.** (*d*) **187.** (*d*) **188.** (*c*) **189.** (*a*) **190.** (*a*)
191. (*d*) **192.** (*c*) **193.** (*b*)

Learning Plus

1. (*b*) **2.** (*a*) **3.** (*c*) **4.** (*c*) **5.** (*a*) **6.** (*b*) **7.** (*c*) **8.** (*b*) **9.** (*c*) **10.** (*d*)
11. (*a*) **12.** (*c*) **13.** (*d*) **14.** (*a*) **15.** (*c*) **16.** (*c*) **17.** (*a*) **18.** (*d*) **19.** (*c*) **20.** (*c*)
21. (*b*) **22.** (*a*) **23.** (*c*) **24.** (*c*) **25.** (*b*) **26.** (*d*) **27.** (*b*) **28.** (*a*) **29.** (*b*) **30.** (*a*)
31. (*d*) **32.** (*a*) **33.** (*c*) **34.** (*d*) **35.** (*a*) **36.** (*c*)

Multiconcept MCQs

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

NEET Past 10 Years Questions

1. (*b*) **2.** (*d*) **3.** (*b*) **4.** (*c*) **5.** (*d*) **6.** (*b*) **7.** (*d*) **8.** (*d*) **9.** (*c*) **10.** (*d*)
11. (*b*) **12.** (*d*) **13.** (*a*) **14.** (*d*) **15.** (*c*) **16.** (*d*) **17.** (*a*) **18.** (*c*) **19.** (*a*) **20.** (*c*)
21. (*a*,*d*) **22.** (*c*) **23.** (*b*) **24.** (*b*) **25.** (*a*) **26.** (*a*) **27.** (*c*) **28.** (*b*) **29.** (*a*) **30.** (*b*)
31. (*a*) **32.** (*a*) **33.** (*b*) **34.** (*c*) **35.** (*b*) **36.** (*a*) **37.** (*d*) **38.** (*b*) **39.** (*c*) **40.** (*a*)
41. (*a*)

Solution

Topicwise Questions

1. (*c*) ∵ The valency of O is 2, by cross method the compound becomes M_2O_4 which is same as MO_2.

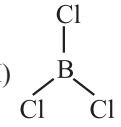
2. (*b*) Valency is double to the number of oxygen atoms reacting.

3. (*b*) Differentiating electrons enters into (n-1)d electrons participate in bonding. It is property of transition metal.

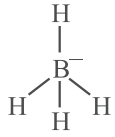
4. (*b*) The central atom of Be contains 4 electrons

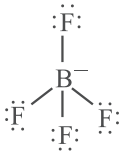
5. (*d*) XeF_4 has 10 electrons in valence shell of Xe

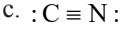
6. (*c*) PCl_5 has 10 electrons in valence shell of P.

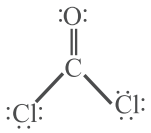
7. (*b*) (I) Cl–Be–Cl (II)  (III) H–H

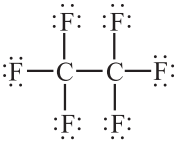
8. (*b*) Strongest Lewis Acid is one with maximum O.S.

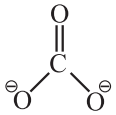
9. (*b*) a.  lp = 0

b.  lp = 12

c.  lp = 2

d.  lp = 8

10. (*d*)  all have complete octet

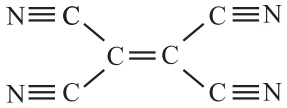
11. (*d*) 

12. (*c*) $CO(NH_2)_2$ **13.** (*d*) 14

14. (*b*) Due to Fajan's rule, according to which the cation distorts the electron cloud of the anion and pulls the electron density towards itself. Thus, it introduces some covalent character in the ionic bond. Now, the smaller the cation, greater will be its polarizability.

15. (*c*) Ionic character increases as: $BeCl_2 < MgCl_2 < CaCl_2 < BaCl_2$ because larger the size of the cation, higher is the ionic character.

16. (*b*) **17.** (*c*) **18.** (*d*)

19. (*a*) 

20. (*c*) Since the molecule possess 6 pairs of electrons, it should exhibit octahedral geometry. But as its structure is planar, it will have 2 lone pairs of electrons, and 4 sigma bonds.

21. (*a*)

22. (*a*) More electropositive and more electronegative combination gives more ionic compound.

23. (*d*) Ionic bond is formed by ions.

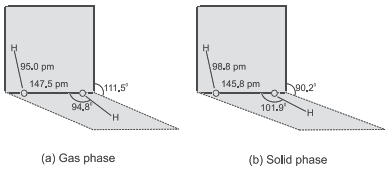
24. (*b*) Small anion forms more ionic compound.

25. (*c*) All the C–atoms in benzene are sp^2 hybridised.

26. (*c*) Greater the strength of bond, more is the directional character.

27. (*a*) Sodium Fluoride

28. (*d*) The electronegativity of S is least among others. Hence H-S bond is least polar in nature.

29. (*c*) 

30. (*c*) ΔEN. = 0 ∴ non-polar.

31. (*b*) 4.8 e.s.u Cm = 1.6×10^{-29} col.m

1 e.s.u. Cm = $\frac{1.6\times10^{-29}}{4.8}$

$= 0.33 \times 10^{-29}$ col.m

$= 3.33 \times 10^{-30}$ col.m

32. (*b*) In HCl, H–atom and Cl–atom shares one electron pair to form covalent bond.

33. (*a*) Li^+ being smaller, according to Fajan rule, Li^+ has high polarising power. Hence LiCl should have maximum ionic character.

34. (*b*) Symmetrical with linear structure will have zero dipole moment.

35. (*b*) $O = \overset{\oplus}{N} = O$ Bond order = 2

36. (*b*) In diamond, one carbon atom is attached to more carbon atoms by means of covalent bonds.

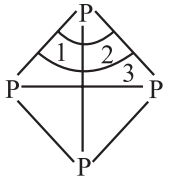
37. (*c*) NH_4^+ is sp^3 hybridised

$$\left[H = \frac{1}{2}[5+4-1+0] = 4\right]$$

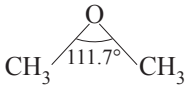
38. (*d*) PH_3, $\theta = 94^o$

39. (*b*) $\widehat{FClF}, \widehat{FBrF}$ as electronegativity of central atom increases angle of molecule increases.

40. (*c*) Thus for each P–3 P–P–P angle

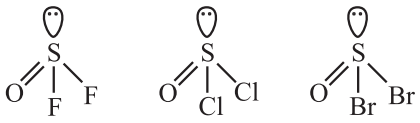


$\therefore$ for 4 it is 12

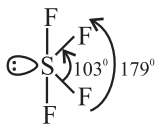
41. (*c*) 

42. (*c*) Bond angle of hydrides of group 16 elements are
$H_2O > H_2S > H_2Se > H_2Te$

43. (*a*) H_2O (104.5^0) ; BeF_2 (180^0)
CH_4 (109^0) ; NH_3 (107^0)

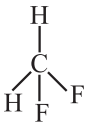
44. (*b*) 

In $SOBr_2$, S–O bond has maximum bond length in comparision to S–O bond lengths in SOF_2 and $SOCl_2$, because in $SOBr_2$, S–O bond has been formed by hybrid orbital containing less s-character. According to Bent Rule.

45. (*a*) 

Equatorial bond angle has been reduced from 120° to 103°

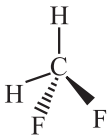
46. (*a*) The hybrid orbitals used for forming C–H bonds contain more s–character than hybrid



orbitals used for forming C–F bonds. As more the s–character in hybrid orbital larger will be bond angle.

47. (*c*) In H–C≡C–H, the C–H bond is made of sp-s overlap which imparts shorter bond length.

48. (a) As the electro-negativity of the central atom decreases and the size increases, the bond angle decrease in similar molecules.

49. (b) 

Total number of angles are six.

$\widehat{HCH}, \widehat{HCF}, \widehat{FCF}, \widehat{HCF}, \widehat{HCF}, \widehat{HCF}$

Maximum number of identical angles are four ($\widehat{HCF}$)

50. (*b*) Due to absence of lone pair on N in NH_4^+, lp–lp, or lp –bp repulsion is not there.

51. (*a*) O is more electronegative than N & S take greater strength of H-bonds as the strength of H-bonds is attributed to the electro-negativity of the atom bonded to hydrogen.

52. (*b*) In H_2S, due to low electronegativity of sulphur the l.p.-l.p. repulsion is more than b.p. -b.p. repulsion and hence the bond angle is 92°.

53. (*d*) $:\ddot{O}=\ddot{O} \longrightarrow \ddot{O}:$ is one of the resonating structure of ozone.

54. (*c*) Hint: Due to maximum lp-bp repulsion present in H_2O.

55. (a) The order of bond angles
$BF_3 > SiH_4 > NH_3 > H_2S$
120° 109°28' 107° 92.5°

56. (*d*)

57. (*c*) Bond angles are:
BF_3 – 120°, NF_3 – 106°, PF_3 – 101°, ClF_3 – 90°

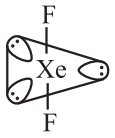
58. (*d*) Bond angle of water is 104.5. In water, oxygen is sp^3 – hydridised with two lone pairs and two bond pairs and lp - lp repulsions is more than l.p - b.p. repulsion, which in turn is more than bp - bp repulsion.

59. (*d*) O–O bond length in O_2 is 120.7 pm, O_3 is 127.8 pm, H_2O_2 is 147.5 pm

60. (*a*)

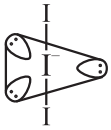
61. (*d*) Shape of the molecule depends upon the total number of electron pairs around the central atom which include l.p as well as b.p.

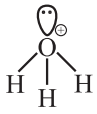
62. (*b*) SF_4 has see-saw shape due to 1 lp and 4 bp and XeF_4 has square planar shape due to 2 lp and 4 bp.

63. (*b*) a. 

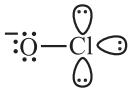
$O = C = O$

Hybridisation : sp^3d Hybridisation : sp
shape : linear shape : linear

b. 

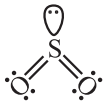


Hybridisation : sp^3d Hybridisation : sp^3
shape : linear shape : linear

c. 

$:C \equiv O:$

Hybridisation of Cl : sp^3 Hybridisation of C : sp
shape : linear shape : linear

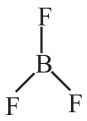
d. 

Hybridisation : sp^2
shape : V-shape/bent/angular



Hybridisation of O : close to sp^3
shape : V-shape/bent/angular

64. (*d*) 



Tetra atomic

$\mu_d \neq 0$ $\mu_d = 0$

Planar & Polar Planar & non-polar

65. (*d*) Values of dipole moment can't be predicted as we don't know about the electronegativities of the atoms.

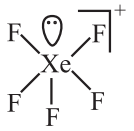
66. (*c*) To have minimum repulsions, the two lone pair occupy the trans positions in octahedral geometry.

67. (*c*) Hybridisation of SO_4^{2-}

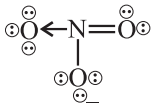
$$H = \frac{1}{2}[6 + 0 - 0 + 2] = 4$$

68. (*d*) a. XeO_3 is trigonal pyramid.
b. IOF_4^+ is see-saw.
c. PCl_5 is trigonal bipyramidal.

d. 



square pyramidal.

69. (*d*) 

70. (*c*) sp^3d^3

71. (*b*) Polarity of molecule decreases with decrease in dipole moment. The dipole moment of HF, H_2O, SO_2 and NH_3 are 1.92 D, 1.84 D, 1.60 D, and 1.4 D respectively.

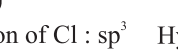
72. (*b*) 3 B.P and 1 L.P

73. (*d*) sp^3- Tetrahedral.

74. (*b*) CO_2 and BeF_2 both are linear.

75. (*c*) Regular geometry of CH_4 is attributed to zero dipole moment ($\mu = 0$) and absence of lone pair on carbon atom.

76. (*b*) OF_2 has 2 bond pairs and 8 lone pairs.



77. (*c*) Both XeF_4 and $[Ni(CN)_4]^{2-}$ are square planar. In XeF_4, Xe is sp^3d^2 hydridised with 2 lp and Ni in $[Ni(CN)_4]^{2-}$ is dsp^2 hybridised.

78. (*b*) In aldehydes hybridisation of carbon = sp^2

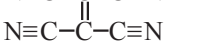
79. (*a*) s-character = 20%.

80. (*b*) In decane $4 \times 10 = 40$ hybrid orbitals present.

81. (*c*) C-C bond is $sp^3 - sp^3$ and C-Cl is $sp^3 - p$

82. (*d*) sp^3 hybridization, hence s - character is 25%

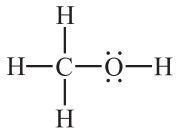
83. (*c*) $C_2(CN)_4$ is

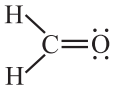


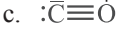
C = C is sp^2 hybridised and C ≡ N is sp hybridised.

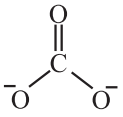
84. (*b*)

85. (*b*) N ≡ C –C ≡ N

86. (*c*) a. 

b. 

c. 

d. $2[Na^+]$ 

c. → triple bond therefore bond length shortest.

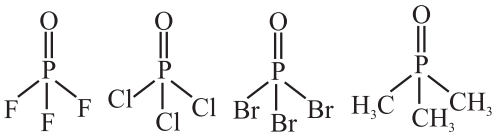
87. (*c*) $\sigma = 12$, $\pi = 3$

88. (*a*) s - p bond involves head-on overlapping. Therefore; it is a sigma (σ) bond.

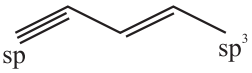
89. (*b*) S = C = S (linear) is best explained by sp hybrid state of C.

90. (*b*) a σ bond is formed by axial overlapping.

b. p-orbital have both axial and side ways overlapping

91. (*a*) 

In OPF_3, molecule % s-character in O – P is maximum according to Bent's Rule, hence it is strongest among all.

92. (*d*) 

93. (*a*) a. $\underset{sp^3}{CH_4} + 2O_2 \rightarrow \underset{sp}{CO_2} + 2H_2O$

94. (*b*) Do it yourself

95. (*c*) XeF_5^- : 2 lone pairs ; pentagonal planar

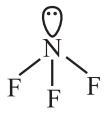
BrF_3 : 2 lone pairs ; Bent-T-shape

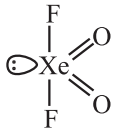
XeF_2 : 3 lone pairs ; Linear

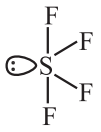
H_3S^+ : 1 lone pairs ; Pyramidal

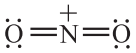
$\overset{\uparrow\uparrow}{CH_2}$: 2 unpaired ; Linear

(No lone pair)

96. (*b*) 

sp^3 Hybridisation  sp^3d Hybridisation



sp^3d Hybridisation  sp Hybridisation

97. (*b*) $\underset{sp^3}{NH_3} + \underset{}{H^+} \rightarrow \underset{sp^3}{NH_4^+}$ $\quad 2\underset{sp^3}{NH_3} \rightarrow \underset{sp^3}{NH_2^-} + \underset{sp^3}{NH_4^+}$

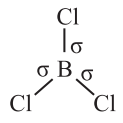
$\underset{sp^2}{H_3BO_3} + \underset{sp^3}{OH^-} \longrightarrow \underset{sp^3}{[B(OH)_4]^-}$ $\quad \underset{sp^3}{H_2O} + \underset{}{H^+} \rightarrow \underset{sp^3}{H_3O^+}$

98. (*b*) N_2O_5 in solid state $\rightarrow NO_2^+\ NO_3^-$

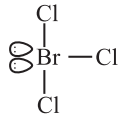
$NO_3^- \longrightarrow$ hybridisation $\longrightarrow sp^2$

99. (*d*) If number of B.P. and L.P are same for the given pairs, they are isostructural.

(*a*)

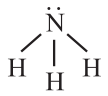


3 σ-bonds
sp^2-hybridization &
trigonal planar geometry

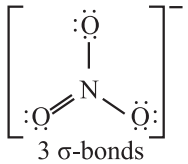


3 σ-bonds + 2L.P.
sp^3d-hybridization and T-shape geometry

(*b*) NH_3 and NO_3^-

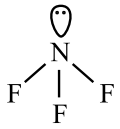


3 σ-bonds + 1 L.P.
sp^3-hybridization and
pyramidal geometry

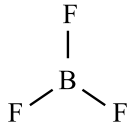


3 σ-bonds
sp^2-hybridization and
trigonal geometry

(*c*) NF_3 and BF_3

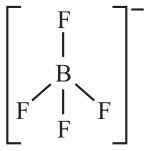


3 σ-bonds + 1 L.P.
sp^3-hybridization and
pyramidal geometry

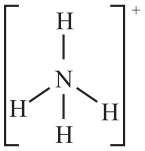


3 σ-bonds
sp^2-hybridization and
trigonal planar geometry

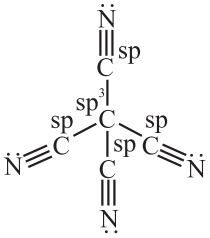
(*d*) BF_4^- and NH_4^+

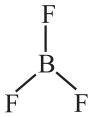


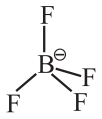
4 σ-bonds
sp^3-hybridization
tetrahedral geometry



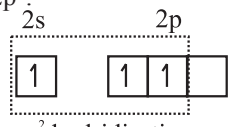
4 σ-bonds,
sp^3-hybridization
and tetrahedral geometry

100. (*b*) 

101. (*a*) 

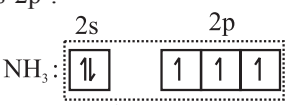
sp^2  sp^3

102. (*c*) (*a*) Electronic configuration of Beryllium in ground state is $1s^2 2s^2 2p^1$.



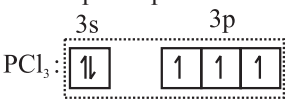
sp^2 hydridisation

(*b*) Electronic configuration of nitrogen in ground state is $1s^2 2s^2 2p^3$.



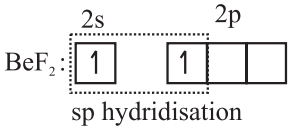
sp^3 hydridisation

(*c*) Electronic configuration of phosphorus in ground state is $1s^2 2s^2 2p^6 3s^2 3p^3$.



sp^3 hydridisation

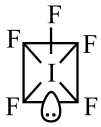
(*d*) Electronic configuration of boron in ground state is $1s^2 2s^2$.



103. (*c*) $IF_5 = 7 + 5 \times 7 = \frac{42}{8} = 5+1$

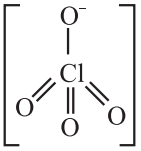
sp^3d^2 hybridisation, hence it should be octahedral.

Due to the presence of one lone pair of electrons, its actual geometry is square pyramidal.



104. (*b*) A fact, to be remembered.

105. (*d*) In perchlorate ions,
Central Cl atom is sp^3 - hybridised. So; all the four sigma bonds are sp^3 - p bonds.



106. (*c*) a. $CH_4 \rightarrow CH_3—CH_3$
4 b.p 3 b.p 4 b.p
Hybrisation: Tetrahedral Tetrahedral

b. $NH_3 \rightarrow NH^+_4$
3 b.p +1 l.p 4 b.p
Hybrisation: Pyramidal Tetrahedral

c. $BF_3 \rightarrow BF_4^-$
Hybridisation: Tetrahedral trigonal planar

d. $H_2O \rightarrow H_3O^+$
2b.p.+2 l.p 3b.p. + 1 l.p
Hybridisation: Angular Pyramidal

Thus, conversion of BF_3 into BF_4^- involves changes in both hybridisation and shape.

107. (*d*) Perfectly tetrahedral $\mu_D = 0$ non polar



sp^3-hybridised

108. (*a*) Order of overlapping is
$\sigma_{p-p} > \sigma_{s-p} > \sigma_{s-s}$

109. (*a*) hybrid orbitals form sigma bonds only.

110. (*c*) pi bond is formed by the lateral overlaping.

111. (*a*) sigma bond is stronger than pi bond.

112. (*d*) sp^3d^2

113. (*d*) In propane carbon undergoes sp^3 hybridisation.

114. (*d*) Both PO_4^{3-} & SO_3^{2-} have hybridisation sp^3.

115. (*b*) In propene 10 hybrid orbitals (3+3+4) are present.

116. (*a*) In sp^3d^2 hybridization, angle between two hybrid orbital is 90°.

117. (*b*) p-character = 75%

118. (*b*) In hybridisation combination and redistribution of atomic orbitals takes place.

119. (*b*) sp hybridisation is formed.

120. (*a*) bond angle $\propto$ s- character.

121. (*c*) sp^3d hybridization gives TBP.

122. (*d*) s-orbital of hydrogen and p-orbital of fluorine overlaps.

123. (*a*) Overlapping of atomic orbitals determines direction of the bond.

124. (*a*) sigma bond is directional.

125. (*a*) Order of overlapping is
$\sigma p - p > \sigma s - p > \sigma s - s$, 1s-1s bond is stronger because 1s orbital is close to the nucleus.

126. (*b*) Do it yourself

127. (*d*) 104° 28

128. (*a*) In H_2O, hybridization is sp^3 and the shape is angular due to 2 bp & 2 lp

129. (*c*) In $BeCl_2$, hybridization sp kind the shape is linear due to 2 bp & 0 lp

130. (*b*) $NH_4^+ \; T.P = \frac{5+4-1}{2} = 4\left(sp^3\right)$

131. (*d*) sp^3d has bond angles 90° & 120°

132. (*c*) All of them have total four electron pairs with 0 lp . Hence tetrahedral in shape.

133. (*d*) CH_3CH_2OH is saturated hence all carbon atoms are sp^3 hybridized.

134. (*d*) Ethylene on hydrogenation into ethane : hybradization , bond angle & C - C bond length change

135. (*c*) Do it yourself

136. (*d*) $\pi^*2p_x^2.\pi^*2p_y^1$ antibonding molecular orbitals contain 3 electrons in O_2^-. There energy being more , ionisation potential is less.

137. (*c*) KO_2 has K^+ and O_2^- structure having one unpaired electron.

138. (*b*) $B.O.\ of\ \ O_2 = 2$

$O_2^+ = 2.5$

$O_2^- = 1.5$

139. (*b*) Both O_2^+ and O_2 are paramagnetic bond order of $O_2^+ = 2.5$ and $O_2 = 2$

140. (*b*) (a) $C_2 \longrightarrow C_2^+$
B.O.= 2 B.O. = 1.5
$\mu = 0$, $\mu \neq 0$,
i.e. diamagnetic i.e. paramagnetic

(b) NO $\longrightarrow$ NO^+
B.O.= 2.5 B.O. = 3.0
$\mu \neq 0$, $\mu = 0$,
i.e. paramagnetic i.e. diamagnetic

(c) $O_2 \longrightarrow O_2^+$
B.O.= 2 B.O. = 2.5
$\mu \neq 0$, $\mu \neq 0$,
i.e. paramagnetic i.e. paramagnetic

(c) $N_2 \longrightarrow N_2^+$
B.O.= 3 B.O. = 2.5
$\mu = 0$, $\mu \neq 0$,
i.e. diamagnetic i.e. paramagnetic

141. (*c*) $\sigma 1s^2\ \sigma^*1s^1$

142. (*b*) It does not have any π - bond because all the molecules have O - atom with lone pairs but in H_2O, the H - atom has no vacant orbital for π - bonding.

143. (*a*) O_2^{2-} (Number of electron = 18 does not contain unpaired electron)
$O_2^{2-} = \sigma 1s^2 , \sigma^*1s^2 , \sigma 2s^2 , \sigma^*2s^2 , \sigma 2p_z^2,$
$\pi_2 p_x^2 \approx \pi_2 p_y^2\ \pi^*2p_x^2 = \pi^*2p_y^2$

144. (*b*) Since anti-bonding orbitals are formed by substraction of orbitals.

145. (*c*) Bond order of Be_2 is zero, it does not exist.

146. (*c*) According to MOT oxygen is paramagnetic with two unpaired electrons.

147. (*a*) NO^- has 16 electron system bond order = 2
NO^-, CN^-, N_2 are isoelectronic with CO.
$\therefore$ All have same bond order = 3

148. (*a*) Bond order $\propto$ bond length.
B.O of $O_2^+ = 2.5$ and rest has B.O. less than 2.5
Thus, it has the strongest O – O bond among the given.

149. (*c*) Total number of electrons in $N_2^- = 7 + 7 + 1 = 15$
Total number of electrons in $N_2^{2-} = 7 + 7 + 2 = 16$
Total number of electrons in $O_2^{2-} = 8 + 8 + 2 = 18$
Total number of electrons in $O_2^- = 8 + 8 + 1 = 17$
Total number of electrons in $O_2^+ = 8 + 8 - 1 = 15$

Configuration will be same as N_2^- and O_2^+ contain the same total number of electrons. Thus N_2^- and O_2^+ has the same bond order.

150. (*b*) $B_2 = KK * (\sigma 2s)^2 (\sigma^*2s)^2 (\pi p_x)^1 \approx (\pi p_y)^1$
Bond order $= \frac{1}{2}(4-2) = 1$

151. (*b*) C (12) $= \sigma 1s^2 , \sigma^*1s^2 , \sigma 2s^2 , \sigma^*2s^2 , \pi 2p^2{}_x \approx \pi 2p_y{}^2$
C_2 has two pairs of electrons in $\pi 2p_x$ and $\pi\, 2p_y$, therefore; it has only 2π - bonds.

152. (*c*) No unpaired electrons, therefore diamagnetic.
eg: $Li_2 = 6e^- = \sigma 1s^2 , \sigma^*1s^2 , \sigma 2s^2$

153. (*c*) Mulliken gave the molecular orbital theory.

154. (*a*) Given electronic configuration of anion X is-
$\sigma 1s^2 , \sigma^*1s^2 , \sigma 2s^2 , \sigma^*2s^2 , \pi 2p^2{}_x \approx \pi 2p_y{}^2 , \sigma 2p_z{}^2 , \pi^*2p_x{}^1$
$\therefore$ Total number of electron of anion X = 15
Hence; the anion X is N_2^-

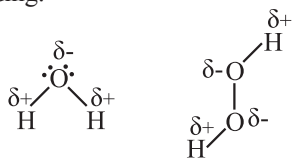
155. (*a*) In the anti-bonding molecular orbital, most of the electron density is located away from the space between the nuclei, as a result of which there is a nodal plane (i.e., a plane at which the electron density is zero) between the nuclei.

156. (*b*) According to Molecular Orbital Theory, $H_2^+ = \sigma 1s^1$
Bond order $\frac{N_b - N_a}{2} = \frac{1-0}{2} = 0.5$

157. (*a*) Isostructural species will have same bond order

158. (*b*) 1.5 **159.** (*b*) N_2 **160.** (*c*)

161. (*b*) More is the partial charges generated more is the strength of H-bonding.



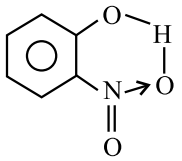
The strength of H-bonding is more in H_2O.

162. (*c*) F has highest EN.

163. (*d*) o-Salicylaldehyde has intramolecular H-bonding.

164. (*b*) Ammonia has hydrogen bonds

165. (*b*) Ortho-nitrophenol



166. (*a*) H-bond by simply vander waal's attraction.

167. (*d*) Because of high EN of 'O' forms intermolecular H-bonding.

168. (*d*) Bromine has no hydrogen bonds.

169. (*a*) Ethyl alcohol has hydrogen bonds.

170. (*a*) Due to +I effect. An electron-releasing group increases electron density on oxygen tending to decrease the polarity of O—H bone. This decreases the Hydrogen bond strength.

171. (*b*) C_2H_5OH can form intermolecular H-bonding with H_2O.

172. (*a*) Because base units in DNA has H,N & O as component elements.

173. (*b*) Because o-nitrophenol has intramolecular H-bonding.

174. (*d*) More number of H-bonds

175. (*a*) Only O,N,F (because of high EN) can form H-bonding.

176. (*a*) Ionic bond > covalent bond > Hydrogen bond > vander waals forces > metallic bond.

177. (*c*) NH_3 has intermolecular H-bonding.

178. (*b*) 2 lone pairs of 'O' and 2'H' atoms in H_2O molecule involve in H-bonding

179. (*a*) Ice occupies more volume than water

180. (*b*) Intermolecular H-bonding is formed between H_2O molecules.

181. (*c*) In aqueous ammonia hydrogen bonding is present between NH_3 & H_2O .

182. (*d*) None

183. (*b*)

184. (*d*) Hydrogen bond is formed between molecules of compounds having O, F and N with hydrogen.

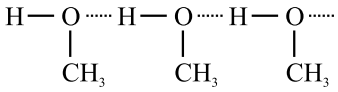
185. (*c*) Ethyl alcohol forms stronger H - bonds than ethylamine or ammonia due to greater electronegativity of oxygen than nitrogen atom. Diethyl ether, however, does not form H-bonds, since it does not have a H-atom attached to any highly electronegative atom.

186. (*d*) The higher boiling point of water is due to hydrogen bonding. O-atom of each H_2O molecule is covalently linked with two H-atoms of its own molecule and with another H-atom of adjacent H_2O molecule by H-bonding.

187. (*d*) Ice has H-bonding therefore; when two ice-cubes are pressed over each other, they unite to form one cube.

188. (*c*) In ethanol, hydrogen bonding is maximum.

189. (*a*) The molecule of CH_3OH are associated by hydrogen bonding.

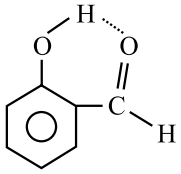


190. (*a*) Water (H_2O) is least volatile due to the association of H_2O molecules through H-bonding.

191. (*d*) Strength of hydrogen bond between H----F is the strongest and hence; it has the strongest bond.

192. (*c*) Intermolecular H - bonding is found in $(HF)_n$ due to higher electro-negativity of fluorine atoms. Hydrogen bonding is helpful in the association of HF molecule. So, HF form in liquid form.

193. (*b*)



Learning Plus

1. (*b*) **2.** (*a*)

3. (*c*) CO_2 being symmetrical has zero dipole moment
Among HI, SO_2 and H_2O dipole moment is highest for H_2O as in it the central atom contains 2 lone pairs.

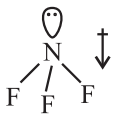
4. (*c*) **5.** (*a*)

6. (*b*) Strength of H-bond is in the order H....F > H…O > H…..N.

But each H_2O molecule is linked to four other H_2O molecules through H-bonds whereas each HF molecule is linked only to two other HF molecules. Hence, b.p of H_2O > b.p of HF > b.p. of NH_3

7. (*c*) (*a*) NH_3 bond angle is 107.5°

(*b*)



(*c*) Percent 's' character $= \frac{1}{3} \times 100 = 33.33\%$

Percent 'p' character = 66.67 %

8. (*b*)

9. (*c*) In PO_4^{3-} ion, formal charge on each O-atom of P–O bond

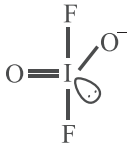
$$= \frac{\text{total charge}}{\text{Number of O} - \text{atom}} = -\frac{3}{4} = -0.75$$

10. (*d*) In N-atom, number of valence electrons = 5
$2e^-$ of N atom are covalently bonded with one O atom, other $2e^-$ of N atom are involved in coordinate bond with other O atom, one e^- lest of N atom is covalently bonded with the 3rd O atom & due to the -ve charge present on O atom octet is completed.
Thus, 3 O-atoms shared with 8 electrons of N-atom.
∴ Number of bond pairs (or shared pairs) = 4
Number of lone pairs = 0

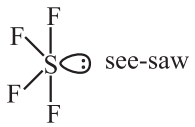
11. (*a*) A fact, to be remembered.

12. (*c*)

13. (*d*)



sp^3d but see-sawshape



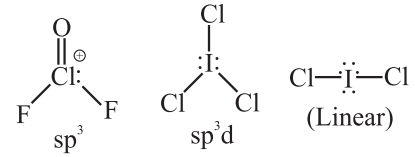
see-saw

14. (*a*) (BH_4^- ⇒ 4 bond pairs + 0 lone pair ⇒ sp^3 hybridised = tetrahedral geometry)
(CO_3^{2-} = triangular planar
NH_2^- = V – shape
H_3O^+ = pyramidal)

15. (*c*) The given compound will have the correct structure as 19 σ-bonds are present in the given molecule i.e., there are 5 π-bonds and 8 C–H + 11 C–C σ-bonds.

16. (*c*) NO^+_2 involves an sp type of hybridization.

17. (*a*)



(Pyramidal) (sp^3) (T-shaped) (sp^3d) (Linear)

18. (*d*) Cl – Be – Cl $O = \overset{+}{N} = O$



19. (*c*) The electronic configuration of the given molecules are

$N_2^+ = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi p_x^2 = \pi p_y^2, \sigma 2p_z^1$

It has one unpaired electron.

$O_2 = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 \approx \pi 2p_y^1 \pi^* 2p_z^1 = \pi^* 2p_y^1$

O_2 has two unpaired electrons.

$O_2^{2-} = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 \approx \pi 2p_y^2, \pi^* 2p_x^2 \approx \pi^* 2p_y^2$

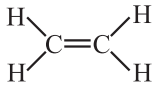
Thus, O_2^{2-} has no unpaired electrons.

$B_2 = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2 \pi 2p_x^1 \approx \pi 2p_y^1$ Thus, B_2 has two unpaired electrons.

20. (*c*) $XeF_4 \Rightarrow 4bp + 2\ell p$ square planar $\Rightarrow$ all bonds are equal

$BF_4^- \Rightarrow 4bp + 0\ell p \Rightarrow$ tetrahedral (all bonds are equal)

SiF $\Rightarrow 4bp + 0\ell p \Rightarrow$ tetrahedral (all bonds are equal)

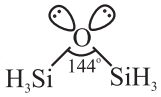


Thus, in C_2H_4 all the bonds are not equal.

21. (*b*) In $\overset{\cdot\cdot}{N}H_3$ central atom is nitrogen which is sp^3 hybridized hence it will be at the centre of tetrahedron with H–atoms at three vertices.

22. (*a*) In trigonal bypyramidal geometry lone pair electrons cannot occupy axial positions – (VSEPR theory)

23. (*c*) $O(SiH_3)_2$



24. (*c*) H_2S– No hybridisation bond angle ≈ 93°

NH_3 – Pyramidal 104.5°

CH_4 – Tetrahedral 109°28'

BF_3 – Trigonal plannar120°

25. (*b*) As the electronegativity of central atom increases the bond angle increases due to repulsion between bond pair and bond pair as bond pairs are more close to the central atom.

26. (*d*) The given belongs to d-block of the periodic table. In transition elements i.e., d-block elements, electrons of ns and (n – 1) d/ subshell take part in bond formation electronic configuration shows that the element is titanium (Z = 22).

27. (*b*) For sp^2 hybridization, the geometry is generally triangular planar.

Thus, bond angle is 120°.

28. (*a*) As electronegativity of P is much lower than N, so in NH_3

Nitrogen attracts all bond electron towards the centre and electron-electron repulsion is higher so it has a higher bond angle. Bond angle in ammonia is 107° & that in phosphene is 93°.

29. (*b*) C_2 : $(\sigma 1s)^2 < (\overset{*}{\sigma} 1s)^2 < (\sigma 2s)^2 < (\overset{*}{\sigma} 2s)^2 < (\pi_z^2 px = \pi_z^2 py)\ (2py)\Big]$

$\mu_{eff} = 0$; i.e. diamagnetic

B_2 : $(\sigma 1s)^2 < (\overset{*}{\sigma} 1s)^2 < (\sigma 2s)^2 < (\overset{*}{\sigma} 2s)^2 < (\pi_{2px})^1$

$= (\pi_{2py})^1 (\pi_{2px}) = (\pi_{2py})$

$\mu_{eff} \neq 0$; i.e. paramagnetic

O_2 : $(\sigma 1s)^2 < (\overset{*}{\sigma} 1s)^2 < (\sigma 2s)^2 < (\overset{*}{\sigma} 2s)^2 < (\sigma 2pz)^2$

$< (\pi_{2px})^2 = (\pi_{2py})^2 < (\overset{*}{\pi} 2px)^1 = (\overset{*}{\pi} 2py)^1$

$O_2 : \mu_{eff} \neq 0$

$O_2^{2+} : \mu_{eff} = 0$

$O_2^{2-} : \mu_{eff} = 0$

N_2 :

$(\sigma 1s)^2 < (\overset{*}{\sigma} 1s)^2 < (\sigma 2s)^2 < (\overset{*}{\sigma} 2s)^2 < (\pi_{2px})^2$

$= (\pi_{2py})^2 < (\sigma 2pz)^2$

N_2 : $\mu_{eff} = 0$

N_2^+ : $\mu_{eff} \neq 0$,

N_2^{2-} : $\mu_{eff} \neq 0$

N_2^{2+} : $\mu_{eff} = 0$

30. (*a*) The correct increasing order of energies of molecular orbitals of N_2 is given below

$\sigma 1s^2 < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x \approx \pi 2p_y) < \sigma 2p_z$

$(\pi^* 2p_x \approx \pi^* 2p_y) < \sigma^* 2p_z$

31. (*d*) Existence of molecule, bonding nature and energy order of molecular orbitals can be explained on the basis of molecular orbital theory as follows

(*i*) Molecules having zero bond order never exists while molecular having non-zero bond order either exists or expected to exist.

(*ii*) Higher the value of bond order, higher will be its bond strength.

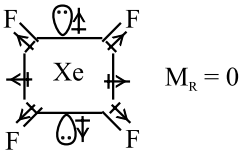
Electrons present in bonding molecular orbital are known as bonding electrons (N_b) and electrons present on anti-bonding molecular orbital are known as anti-bonding electrons (N_a) and half of their difference is known as bond order.

(*d*) is incorrect. The correct order of energies of molecular orbitals in N_2 molecule is

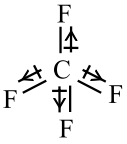
$\sigma 2s < \sigma^* 2s < (\pi 2p_x \approx \pi 2py) < \sigma 2p_z < \pi^* 2p_x \approx \pi^* 2p_y < \sigma^* 2p_z$

32. (*a*) In KO_2 , O_2^- (superoxide ion) has one unpaired electron.
In BaO_2, O_2^{2-} (peroxide ion) has paired electrons.
In AlO_2^- , $[Al^{3+}\ O_2^{2-}\ ]^-$, O_2^{2-} (peroxide ion) has paired electrons

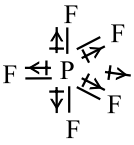
33. (*c*) XeF_4 is square planar in shape.



$M_R = 0$



$M_R = 0$



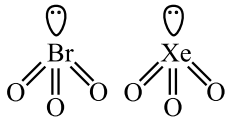
$M_R = 0$

34. (*d*) PF_5 exists as PF_4^+ PF_6^-
(s) sp^3 sp^3d^2

PCl_5 exists as PCl_4^+ PCl_6^-
(s) sp^3 sp^3d^2

PBr_5 exists as PBr_4^+ PBr_6^-
(s) sp^3 sp^3d

PBr_5
(g) sp^3d

35. (*a*) BrO_3^- 

36. (*c*) Only in BF_3 pπ-pπ back bonding is possible due to which B–F bond length is shortest. In other compounds, p-orbital at boron is not vacant hence, pπ-pπ back bonding is not possible.

Multiconcept MCQs

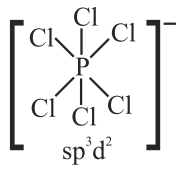
1. (*d*)

2. (*c*)

3. (*d*) (*i*) 

sp^3d

(*ii*) $\left[N=\overset{+}{N}=N\right]^{2-}$
sp

(*iii*) PCl_5 (s) exists as $[PCl_4]^+$ and $[PCl_6]^-$, 

sp^3d^2

(*iv*) $I_2Cl_6(\ell) \rightleftharpoons [ICl_2]^+ + [ICl_4]^-$
self ionisation

4. (*a*)

5. (*c*) XeF_5^- : 2 lone pairs ; pentagonal planar
BrF_3 : 2 lone pairs ; Bent-T-shape
XeF_2 : 3 lone pairs ; Linear
H_3S^+ : 1 lone pairs ; Pyramidal
$\overset{\uparrow\uparrow}{CH_2}$: 2 unpaired ; Linear
(No lone pair)

6. (*c*)

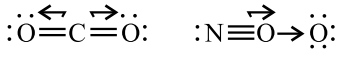
7. (*d*) N_2^+ has one unpaired electron therefore it is paramagnetic and not diamagnetic

8. (*b*) (*i*) (T)
(*ii*) (F) In general as the number of lone pair of electrons on central atom increases, value of bond angle from normal bond angle decreases due to *l*p – *l*p > *l*p – bp
(*iii*) (T)
(*iv*) (F) Structures of Xenon fluorides and xenon oxy fluoride are explained on the basis of VSEPR theory.

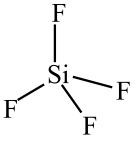
9. (*c*)

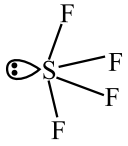
10. (*c*)

11. (*c*) Both CO_2 and N_2O are linear in shape. CO_2 is non-polar but N_2O is polar



12. (*a*)

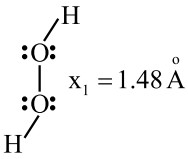
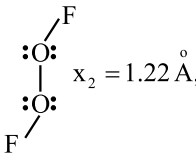
13. (*b*) 



Si is sp^3 hybridised in SiF_4 and it has tetrahedral shape
S is sp^3d hybridised and SF_4 has bent see-saw shape. So, not isostructural.

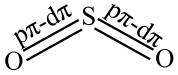
14. (*c*) Linear dipole is created in p-nitroaniline, dipole moment in opposite direction cancel out.

15. (*a*) O_2 have two unpaired electron, O_2^+ have one unpaired electron, O_2^{2-} have zero unpaired electron. As number of unpaired electrons increases paramagnetic character increases.

16. (*b*)  $x_1 = 1.48\,\text{Å}$  $x_2 = 1.22\,\text{Å}, x_1 > x_2$

According to Bent's Rule in O_2F_2, there is more p-character in O–F bond in comparison to O–H bond in H_2O_2 ; hence s-character in O–O bond is greater in O_2F_2.

17. (*a*) $R = \sqrt{P^2 + Q^2 + 2PQ\cos\theta}$; as θ increases from 90°, cosθ becomes –ve. So R ↓.

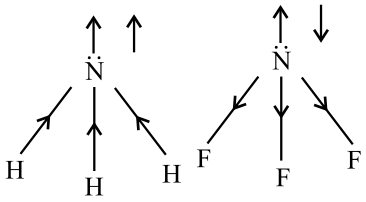
18. (*d*) 

SeF_4 has sp^3d hydridisation and see-saw shape I_3^+ has sp^3 hybridisation and bent shape.

19. (*b*) Bond order = $\frac{1}{2}(N_b - N_a)$

Bond order of $Li_2 = 1$, $N_2 = 3$, $Be_2 = 0$, $O_2 = 2$

20. (*c*) A is a VII A group and C is a IA group element. ∴ The compound is ionic.

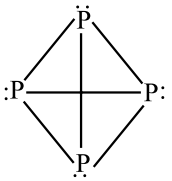
21. (*d*) 

22. (*b*) (I) $[PCl_4]^+ \rightarrow sp^3$



(II) H H has 104.5° bond angle due to lp-lp repulsion.

(III) All have sp^3 hybridisation and one lone pair.

(IV) 

23. (*a*) % s-character ∝ bond angle

$$\propto \frac{1}{\text{bond length}} \propto \text{bond strength}$$

24. (*d*)

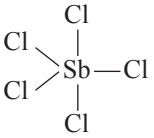
NEET Past 10 Years Questions

1. (*b*) 

sp^2, Trigonal planar 6e– around central atom.

2. (*d*)

	Hybridisation	L.P.	Shape
(A) PCl_5	sp^3d	0	Trigonal bipyramidal
(B) SF_6	sp^3d^2	0	Octahedral
(C) BrF_5	sp^3d^2	1	Square pyramidal
(D) BF_3	sp^2	0	Trigonal planar

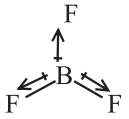
3. (*b*) 

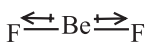
sp^3d Dipole moment (μ) = 0

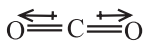
Trigonal bipyramidal Non-polar

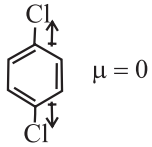
4. (*c*)

The given set of molecules have dipole moment zero. Because dipoles of the bond cancel each other.

 μ = 0

 μ = 0

 μ = 0

 μ = 0

5. (*d*)

For He_2 molecule

Electronic configuration is $\sigma 1s^2$, $\sigma^* 1s^2$

$$B.O = \frac{1}{2}(N_b - N_a)$$

$$= \frac{1}{2}(2-2)$$

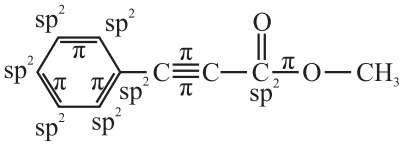
$$= 0$$

The bond order comes out to be zero. This indicates that there is no bond formation between 2 He atoms and hence the He_2 molecule does not exist.

6. (*b*)

(*i*) Number of sp^2 hybridised carbon atoms is 7

(*ii*) Number of pi bonds is 6



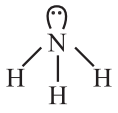
7. (*d*)

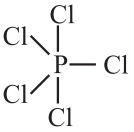
Potential energy of two H atoms at infinite distance = a

Potential energy of two H atoms at distance equal to bond length = b

So, The bond energy of H_2 = (b – a)

8. (*d*)

NH_3  Trigonal pyramidal

PCl_5  Trigonal bipyramidal

SF_6  Octahedral

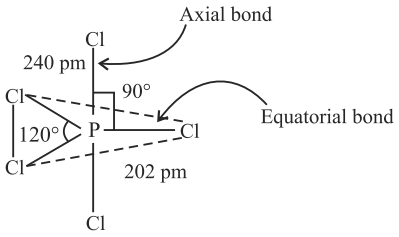
$BeCl_2$ Cl—Be— Cl Linear

9. (*c*)

$$\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_x^2 = \pi 2p_y^2$$

Double bond in C_2 consists of both π bonds because of the presence of four electrons in two π molecular orbital. In other molecule a double bond is made up of a sigma bond and a pi bond.

10. (*d*)



(*d*) False

Due to longer and thus weaker axial bonds, PCl_5 is a reactive molecule.

11. (*b*)

NO : $(\sigma 1s)^2$, $(\sigma^* 1s)^2$, $(\sigma 2s)^2$, $(\sigma^* 2s)^2$, $(\sigma 2p_z)^2$, $(\pi 2p_x)^2 = (\pi 2p_y)^2$, $(\pi^* 2p_x)^1 = (\pi^* 2p_y)^0$

$$BO = \frac{10-5}{2} = 2.5$$

CN^- : $(\sigma 1s)^2$, $(\sigma^* 1s)^2$, $(\sigma 2s)^2$, $(\sigma^* 2s)^2$, $(\pi 2p_x)^2 = (\pi 2p_y)^2$, $(\sigma 2p_z)^2$

$$BO = \frac{10-4}{2} = 3$$

CN : $(\sigma 1s)^2$, $(\sigma^* 1s)^2$, $(\sigma 2s)^2$, $(\sigma^* 2s)^2$, $(\pi 2p_x)^2 = (\pi 2p_y)^2$, $(\sigma 2p_z)^1$

$$BO = \frac{9-4}{2} = 2.5$$

CN^+ : $(\sigma 1s)^2$, $(\sigma^* 1s)^2$, $(\sigma 2s)^2$, $(\sigma^* 2s)^2$, $(\pi 2p_x)^2 = (\pi 2p_y)^2$

$$BO = \frac{8-4}{2} = 2$$

Hence, option (*b*) should be the right answer.

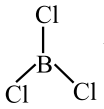
12. (*d*)

$CN^- = 6 + 7 + 1 = 14$

$CO = 6 + 8 = 14$

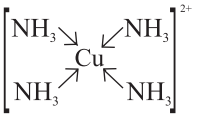
These two species are isoelectronic and iso structural in nature. Therefore, they both have same bond order.

13. (*a*)

BCl_3 having bond angles of 120^0.  VESPR Theory

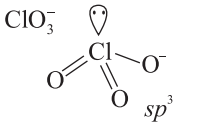
14. (*d*)

$[Cu(NH_3)_4]^{2+}$ is not tetrahedral. It is a square planar complex.



sp^2d

15. (*c*)



sp^3 sp^2

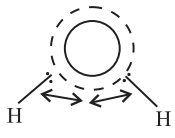
tetrahedral geometry trigonal planar pyramidal shape.

Hence, they both are not isostructural.

16. (*d*)

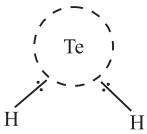
O, Se, Se, Te belong to Group 16.

On moving down the group, size of atom increases.



'O' is most electronegative and lone pairs lie close to the atom electron cloud. This causes repulsion in lone pairs of oxygen and bond pairs of hydrogen.

$\therefore$ Angle maximum due to l.p - b.p repulsion.



Te has maximum size: lone pair lie far away from the atom electron cloud. Lone pair - bond pair repulsion is the least.

17. (*a*)

Intramolecular hydrogen bonding takes place within the same molecule i.e., between the atoms of the same molecule. Since, cellulose is a complex structure containing oxygen and hydrogen the bonding occurs between them easily.

Whereas, in HCN, H_2O_2 and concentrated acetic acid inter-molecular hydrogen bonding occurs.

Eg.: $^{+}H-CN^{-}\cdots\cdots{}^{+}H-CN^{-}$

$\quad\quad\quad

25. (*a*)

Each pi bond have $2e^-$ involved so 4 pi bonds = $8e^-$.

26. (*a*)

Bond order: $O_2^- = 1.5 \quad O_2 = 2 \quad O_2^+ = 2.5$

$O_2^+ > O_2 > O_2^-$

27. (*c*)

Indentation Bond order:

$O_2^{2+} = 3, O_2^- = 1.5, O_2^+ = 2.5 \rightarrow O_2^{2+} > O_2^+ > O_2^-$

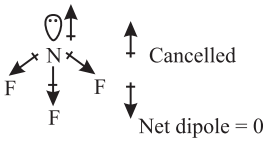
28. (*b*)

29. (*a*)

Species →	NO_3^-	NO_2	NO_2^-	NO_2^+
Hybd. →	sp^2	sp^2	sp^2	sp
Bond angle →	120°	134°	115°	180°
		↓	↓	
		B.P.– B.P. repulsion dominant	L.P. – B.P. repulsion dominant So, Bond angle ↓ So bond angle↑	

30. (*b*)

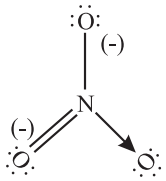




= 0.23D

31. (*a*)

NO_3^- resonating structure



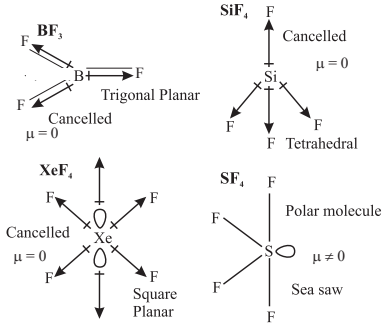
Planar structure

32. (*a*)

Hybridisation of CO_2 is *sp* O = C = O

and of ethyne C_2H_2 is *sp* H – C ≡ C – H

33. (*b*)



34. (*c*)

Among the 4 options $(CH_3)_2$, $(SiH_3)_2$ and PH_3 all have complete octets with $8e^-$. But BH_3 has only $6e^-$ and act as a Lewis acid.

35. (*b*)

In paramagnetic species there is a presence of unpaired electrons.

Except O_2^- all contains paired electrons.

36. (*a*)

By MOT for filling of e^- in molecular orbitals

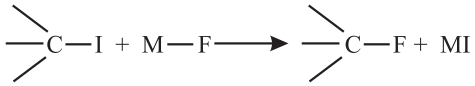
Bond order of $He_2^+ = 0.5, NO = 2.5, O_2^- = 1.5$

$C_2^{2-} = 3$

So, $C_2^{2-} > NO > O_2^- > He_2^+$

37. (*d*)

For the reaction



The reaction will be faster with Rb because lattice energy of RbF is less than LiF, NaF, KF due to large size of Rb^+

38. (*b*)

O_2^{2-} : 1	B_2 : 1
O_2^+ : 2.5	NO^+ : 3
NO : 2.5	CO : 3
N_2 : 3	O_2 : 2

39. (*c*)

$O_2 \Rightarrow \sigma ls^2 \sigma^* ls^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \left(\pi 2p_x^2 = \pi 2p_y^2\right)\left(\pi^* 2p_x^1 = \pi^* 2p_y^1\right)$

O_2 :Bond order $= \frac{1}{2}(8-4) = 2$

O_2^+ :Bond order $= \frac{1}{2}(8-3) = 2\frac{1}{2}$

O_2^- :Bond order $= \frac{1}{2}(8-5) = 1\frac{1}{2}$

O_2^{2-} :Bond order $= \frac{1}{2}(8-6) = 1$

40. (*a*)

BF_4^- & NH_4^+ have sp^3 hybridised central atom and exhibit tetrahedral geometry.

41. (*a*)

Electronic configuration of O_2

$\sigma(1s)^2, \sigma^*(1s)^2, \sigma(2s)^2, \sigma^*(2s)^2, \sigma(2P_z)^2, \pi(2P_x)^2, \pi(2P_y)^2, \pi^*(2P_x)^1 = \pi^*(2P_y)^1$

So, incoming e^- will enter in $\pi^* 2P_x$ / $\pi^* 2P_y$