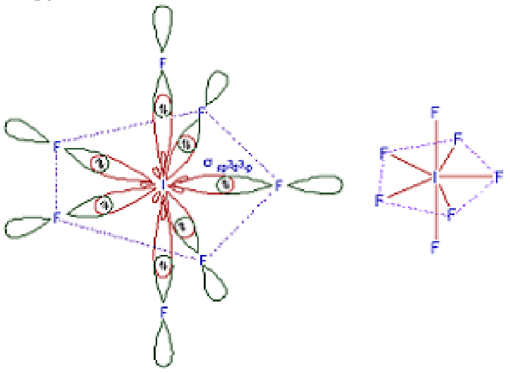
Solution

CHEMICAL BONDING AND MOLECULAR STRUCTURE WS 1

Class 11 - Chemistry

1.

(d) Pentagonal bipyramid

Explanation: 

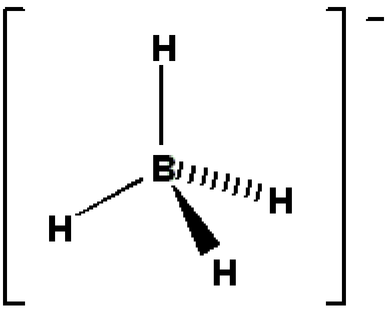
Pentagonal bipyramidal structure of IF_7

In IF_7 Iodine heptafluoride out of 7 Flourine atoms, 5 of them are placed on a plane in Pentagon shape. In the remaining 2 flourines one is placed above the plane and the other below the plane each at 90 degrees

2.

(c) BH_4^-

Explanation:



Boron is surrounded by 4 bond pairs.

In BH_4^- = no. of bond pair = 4

No. of lone pair = 0

It undergoes sp^3 hybridized, so have tetrahedral geometry.

NH_2^- = V - shape H_3O^+ = Pyramidal

CO_3^{-2} = triangular planar

3. **(a)** Ip-lp > Ip-bp > bp-bp

Explanation: According to VSEPR theory, the repulsive interaction of electron pairs decrease in the order.

Lone pair (Ip) - lone pair (lp) > lone pair (lp) - bond pair (bp) > bond pair (bp) - bond pair (bp).

4. **(a)** H-C < H-N < H-O.

Explanation: O is more electronegative than N which is more electronegative than C.

5.

(c) paramagnetic character decreases and the bond order increases

Explanation: For O_2: $\sigma 1s^2$, $\sigma^* 1s^2$, $\sigma 2s^2$, $\sigma^* 2s^2$, $\sigma 2pz^2$

$(\pi 2py^2 = \pi 2px^2)$, $(\pi^* 2py^1 = \pi^* 2px^1)\sigma^* 2pz$

For O_2: $\sigma 1s^2$, $\sigma^* 1s^2$, $\sigma 2s^2$, $\sigma^* 2s^2$, $\sigma 2pz^2$

$(\pi 2py^2 = \pi 2px^2)$, $(\pi^* 2py^1 = \pi^* 2px^0)\sigma^* 2pz$

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

For $O_2 = \frac{10-6}{2} = 2$

For $O_2^+ = \frac{10-5}{2} = 2.5$

6. **(a)** Metal and a Nonmetal

Explanation: Metals are electropositive in nature as they easily lose electrons, so they are reducing agents. On the contrary, Non-metals are electronegative because they gain electrons and thus they are oxidizing agents.

7.

(c) $N_a < N_b$

Explanation: Stability of molecule: It is determined by bond order. Higher is the bond order greater is the stability of the molecule.

Bond order is defined as the number of covalent bonds between the two combining atoms of a molecule.

Bond order = 0.5 ($N_b - N_a$)

If $N_b > N_a$, then the molecule will be stable.

N_b = number of bonding electrons or number of electrons in bonding M.O's.

N_a = number of antibonding electrons.

8.

(d) 0

Explanation: The steric number of BrF_5 is 6. So the geometry is octahedral. It will have 5 bps and one lp which is present in the axial position. Because of distortion caused by lp, there is no 90^o F—Br—F bond angle in BrF_5.

9. **(a)** XeF_2

Explanation:

- **In XeF_2:** Xe has 2 of its electrons in bonding with F. So, since Xe has 8 valence electrons, the no.of lone pairs = $\frac{no.\ of\ non-bonding\ electrons}{2} = \frac{6}{2} = 3$
- **In XeO_3:** Xe has 6 of its electrons in bonding with O (bcoz of double bond). So, Xe will have 2 non bonding electrons and therefore 1 lone pair.
- **In XeF_4**: Xe has 4 of its electrons in bonding with F. So, Xe will have 4 non bonding electrons and therefore 2 lone pairs.
- **In XeF_6**: Xe has 6 of its electrons in bonding with F. So, Xe will have 2 non bonding electrons and therefore 1 lone pair.

So XeF_2 has the maximum no. of lone pairs associated with Xe.

10.

(c) Q = charge, r = distance of separation

Explanation: Q = charge, r = distance of separation

11.

(d) 3,3,3

Explanation: Total number of electrons in N_2 molecule is 7+ 7= 14.

As per the formula Bonded pair of electrons N_b: $\sigma 1s^2\ \sigma 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2$

Total 10 electrons.

Anti bond pairs of electrons N_a: $\sigma 1s *^2\ \ \sigma 2s*^2$ Total 4 electrons.

Bond Order (B.O.) = $\frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

Similarly bond order of CO And NO^+ is 3

12.

(d) resonance hybrid

Explanation: resonance hybrid

13.

(c) NO

Explanation: The number of electrons in Nitrogen is 7 and in oxygen is 8. Hence the number of electrons that would be present in the molecular orbitals in NO is 7 + 8 = 15. As the number of electrons is odd, all the electrons in the NO molecule cannot be paired. Hence, a single electron would be present in a π*2p orbital. Therefore NO is an odd electron species and the gas is hence paramagnetic due to the presence of unpaired electron.

14.

(d) 2.5 and 1. 5

Explanation: The electronic configuration of the O_2^+ ion containing 15 electrons can be written as: $\sigma 1s^2 < \sigma^*1s^2 < \sigma 2s^2 <$ $\sigma^*2s^2 < \sigma 2p_z^2 < (\pi 2p_x^2 = \pi 2p_y^2) < (\pi^*2p_x^1 = \pi^*2p_y) < \sigma^*2p_z$

The bond order can be found as:

B.O = $\frac{N_b - N_a}{2}$

N_b= Number of electrons in the bonding orbitals = 10

N_a = Number of electron in the anti-bonding orbitals = 5

B.O = $\frac{10-5}{2} = \frac{5}{2}$= 2.5

The electronic configuration of the O_2^- ion containing 17 electrons can be written as: $\sigma 1s^2 < \sigma^*1s^2 < \sigma 2s^2 < \sigma^*2s^2 < \sigma 2p_z^2 <$ ($\pi 2p_x^2 = \pi 2p_y^2) < (\pi^*2p_x^2 = \pi^*2p_y^1) < \sigma^*2p_z$

The bond order can be calculated as: B.O = $\frac{10-7}{2} = \frac{3}{2}$ = 1.5

15. **(a)** Dipole moment

Explanation: A dipole moment is a measurement of the separation of two opposite electrical charges. Dipole moments are a vector quantity. The magnitude is equal to the charge multiplied by the distance between the charges and the direction is from negative charge to positive charge: μ = q $\times$ r

16.

(b) number of bonded valence electron pairs. and number of non-bonded valence electron pairs.

Explanation: The shape of a molecule depends upon valence electron pairs i.e. both bonded and non-bonded.

17.

(c) C_2H_4

Explanation: $\underset{H}{\overset{H}{\underset{|}{\overset{|}{C}}}} = \underset{H}{\overset{H}{\underset{|}{\overset{|}{C}}}}$

Each C atom is surrounded by 4σ bonds and 1π bond. Bond length of the double bond is smaller than a single bond.

18.

(b) O_2^{2+}

Explanation: Bond length is inversely proportional to bond order. O_2^{2+} has highest bond order as electrons are removed from antibonding MO.

19.

(d) O_2^{2-}

Explanation: **For $O_2{}^{2-}$: $\sigma 1s^2$, σ^*1s^2, $\sigma 2s^2$, σ^*2s^2 ,$\sigma 2pz^2$, (Л2py^2= Л2px^2), (Л*2py^2= Л*2px^2) σ^*2pz.**

There is no unpaired electrons.

20. **(a)** 3,3,3

Explanation: Total number of electrons in the N_2 molecule is 7+ 7= 14.

As per the formula Bonded pair of electrons N_b: $\sigma 1s^2\ \sigma 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2$

Total 10 electrons.

Anti bond pairs of electrons N_a: $\sigma 1s*^2\ \ \sigma 2s*^2$ Total 4 electrons.

Bond Order (B.O.) = $\frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

Similarly, the bond order of CO and NO^+ is 3.

21. **(a)** NO^+ NO^+

Explanation: Molecules having all their subshells completely filled are diamagnetic, i.e., they are not influenced greatly by a magnetic field. Paramagnetic is the opposite and it is the nature of elements possesing incompletely filled subshell(s).

The Lewis structure of NO molecule can be represented as

: N . = : O : or

: N : = . O :

It is observed that the total no. of unbonded electrons is odd. Therefore, there must be an incompletely filled subshell. Therefore, it is paramagnetic. In NO^+, due to loss of 1 electron, the no. of unbonded electrons becomes even. Therefore, all subshells must be completely filled. Therefore, it is diamagnetic.

22. **(a)** H_2O

Explanation: Because of presence of lp on head of O atom in H_2O therefore it has max dipole moment.

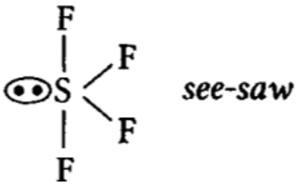
23.

(c) A

Explanation: Octate of A is completely filled.

24.

(b) see-saw

Explanation: 

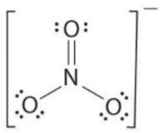
25. **(a)** BC_3

Explanation: B represents phosphorus (P) and C represents Chlorine (Cl). Thecompound formed is PCl_3 i.e., BC_3.

26.

(d) 4, 0

Explanation: The nitrate ion is formed by the loss of the hydrogen ion, and so its structure is:



Around the central nitrogen there are 4 pairs of shared electrons, and no remaining lone pair. The original lone pair has now become a bonding pair. Two of those pairs make up a double bond.

27.

(b) chemical bonds

Explanation: The atoms can achieve the stable octet when they are linked by chemical bonds. It was postulated by Lewis.

28. **(a)** sp, sp^2 and sp^3

Explanation:

- In NO_2^+

  Number of electron pairs =2
  Number of bond pairs = 2
  Number of lone pair = 0
  So, the species is linear with sp hybridisation.
- In NO_3^-
  Number of electron pairs =3
  Number of bond pairs = 3
  Number of lone pair = 0
  So, the species is trigonal planar with sp2 hybridisation
- In NH_4^+
  Number of electron pairs =4
  Number of bond pairs = 4
  Number of lone pair = 0
  So, the species is tetrahedral with sp^3 hybridisation.

29.

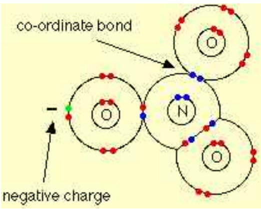
(d) Covalent

Explanation: Both B and C are non-metals. So, the bond formed between them will be covalent. B is phosphorous and C is chlorine.

30.

(d) 4,0

Explanation: The nitrate ion is formed by the loss of the hydrogen ion, and so its structure is:



Around the central nitrogen there are 4 pairs of shared electrons, and no remaining lone pair. The original lone pair has now become a bonding pair. Two of those pairs make up a double bond. The double bond unit and the two single bonds arrange themselves as far apart as possible in a trigonal planar arrangement - exactly the same as the carbonate ion.

31. **(a)** $\mathrm{H_2O} > \mathrm{HF} > \mathrm{NH_3}$

Explanation: The smaller the size of the atom, the greater is the electronegativity, and hence stronger is the H−bonding. Thus, the order of strength of H-bonding is HF > H_2O > NH_3.

But each HF molecule is linked only to two other HF molecules while each H2O molecule is linked to four other H2 O molecules through H−bonding.

Hence, the decreasing order of boiling points is H_2O > HF > NH_3.

32. **(a)** 99 pm

Explanation: Radius of Cl-atom = $\frac{Bond\ distance\ in\ CI2}{2}$

$= \frac{198}{2}$ = 99 pm

33.

(c) H_2SO_4

Explanation: In H_2SO_4, S has more than 8 electrons in the valence shell. It has 12 electrons in its valence shell.

34. **(a)** determining molecular shape

Explanation: Bond angles also contribute to the shape of a molecule. Bond angles are the angles between adjacent lines representing bonds. The bond angle can help differentiate between linear, trigonal planar, tetrahedral, trigonal-bipyramidal, and octahedral. The ideal bond angles are the angles that demonstrate the maximum angle where it would minimize repulsion, thus verifying the VSEPR theory.

35.

(c) $\mathrm{NO}^{+}$and N_2

Explanation: CO, $N_{2,}$ and NO^+ have same number of electrons = 14.

36.

(b) – 0.75

Explanation: Formal change = $\frac{Total\ charge}{no.\ of\ o-atom}$

= $\frac{-3}{4}$ = -0.75

37.

(d) 

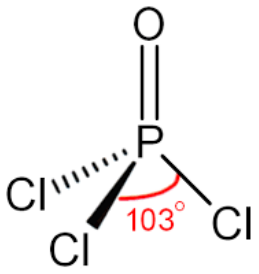
Explanation: Out of phase overlap occurs in

Thus, it shows negative overlap

38.

(d) sp^3, 0

Explanation: 4σ bonds are present without any lone pair of electrons so sp^3 hybridisation.



39. **(a)** C_2H_2

Explanation: C_2H_2

40.

(b) triple bond

Explanation: The bond length depends on the strength of the bond. The stronger the bond is, the shorter it will be. The triple bonds are the strongest and hence the shortest. Then comes double bonds which are of intermediate strength between the triple and single bonds. And finally, the single bonds are weaker than the other two. This way, Triple bonds are the shortest. Then comes double bonds. Finally, single bonds are the longest among the three.

The order of bond lengths is given as Triple bond < Double bond < Single bond.

41. **(a)** one sigma and two pi bonds

Explanation: Acetylene is C_2H_2.

C atoms are bound with triple bond i.e. 1 sigma and 2 pie bonds.

42. **(a)** Both overlapping of atomic orbitals and hybridisation of atomic orbitals

Explanation: The valence bond theory explains the shape, the formation and directional properties of bonds in polyatomic molecules like CH_4, NH_3 and H_2O etc., in terms of overlap and hybridization of atomic orbitals.

43.

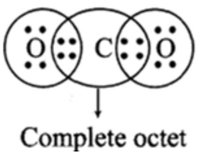
(b) inertness of noble gases

Explanation: inertness of noble gases

44.

(c) CO_2

Explanation: Lewis structure of CO_2 can be represented as



45.

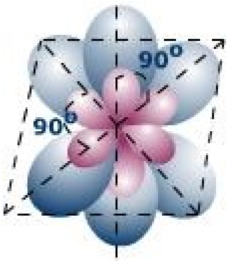
(b) atomic orbitals must have comparable energies and of proper symmetry

Explanation: According to molecular orbital theory (MOT), the atomic orbitals of comparable energies and proper symmetry combine to form molecular orbitals.

46.

(b) Octahedral geometry

Explanation: In sp^3d^2 Hybridization one s, three p and two d-orbitals get hybridized to form six sp^3d^2 hybrid orbitals which adopt octahedral arrangement as given in image below



47.

(d) Sharing of electrons contributed by one atom only

Explanation: Covalent bond is formed by two atoms sharing a pair of electrons. The atoms are held together because the electron pair is attracted by both of the nuclei.

In the formation of a simple covalent bond, each atom supplies one electron to the bond. But a **co-ordinate bond** (also called a dative covalent bond) is a covalent bond (a shared pair of electrons) in which both electrons come from the same atom.

48.

(d) All of these

Explanation: VSEPR provides a simple procedure to predict the shapes of covalent molecules.

49.

(b) All of these

Explanation: The molecular orbitals like atomic orbitals are filled in accordance with the Aufbau principle, Pauli's exclusion principle and Hund's rule.

50.

(d) 120°

Explanation: sp^2 hybridisation have triangular planar structure so bond angle is 120^o

51.

(b) molecular orbital

Explanation: The electron probability distribution around a group of nuclei in a molecule is given by a molecular orbital.

52.

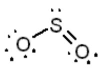
(b) Electronic theory of chemical bonding

Explanation: Ionic crystalline compounds formed by ion-formation by electron transfer proves the electronic theory of chemical bonding. According to this theory, ionic bonds are formed when an atom loses electron/ electrons to become a positive ion and another atom gains the electron/ electrons to become a negative ion. Ionic bonding is a process of complete transfer of Valence electrons.

53.

(c) sP^2 hybridized

Explanation: The Lewis Structure of SO_2.



there are 3 things or regions around the central atom. 1 single bond, 1 double bond and a lone pair.

The hybridization for S is sp^2 Where: s has 1 orbital, p has 3 orbitals, d has 5 orbitals, and f has 7 orbitals.

54. **(a)** Isoelectronic

Explanation: Isoelectronic species are elements or ions that have the same, or equal number of electrons. Although isoelectronic species have the same number of electrons, they are different in their physical and chemical properties. All of given species have 14 electrons.

55.

(d) hybrid orbitals

Explanation: According to Pauling the atomic orbitals combine to form new set of equivalent orbitals known as **hybrid orbitals**. Unlike pure orbitals, the hybrid orbitals are used in bond formation.

56.

(b) Repulsive effect

Explanation: Greater repulsion between lone pairs of electrons as compared to the lone pair-bond pair and bond pair-bond pair repulsions. These repulsive effects result in deviations from idealized shapes and alterations in bond angles in the molecules.

57.

(c) 104.5^0

Explanation: Due to the presence of two lone pairs on O in the H_2O bond angle reduce to 104.5^0 from 109^0.

58.

(c) energy decreases

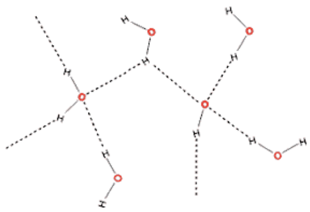
Explanation: A chemical bond is a lasting attraction between atoms that enables the formation of chemical compounds. The bond may result from the electrostatic force of attraction between atoms with opposite charges, or through the sharing of electrons as in the covalent bonds. When a bond forms, electrons are attracted to the space between nuclei where the electrostatic force of attraction is greater. As the electrons fall to a position of lower potential energy, the total mechanical energy of the molecular system decreases. Part of the mechanical energy of the unbound atoms is lost when they form the molecular system with a lower total mechanical energy.

59. **(a)** Nyholm and Gillespie

Explanation: Nyholym and Gillespie further developed and refined the VSEPR theory.

60.

(d) 4

Explanation: 

The two hydrogens of the water molecule can form hydrogen bonds with other oxygens in water, and the two lone pair of electrons on oxygen of the water molecule can attract other hydrogens in water. Hence, 4 possible hydrogen bonds.

61. **(a)** Outermost orbitals of the noble gases are fully filled**.**

Explanation: The six noble gases (Helium, Neon, Argon, Krypton, Xenon, Radon) are found in group 18 of the periodic table. These elements were considered to be inert gases until the 1960s because their oxidation number of 0 prevents the noble gases from forming compounds readily. All noble gases have the maximum number of electrons possible in their outer shell (2 for Helium, 8 for all others), making them stable.

62.

(c) All of these

Explanation: VET is based on all the given criteria.

63.

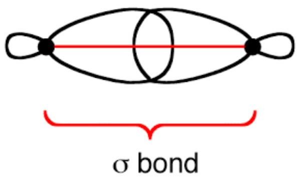
(d) sp^2 , sp , sp^3

Explanation: Hybridization of orbitals of N atom in NO_3^- , NO_2^+ and $NH_4{}^+$ are sp^2, sp, sp^3 which can be explained by their Lewis structures. The empty p- orbitals of N take part in hybridization.

64.

(c) sigma bond

Explanation: Sigma bond: This type of covalent bond is formed by the axial or end to end overlapping of half-filled atomic orbitals of the atoms participating in bonding. The electron cloud formed as a result of axial overlap is cylindrically symmetrical about an internuclear axis.



65.

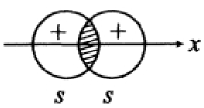
(d) lattice enthalpy

Explanation: Lattice enthalpy is simply the change in Enthalpy associated with the formation of one mole of an ionic compound from its oppositely charged ions in their standard states under standard conditions.

66. **(a)** have no real existence

Explanation: The canonical forms have no real existence.

67.

(d) 

Explanation: H-atom has only s orbital. Therefore, s-s overlapping takes place in the formation of H_2 molecule.

68.

(b) p_z orbitals

Explanation: In such a condition only p_z orbitals form σ-bond along the z-axis while p_x and p_y form π-bonds.

69. **(a)** C-O > C=O > C $\equiv$ O

Explanation: The length of the bond is determined by the number of bonded electrons (the bond order). The higher the bond order, the stronger the pull between the two atoms and the shorter the bond length. Generally, the length of the bond between two atoms is approximately the sum of the covalent radii of the two atoms. CO has a triple bond so has minimum bond length.

70. **(a)** the number of valence electrons of the element

Explanation: Gilbert N. Lewis is widely known for his use of simple symbolic representations of elements that show valence electrons as dots.

The Lewis electron-dot symbols focus on the electrons in the highest principal energy level in the atom, the valence electrons. After all, these are the electrons that participate in chemical reactions. Lewis electron-dot symbols work well for the representative elements.

71.

(c) valence shell electrons with opposite spins.

Explanation: Overlapping of atomic orbitals having electrons of opposite spin take place in the formation of a molecule to cancel the dipole moment.

72. **(a)** lattice enthalpy

Explanation: Lattice energy is an estimate of the bond strength in ionic compounds. It is defined as the heat of formation for ions of opposite charge in the gas phase to combine into an ionic solid. The stability of the ionic bond is directly proportional to lattice energy.

73.

(c) positive

Explanation: Bond order $\propto$ Stability

Hence, for a stable molecule, the value of bond order must be positive. When bond order is zero and negative the molecule will not exist.

74. **(a)** between the nuclei of the bonded atoms

Explanation: In case of bonding molecular orbital the electron density is located between the nuclei of the bonded atoms.

75.

(b) C≡O > C=O > C-O

Explanation: The more bonds between atoms, the shorter the bond is. For example, a triple bond will be shorter than a double bond. The shorter the bond is the stronger it is (higher dissociation energy). Bond order is basically the number of bonds between atoms, so a larger number bond order=more bonds=shorter bonds=stronger bonds. i.e., C≡O > C=O > C-O.