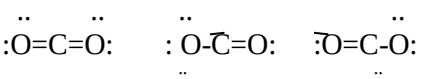
Chapter 6: Chemical Bonding

Review Questions

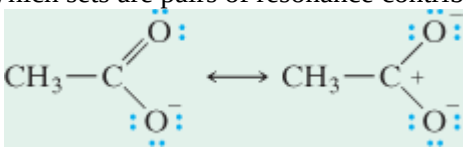
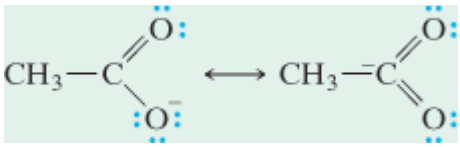
1. In the CO_2 molecule, where are the non-bonding electrons? How many electrons are in bonding domains?

Solution

In CO_2, the non-bonding electrons are located in the non-bonding electron domains of the oxygen atoms. Also, each oxygen in CO_2 has two non-bonding domains. There are two electrons in each non-bonding domain, but four electrons are located in each of the bonding domains between carbon and oxygen. Four electrons are required to form a double bond.



2. Which sets are pairs of resonance contributing structures?

(a)  (b) 

Solution

Pair (a) is a set of resonance structures.

3. Which of the following molecules/ions have resonance structures?

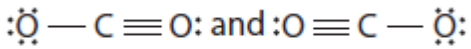
(a) HCN (b) SO_2 (c) PCl_3 (d) OCN^-

Solution

The molecule (b) SO_2 and the ion (d) OCN^-

4. A possible Lewis structure for CO_2 involves a triple bond and a single bond between C and the 2 O. Write the two resonance structures involving the triple bond. What is implied about the nature of the C–O bond by these two structures? How does this relate to the common Lewis structure for CO_2 involving two double bonds?

Solution



These resonance structures imply either that the bonds are unequal (which they are not) or that they average out to two double bonds (which they do). Resonance structures are not necessary in this case since they lead to the same conclusion as the one structure with two double bonds.

5. What is the theory proposed to explain the bonding in metals?

Solution

Molecular orbital theory may be extended to explain bonding in metals and is known as the band theory.

6. How does the theory for metallic bonding explain the following properties of metals: Luster, malleability, and conductivity?

Solution

The electrons have high degree of mobility and this explains the high electrical and thermal conductivity of metals. In presence of an electric field, the mobile electrons, which move in all directions in absence of a field, begin to move toward the anode, causing electric current to flow.

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Similarly when the metal is heated, the mobile electrons gain energy and move toward the colder part of the metal, thus transferring heat. The metallic luster arises because of absorption of light of any wavelength by metal atoms on the surface and its re-emission. The electrons absorb light, move to higher energy levels, and then emit the same amount of light to come back to the ground state.

7. What are the molecular orbitals obtained by the linear combination of the atomic orbitals 1*s*, 2*s* and 2*p*? Give their increasing order of energy.

Solution

The orbitals are $\sigma_g 1s,\ \sigma_u^* 1s,\ \sigma_g 2s,\ \sigma_u^* 2s,\ \pi_u 2p_x,\ \pi_u 2p_y,\ \sigma_g 2p_z,\ \pi_g^* 2p_x,\ \pi_g^* 2p_y$ and $\sigma_u^* 2p_z$. The energetic ordering of these orbitals from Li to N is

$$\sigma_g 1s < \sigma_u^* 1s < \sigma_g 2s < \sigma_u^* 2s < \pi_u 2p_x = \pi_u 2p_y < \sigma_g 2p_z < \pi_g^* 2p_x = \pi_g^* 2p_y < \sigma_u^* 2p_z$$

whereas for O, F, and Ne, the order is

$$\sigma_g 1s < \sigma_u^* 1s < \sigma_g 2s < \sigma_u^* 2s < \sigma_g 2p_z < \pi_u 2p_x = \pi_u 2p_y < \pi_g^* 2p_x = \pi_g^* 2p_y < \sigma_u^* 2p_z$$

The reason for this difference can be understood roughly as follows: The $\sigma_g 2p_z$ orbital has considerable electron density along the *z*-axis. For $Z \geq 5$, the contributing $2p_z$ orbitals are sufficiently delocalized along this axis so that electrons occupying them overlap and, therefore, contribute significant Coulomb repulsion to the total energy, which raises the energy of the $\sigma_g 2p_z$ orbital above the σ orbitals, for which the Coulomb repulsion is considerably less, as they do not peak along the *z*-axis. For $Z \geq 6$, however, the nuclear charge is sufficiently high so that the $2p_z$ orbitals decay more quickly, are more localized, and the extent of overlap reduces and Coulomb repulsion is reduced to such an extent that the energy of the $\sigma_g 2pz$ orbital is now below that of the π orbitals.

8. Give the molecular orbital configuration and the bond order in the molecules N_2, F_2, CN^- and O_2^-.

$$F_2 = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi^2 p_y^2 \pi^* 2p_x^2 = \pi^* 2p_y^2$$

Bond order = (10 – 8)/2 = 1

$$N_2 = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2 \sigma 2p_z^2$$

Bond order = (10 – 4)/2 = 3

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

Bond order = (10 – 4)/2 = 3

$$O_2^- = \sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \sigma 2p_z^2\ \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^2 = \pi^* 2p_y^1$$

$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{10-7}{2} = \frac{3}{2}$$

9. How does molecular orbital theory account for magnetic properties of molecules?

Solution

The molecules that contain unpaired number of electrons are paramagnetic in nature, while those will filled electrons are diamagnetic.

10. Interpret the non-linear shape of H_2S and non-planar shape of PCl_3 using valence shell electron pair repulsion (VSEPR) theory.

Solution

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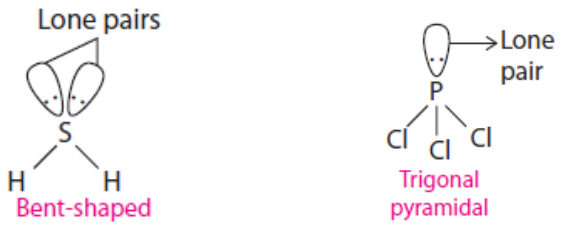

In H_2S, S is in sp^3-hybridized state having two lone pairs on it as shown above. The angle is contracted due to lone pair effect to produce V-shaped structure. On the other hand, in PCl_3, P is also sp^3 hybridized having one lone pair on it. This time the tetrahedron structure is distorted to produce pyramidal shape due to lone pair effect.

11. Briefly describe valence bond theory of covalent bond. How can you interpret it in terms of energy considerations?

Solution

It is a theory of chemical bonding that explains the formation of covalent bonds on the basis of overlapping of atomic orbitals and pairing of electrons with opposite spins.

It was proposed by Heitler and London in 1927. The theory was modified by Pauling and Slater in 1931.

The important features of valence bond theory are as follows:

(a) A covalent bond is formed by overlapping of orbitals having two unpaired electrons of opposite spins.

(b) When the overlapping between orbitals is greater, the covalent bond formed is stronger.

(c) When a covalent bond is formed, the energy of the molecule decreases and the molecule becomes more stable.

(d) The extent of overlapping of orbitals follows the following order:
s–s overlapping > *s–p* overlapping > *p–p* overlapping.

(e) Overlapping along the axis is stronger than overlapping sideways.

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12. What are σ and π bonds? How do they differ from each other?

Solution

Sigma (σ) bond is a type of covalent bond formed by the axial overlapping of half-filled atomic orbitals.

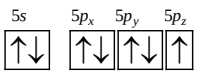
Pi (π) bond is a type of covalent bond formed by the lateral or sidewise overlap of the atomic orbitals.

Sigma bond	**Pi bond**
(a) A σ bond is stronger because the extent of overlapping is greater.	(a) A π bond is weaker because the extent of overlapping is less.
(b) A σ bond involves overlap of *s–s*, *s–p* or *p–p* orbitals.	(b) It involves the overlap of *p*-orbitals only.
(c) The bond may be present between two atoms either alone or along with π bond.	(c) The bond is always present between the two atoms along with a σ bond.

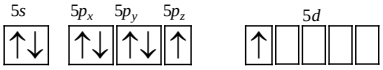
13. Explain the shape of I_3^- ion.

Solution

Iodine (ground state): $5s^2\,5p^5$



Iodine (excited state): $5s^2\,5p^5\,5d^1$



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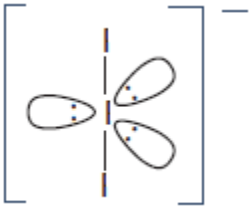
Hybridization: $\overset{5s}{\boxed{\uparrow\downarrow}}\ \overset{5p_x}{\boxed{\uparrow\downarrow}}\overset{5p_y}{\boxed{\uparrow\downarrow}}\overset{5p_z}{\boxed{\uparrow\times}}\quad \overset{5d}{\boxed{\uparrow\times}\boxed{}\boxed{}\boxed{}\boxed{}}$ sp^3d hybridization.

Consider the I_3^- anion. In this case, iodine orbitals are involved in sp^3d hybridization. One of the hybrid orbitals is utilized for accepting an electron pair from 5*p* orbital of another I^- as shown above. The singly filled sp^3d hybrid orbital of iodine overlaps with the half-filled 2*p* orbital of one iodine atom to form I–I σ bond. Due to the symmetrical arrangement of three lone pairs around the central iodine atom, the net bond pair–lone pair repulsion is the same along the axial directions. The anion, therefore, has a linear structure.



14. Phosphorus pentachloride dissociates in the vapor phase according to the equation: $PCl_5 \longrightarrow PCl_4^+ + Cl^-$.

The hybridization of phosphorus in PCl_5 is sp^3d. What is the hybridization of P in PCl_4^+?

Solution

PCl_4^+:

$$H = S + \frac{1}{2}[E - V \pm C] = 4 + \frac{1}{2}[5 - 4 - 1] = 4 = sp^3$$

P is in sp^3-hybridized state.

15. An energy level scheme for the orbital of second period diatomic molecules O_2 to Ne_2 lists the molecular orbitals in the following order of increasing energy.

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

Based on this energy level scheme, how many unpaired electrons are in all of the antibonding molecular orbitals of the O_2 molecule in its ground state?

Solution

There are two unpaired electrons in all of the antibonding molecular orbitals of the O_2 molecule in its ground state.

It has 16 electrons.

MOC of O_2: $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_x^2 \pi 2p_z^2 \pi 2p_y^2 \pi^* 2p_z^2 \pi^* 2p_y^2$.

That is, two unpaired electrons.

16. What factors control the bond length in a covalent bond?

Solution

The bond distance in a covalent bond is determined by a balance (compromise) between the separate attractions of the nuclei for the electron density that is found between them, and the repulsions between the like-charged nuclei and those between the like-charged electrons. These attractions and repulsions oppose one another, and a bond distance is achieved that maximizes the attraction, while minimizing the repulsions.

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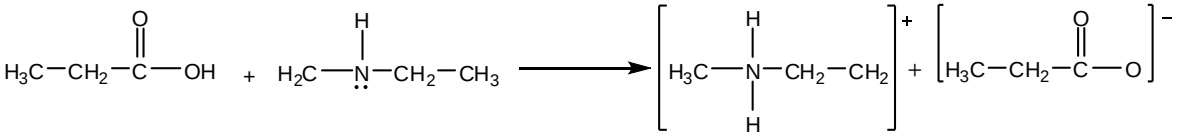
17. Why do elements of the second period never form more than four covalent bonds? Why are elements of the third period able to exceed an octet?

Solution

The valence shell of a period 2 element can hold only eight electrons since only two types of subshells, *s* and *p*, are available in this period. The valence shell of elements in row three can hold as many as eighteen electrons. This results because there are three types of subshells, *s*, *p*, and *d*, that can hold electrons. The *s* subshells hold 2 electrons, *p* subshells hold 6 electrons and *d* subshells hold 10 electrons. The total is, therefore, 18 electrons.

18. Use Lewis structures to diagram the reaction between the propanoic acid and ethylmethylamine.

Solution



19. Which elements are assigned electronegativities of zero? Why?

Solution

The noble gases are assigned electronegativity values of zero because they have a complete shell of electrons.

20. What is the minimum number of electrons that would be expected to be in the valence shell of As in $AsCl_5$?

Solution

We expect a minimum of ten electrons, in bonding pairs between As and each of five Cl atoms.

21. Define bond order. How are bond energy and bond length related to bond order? What can be inferred from these relationships?

Solution

The number of electron pairs that comprise a covalent linkage (bond) between two atoms is called the bond order. As bond order increases, so does the strength of the bond. Therefore, as bond order increases, bond energy increases and bond length decreases, in keeping with increased bond strength.

22. The energy required to break the H–Cl bond to give H+ and Cl^- ions would not be called the H–Cl bond energy. Why?

Solution

The H–Cl bond energy is defined to be the energy required to break the bond to give atoms of H and Cl, not the ions H^+ and Cl^-. In other words, when the bond is broken between the H and the Cl atoms, one of the two electrons of the bond must go to each of the atoms. When ions are obtained, however, both electrons of the bond are given to the chlorine atom and the lattice energy would be the more appropriate term to use.

23. What is a resonance hybrid? How does it differ from the resonance structures drawn for a molecule?

Solution

A resonance hybrid is the true structure of a molecule or polyatomic ion, whereas the various resonance structures that are used to depict the hybrid do not individually have any reality. The hybrid is a mix, or average, of the various resonance structures that compose it.

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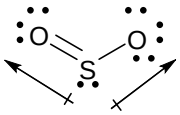
24. Below are two structures for $BeCl_2$. Give two reasons why the one on the left is the preferred structure.

Solution
The formal charges on all of the atoms of the left structure are zero; therefore, the potential energy of this molecule is lower and it is more stable.

25. What arrangements of domains around an atom are expected when there are (a) three domains, (b) six domains, (c) four domains, or (d) five domains?
Solution
(a) Planar triangular, otherwise known as trigonal planar
(b) Octahedral
(c) Tetrahedral
(d) Trigonal bipyramidal

26. Use a drawing to show why the SO_2 molecule is polar.
Solution



The individual bond dipoles do not cancel one another.

27. (a) Are all dissymmetric molecules polar?
(b) What condition must be met if a molecule having polar bonds is to be non-polar?
Solution
(a) Dissymmetric molecules are polar only if there is a difference is the electronegativity of the atoms.
(b) A molecule having polar bonds will be nonpolar only if the bond dipoles are arranged so as to cancel one another's effect.

28. Why do elements of the second period never use sp^3d or sp^3d^2 hybrid orbitals for bond formation?
Solution
Elements in the second period do not have a *d* subshell in the valence level.

29. What relationship is there, if any, between Lewis structures and the valence bond descriptions of molecules?
Solution
Lewis structures are something like shorthand representations of VB descriptions of molecules.
30. How can the VSEPR model be used to predict the hybridization of an atom in a molecule?
Solution
The VSEPR model gives the shape of the electron domains around the central atom; this information is used to predict the hybridization of the atom.
31. If the central oxygen in the water molecule did not use sp^3-hybridized orbitals (or orbitals of any other kind of hybridization), what would be the expected bond angle in H_2O (assuming no angle-spreading force)?
Solution
This angle would have had to be 90°, the angle between one atomic *p*-orbital and another.

32. How does the geometry around B and O change in the following reaction? How does the hybridization of each atom change?

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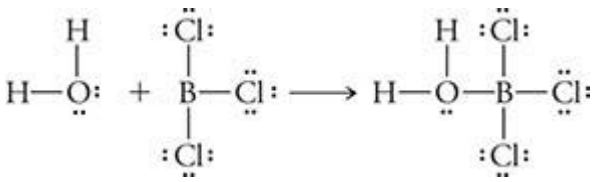


Solution

The geometry of the boron BCl_3 is trigonal planar; after the addition of the water molecule, it becomes tetrahedral. The hybridization changes from sp^2 to sp^3. The geometry of the oxygen in the water molecule changes from bent to trigonal pyramidal after coordination to the boron atom. Its hybridization does not change, but remains at sp^3.

33. On the basis of MO theory, explain why Li_2 molecules can exist but Be_2 molecules cannot. Could the ion Be_2^+ exist?

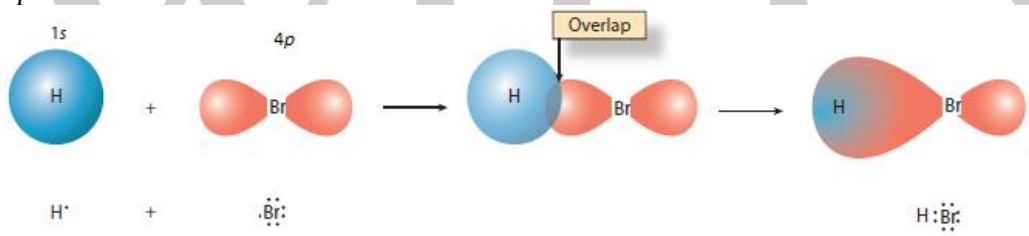
Solution

The bond order of Li_2 is 1.0. The bond order of Be_2 would be zero. Yes, Be_2^+ could exist since the bond order would be 1/2.

34. Use sketches of orbitals to describe how valence bond theory would explain the formation of H–Br bond in hydrogen bromide.

Solution

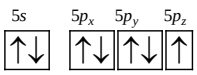
Electronic configuration of hydrogen is $1s^1$. Electronic configuration of bromine is $[Ar]3d^{10}4s^24p^5$. When hydrogen bromide molecule is formed, the 1*s* atomic orbital of hydrogen atom overlaps with $4p_z$ atomic orbital of bromine atom along the axis. This leads to overlapping of *s*–*p* orbitals, forming *s*–*p* σ bond.

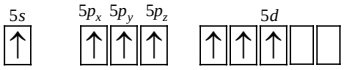


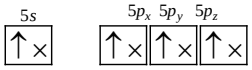
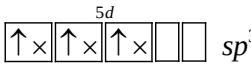
35. On the basis of hybridization, explain the shape of iodine heptafluoride.

Solution

Iodine heptafluoride (IF_7): The electronic configuration of iodine in ground and excited states are as follows:

Iodine (ground state): $5s^2\,5p^5$ 

Iodine (excited state): $5s^1\,5p^3\,5d^3$ 

Hybridization:   sp^3d^3 hybridization with 7 electrons coming from 7 fluorine atoms.

The outer electronic configuration of the central atom of iodine in the grand state is $5s^2p_x^2p_y^2p_z^1$, as represented in the above orbitals. On excitation unpairing of 5*s* as well as of $5p_x$ and $5p_y$ electrons take place as shown above. Now the seven orbitals hybridize to give seven sp^3d^3 hybrid orbitals. Each of these hybrid orbitals overlaps with the half-filled 2*p* orbitals of each fluorine atom to form IF_7 molecule as shown below.

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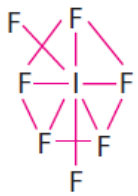

Five hybridized orbitals are coplanar and two perpendicular to this plane. The resulting molecule acquires pentagonal bipyramidal geometry. In this molecule, again all the seven bonds are not equivalent. Five of these are directed towards the vertices of a regular pentagon, making an angle of 72° with one another. The remaining two are oriented at right angles to the plane of the first set of five orbitals. Thus, the bond angles in this molecule are of 72° and 90°.

36. What hybrid orbitals would be expected for the central atom in the following: (a) $AlBr_3$, (b) $BeCl_2$, (c) $SiCl_4$, (d) BeF_3^-, (e) $AsCl_3$?

Solution

The formula

$$H = SA + \frac{1}{2}(E - V) \qquad \text{or} \qquad H = SA + \frac{1}{2}[E - V \pm C]$$

where H is the hybridization of central metal atom, SA = is the number of surrounding atoms or groups, E is number of outer shell electrons in the central atom in the group state, V is valency of the central atom in the given species, and C signifies that the electrical charge of species is ionic.

Using (+) sign with E for anions and (–) sign with E for cations:

(a) $AlBr_3$ $\quad H = 3 + \frac{1}{2}[3 - 3] = 3 = sp^2$

(b) $BeCl_2$ $\quad H = 2 + \frac{1}{2}[2 - 2] = 2 = sp$

(c) $SiCl_4$ $\quad H = 4 + \frac{1}{2}[4 - 4] = 4 = sp^3$

(d) BeF_3^- $\quad H = 3 + \frac{1}{2}[2 - 3 + 1] = 3 = sp^2$

(e) $AsCl_3$ $\quad H = 3 + \frac{1}{2}[5 - 3] = 3 + \frac{1}{2}[2] = 4 = sp^3$

37. When two ice cubes are pressed together, they join to form one cube. Which force holds them together?

Solution

Ice cube contains water molecules and they are present at a particular lattice. When the two cubes are pressed together, they join to form one cube due to the formation of hydrogen bonding between the water molecules of two cubes, which results in forming one cube.

38. Define electronegativity. How does it differ from electron gain enthalpy?

Solution

Electronegativity is the tendency of an atom to attract shared pair of electron toward itself in a covalent bond. Fluorine is the most electronegative element in the periodic table; if it forms a bond with hydrogen atom, then the bonded electrons will be attracted toward fluorine. Whereas electron gain enthalpy is defined as the amount of energy released when an electron gets added to a gaseous atom. It has to be high, so that when an electron is added, a large amount of energy is released (exothermic process). For example, in case of halogens the value of electron gain enthalpy is highly negative means that an electron can be easily added so as to attain stable noble gas configuration.

39. Write a Lewis structure and any equivalent resonance structures for ozone (O_3), where one oxygen serves as the central atom.

Solution

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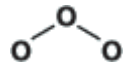
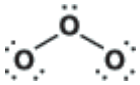
(a) No ions are involved.
(b) There are 18 (3 × 6) outer electrons.
(c) The skeletal structure is

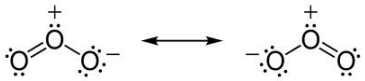
(d) Use 4 electrons for the two bonds.



(e) Add the remaining 14 electrons.



Notice that the central oxygen does not have an octet. A double bond must be made from one of the outer oxygen atoms. Notice that it does not matter which oxygen is chosen; the two oxygen atoms are equivalent. The two resonance structures indicate that each oxygen–oxygen bond is a hybrid between a single and a double bond.



Numerical Problems

1. The chlorine end of the chlorine monoxide molecule carries a charge of +0.167*e*. The bond length is 154.6 pm. Calculate the dipole moment of the molecule in Debye units. (Hint: Be sure to convert the charge to coulombs.)

Solution

For $\mu = q \times r$, we use the given values

$$q = 0.167\, e^{-}\left(\frac{1.602\times10^{-19}\ \text{C}}{1\ \text{e}^{-}}\right) = 2.675 \times 10^{-20}\ \text{C}$$

$r = 154.6\ \text{pm} = 154.6 \times 10^{-12}\ \text{m}$

$\mu = (2.675 \times 10^{-20}\ \text{C}) \times (154.6 \times 10^{-12}\ \text{m}) = 4.136 \times 10^{-30}\ \text{C m}$

In Debye units, we have:

$$q = 4.136 \times 10^{-30}\ \text{C m}\left(\frac{1\ \text{D}}{3.34\times10^{-30}\ \text{C m}}\right) = 1.24\ \text{D}$$

2. Although isolated Na^{+} and Cl^{-} ions are unstable, these ions can exist in the gaseous state as *ion pairs*. An ion pair consists of an NaCl unit in which the bond length is 236 pm. The dipole moment of the ion pair is 9.00 D. What are the actual amounts of charge on the sodium and chlorine atoms in this NaCl pair? What percentage of full 1+ and 1– charges are these? (This is the *percentage ionic character* in the NaCl pair.)

Solution

Given that

$$\mu = 9.00\ \text{D}\left(\frac{3.34\times10^{-30}\ \text{C m}}{1\ \text{D}}\right) = 3.006 \times 10^{-29}\ \text{C m}$$

$r = 236\ \text{pm} = 236 \times 10^{-12}\ \text{m}$

Therefore, $q = \dfrac{\mu}{r} = \left(\dfrac{3.006\times10^{-29}\ \text{C m}}{326\times10^{-12}\ \text{m}}\right) = 1.27 \times 10^{-19}\ \text{C}$

In electron charges

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$$q = 1.27 \times 10^{-19}\ \text{C}\left(\frac{1\ \text{e}^-}{1.602\times10^{-19}\ \text{C}}\right) = 0.795\ e^-$$

On the sodium the charge is +0.795 e^- and on the chlorine, the charge is –0.795 e^-.
This would be 79.5% positive charge on the Na and 79.5% negative charge on the Cl.

3. Calculate the percent of *s*-character between two hybrid orbitals having a bond angle of 105°. Given that cos 105° = 0.2588.

Solution

$$\cos\alpha = -\frac{1}{m} \Rightarrow \cos 105° = -\frac{1}{m} \Rightarrow 0.2588 = -\frac{1}{m} \Rightarrow m = \frac{-1}{0.2588} = -3.8639$$

$$\%\,s\text{-character} = \frac{1}{1-m}\times100 = \frac{1}{1-(-3.8639)}\times100 = \frac{1}{4.8639}\times100 = 20.56\%$$

4. What are the formal charges on the atoms in the HCl molecule? What are the actual charges on the atoms in this molecule? Are formal charges the same as actual charges?

Solution

The formal charges on the atoms in HCl are 0. The actual charges are H: +0.17e^- and Cl: –0.17e^-. The formal charges arise from the bookkeeping in Lewis structures, and are not the same as actual charges.

5. Use data from the tables of ionization enthalpies and electron gain enthalpies to calculate the energy changes for the following reactions.

$$\text{Na(g)} + \text{Cl(g)} \rightarrow \text{Na}^+\text{(g)} + \text{Cl}^-\text{(g)}$$

$$\text{Na(g)} + 2\text{Cl(g)} \rightarrow \text{Na}^{2+}\text{(g)} + 2\text{Cl}^-\text{(g)}$$

Approximately, how many times larger would the lattice enthalpy of $NaCl_2$ have to be compared to the lattice enthalpy of NaCl for $NaCl_2$ to be more stable than NaCl?

Solution

$Na(g) \rightarrow Na^+(g) + e^-$	IE = 496 kJ mol^{-1}
$Cl(g) + e^- \rightarrow Cl^-(g)$	EA = –348 kJ mol^{-1}
$Na(g) + Cl(g) \rightarrow Na^+(g) + Cl^-(g)$	ΔH = 148 kJ mol^{-1}
$Na(g) \rightarrow Na^+(g) + e^-$	IE = 496 kJ mol^{-1}
$Na^+(g) \rightarrow Na^{2+}(g) + e^-$	IE = 4563 kJ mol^{-1}
$2Cl(g) + 2e^- \rightarrow 2Cl^-(g)$	EA = 2(–348 kJ mol^{-1})
$Na(g) + 2Cl(g) \rightarrow Na^{2+}(g) + 2Cl^-(g)$	$\Delta H = 4.36 \times 10^3$ kJ mol^{-1}

In order for $NaCl_2$ to be more stable than NaCl, the lattice energy should be 4.36×10^3 kJ/148 kJ = 29.5 ≅ 30 times larger.

6. Changing 1 mol of Mg(s) and (1/2) mol of O_2(g) to gaseous atoms requires a total of approximately 150 kJ of energy. The first and second ionization enthalpies of magnesium are 737 and 1450 kJ mol^{-1}, respectively; the first and second electron gain enthalpies of oxygen are –141 and +844 kJ mol^{-1}, respectively, and the standard heat of formation of MgO(s) is –602 kJ mol^{-1}. Construct an enthalpy diagram and use it to calculate the lattice energy of MgO. How does the lattice energy of MgO compare with that of NaCl? What might account for the difference?

Solution

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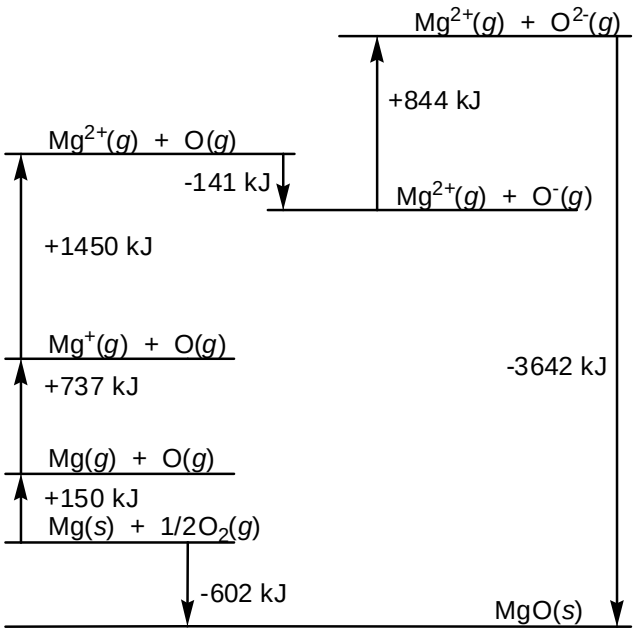

The lattice energy for NaCl is 787 kJ. The large difference in lattice energy is a result of charge (divalent versus univalent) and size of ions.

7. A molecule HX has 87.1% covalent character. Identify X if the electronegativities of halogens are as follows: F = 4.0, Cl = 3.0, Br = 2.8, I = 2.5.

Solution

According to Hanney and Smith equation, the percentage ionic character is given by

$$16\ (\chi_A - \chi_B)\ +\ 3.5\ (\chi_A - \chi_B)^2$$

Where χ_A and χ_B are the electronegantivies of the tow bonded atoms.

Given that the covalent character is 87.1%, it implies that ionic character is 12.9 %.

Using the electronegativity of hydrogen as 2.12 and electronegativities of the different halogens, we find that X is Br.

$16\ (\chi_A - \chi_B)\ +\ 3.5\ (\chi_A - \chi_B)^2 = 12.9$ for $\chi_A = 2.8$ and $\chi_B = 2.1$

8. The value of electronegativity of atoms A and B are 1.20 and 4.0, respectively. What is the percentage of ionic character of A–B bond?

Solution

Percent ionic character = $16\ (X_A - X_B) + 3.5\ (X_A - X_B)^2 = 16 \times 2.8 + 3.5\ (2.8)^2 = 72.24\%$

9. Write formal charge of the atoms in the nitrite ion.

Solution

The nitrite ion can be represented as

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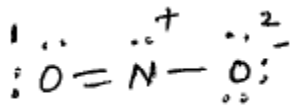


Formal charge on any atom = Number of valence electrons – Number of lone pairs – 1/2 (number of bond pairs)

Formal charge on nitrogen $= 5 - 2 - \frac{1}{2}(6) = 0$

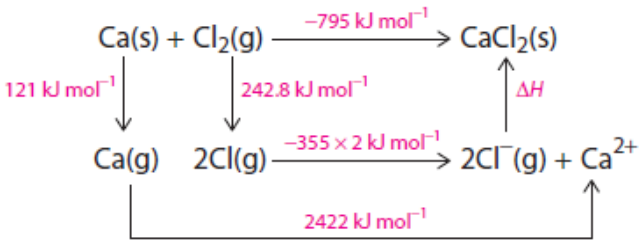
Formal charge on oxygen atom 1 $= 6 - 4 - \frac{1}{2}(4) = 0$

Formal charge on oxygen atom 2 $= 6 - 6 - \frac{1}{2}(2) = -1$

10. Calculate the lattice enthalpy of $CaCl_2$, given that the enthalpy of
(a) sublimation of Ca is 121 kJ mol^{-1},
(b) dissociation of Cl_2 to Cl is 242.8 kJ mol^{-1},
(c) ionization of Ca to Ca^{2+} is 2422 kJ mol^{-1},
(d) electron gain enthalpy for Cl to Cl^- is –355 kJ mol^{-1} and
(e) formation of $CaCl_2$ is –795 kJ mol^{-1}.

Solution

The given values can be represented in the chemical reaction as:



Applying Hess's law of constant heat summation, we get

$$-795 = 121 + 242.8 + 2422 + (-355 \times 2) + \Delta H$$

Therefore, $\Delta H = -2870.8$ kJ mol^{-1}.

11. Calculate the ionic character of HCl. Its measured dipole moment is 1.03 D. The HCl bond length is 1.275 Å.

Solution

Given, measured dipole moment = 1.03 D. Then dipole moment for 100% ionic character would be

$$\mu = q \times d = 4.8 \times 10^{-10} \times 1.275 \times 10^{-8} = 6.12 \text{ D}$$

Therefore, % ionic character $= \frac{\mu_{\text{observed}}}{\mu_{\text{ionic}}} \times 100 = \frac{1.03}{6.12} \times 100 = 16.83\%$

Additional Objective Questions

Single Correct Choice Type

1. Structure of IF_4^+ and hybridization of iodine in this structure is, respectively,

(A) linear, sp^3d (B) K-shaped, sp^3d^2
(C) irregular tetrahedral, sp^3d (D) Square planar, sp^3d^2

Solution

(C) IF_4^+ hybridization; sp^3d – irregular tetrahedral.

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2. Which of the following is/are not correct regarding C_2 molecule?
(A) C_2 molecule has been found to exist in vapor phase.
(B) It has total 12 electrons, out of which 8 electrons occupy bonding orbitals while 4 electrons occupy antibonding orbitals.
(C) The molecule is paramagnetic.
(D) C_2 molecule contains double bond and both are π bonds.

Solution

(C) $(\sigma 1s)^2(\sigma^* 1s)^2(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p_x^2 = \pi 2p_y^2)$

Four electrons are present in 2π molecular orbitals; that is why double bond contains both π bonds. Molecule is diamagnetic.

3. What fraction of charge is present on iodine in covalently bonded $H^{\delta+}-I^{\delta-}$ if the dipole moment of HI is 0.38 D and the bond length is 1.61 Å.
(A) 5% (B) 16% (C) 33% (D) 79%

Solution

(A)

Given that the dipole is 0.38 D = $0.38\times3.34\times10^{-30}$ C m = 1.26×10^{-30} C m.
If the bond was 100% ionic, the magnitude of the dipole moment will be

$|\mu| = qr = 1.6\times10^{-19}\times1.61\times10^{-10}$ C m = 2.57×10^{-29} C m

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Therefore, the fraction of charge covalently bonded is

$$\frac{1.26\times10^{-30}}{2.57\times10^{-29}} = 0.05$$

Therefore, % fractional charge = 5%

4. The correct order for the bond angles of NO_2, NO_2^+, NO_2^-, and NO_3^- is

(A) $NO_2^+ > NO_3^- > NO_2^- > NO_2$ (B) $NO_2^+ > NO_2 > NO_3^- > NO_2^-$

(C) $NO_3^- > NO_2 > NO_2^+ > NO_2^-$ (D) $\overset{+}{N}O_2 > NO_2^- > NO_3^- > NO_2$

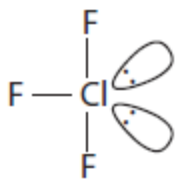
Solution

(B) NO_2^+ $-180°$, NO_2^- $-115°$, NO_2 – $132°$, NO_3^- $-120°$

5. Which of the following molecules is/are having zero dipole moment?
(I) PCl_3Br_2 (stable isomer) (II) XeF_4 (III) SF_6 (IV) ClF_3
(A) only II, IV (B) only I, II (C) only I, II, III (D) only III, IV

Solution

(C) ClF_3 has T-shaped geometry with two lone pairs, so it does not have zero dipole moment.



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6. Bond order of O_2, O_2^+, O_2^- and O_2^{2-} is in order:

(A) $O_2^- < O_2^{2-} < O_2 < O_2^+$ (B) $O_2^{2-} < O_2^- < O_2 < O_2^+$

(C) $O_2^+ < O_2 < O_2^- < O_2^{2-}$ (D) $O_2 < O_2^+ < O_2^- < O_2^{2-}$

Solution

(B) O_2: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2\ \pi^* 2p_y^1\ \pi^* 2p_z^1$

Bond order of O_2 = 1/2 (10 – 6) = 2

O_2^+: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2\ \pi^* 2p_y^1$

Bond order = (1/2) (10 – 5) = 5/2 = 2.5

O_2^-: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2\ \pi^* 2p_y^2\ \pi^* 2p_z^1$

Bond order of O_2^- = 1/2 (10 – 7) = 3/2 = 1.5.

O_2^{2-}: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2\ \pi^* 2p_y^2\ \pi^* 2p_z^2$

Bond order of O_2^{2-} = 1/2 (10 – 8) = 1

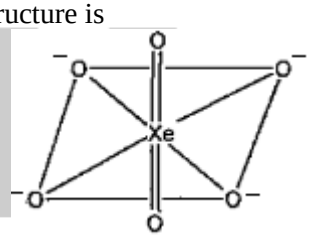
Therefore, bond order is $O_2^{2-} < O_2^- < O_2 < O_2^+$

7. The number of σ and π bonds in tetravalent perxenate anion is

(A) 6σ and 1π (B) 6σ and 2π (C) 4σ and 4π (D) 4σ and 6π

Solution

(B) Perxenate ion is $[XeO_6]^{4-}$. The structure is



8. Which of the following can be calculated from the Born–Haber cycle of formation of Al_2O_3?

(A) Lattice enthalpy of Al_2O_3. (B) Electron gain enthalpy of O atom.

(C) Hydration enthalpy of Al^{3+}. (D) Ionization enthalpy of Al.

Solution

(A) Lattice enthalpy of Al_2O_3 is calculated from Born Haber cycle using enthalpy of formation of Al(g), first, second and third ionization energies of Al(g), enthalpy of formation of O(g), first and second electron gain enthalpies of O(g) and enthalpy of formation of Al_2O_3(s).

9. The correct order of increasing C–O bond strength 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) $CO < CO_2 < CO_3^{2-}$

Solution

(A) Bond strength follows the order single bond < double bond < triple bond.

10. Hybridization, shape, and number of π bonds in $XeOF_2$ molecule are, respectively,

(A) sp^3, pyramidal, 1 (B) sp^3d, trigonal bipyramidal, 2

(C) sp^3d, T-shaped, 1 (D) sp^3d, square pyramidal, 1

Solution

(C) The structure of $XeOF_2$ molecule is T-shaped.

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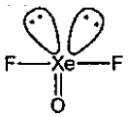


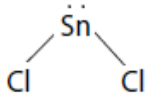
H = 3 + (1/2) (8 – 4) = 3 + 2 = 5 = sp^3d hybridization.

11. Carbon dioxide is isostructural with

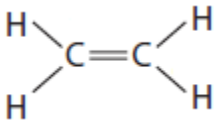
(A) $HgCl_2$ (B) $SnCl_2$ (C) C_2H_4 (D) NO_2^-

Solution

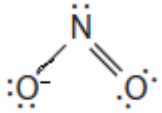
(A) O=C=O (linear) Cl–Hg–Cl (linear)



(bent)



(trigonal)



(bent)

12. Select the incorrect statement among the following:

(A) NF_3 is weaker base then NH_3.

(B) $AlCl_3$ is largely covalent while AlF_3 is largely ionic.

(C) $B_2H_6{\cdot}2NH_3$ is ionic.

(D) NO has one unpaired electron in bonding molecular orbital.

Solution

(D) MO configuration of NO: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2 \sigma 2p_x^2\ \pi 2p_x^2\ \pi 2p_y^2\ \pi^* 2p_y^1$

13. The correct order of dipole moment is

(A) $CH_4 < NF_3 < NH_3 < H_2O$ (B) $NF_3 < CH_4 < NH_3 < H_2O$

(C) $NH_3 < NF_3 < CH_4 < H_2O$ (D) $H_2O < NH_3 < NF_3 < CH_4$

Solution

(A) CH_4 – Net dipole moment is zero.

NF_3 – Resultant dipole moment toward nucleus of nitrogen.

NH_3 – Resultant dipole moment toward lone pair.

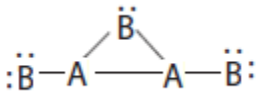
H_2O – Resultant dipole moment 2 Pcos 104.5°.

14. Element A has three electrons in the outermost orbit and B has six electrons in the outermost orbit. The formula of the compound will be

(A) A_2B_3 (B) A_2B_6 (C) A_2B (D) A_3B_2

Solution

(A). A has three electrons in the outermost orbit, it needs five more electrons to complete it octet and B has six electrons so it would need two more electrons to complete the octet. So each A will form two bonds with two B atoms, and single electron remaining on each A will share form bond with the same B atom such that one B atom is shared between the two A atoms. The structure is



15. Polarization of ions is governed by

(A) hybridization. (B) VSEPR theory.

(C) Fajans' rules. (D) Pauling's rule.

Solution

(C) Fajans' rule: Greater the polarization of anion in a molecule, more covalent the molecule will be.

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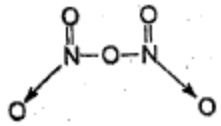
16. The pair of elements from ionic bond is
(A) C + Cl (B) H + F (C) Na + Br (D) O + H
Solution
(C) Ionization potential of sodium is less, electron gain enthalpy of Br is more. So, the type of bond formed is ionic bond.

17. The correct order of decreasing polarizable ions is
(A) Cl^-, Br^-, I^-, F^- (B) F^-, I^-, Br^-, Cl^-
(C) F^- Cl^-, Br^-, I^- (D) I^- Br^-, Cl^-, F^-
Solution
(A) Larger the anion, greater is its polarizability (because the hold on the electron cloud by the nucleus decreases).

18. The bond present in N_2O_5 is
(A) only ionic. (B) covalent and coordinate. (C) only covalent. (D) covalent and ionic.
Solution
(B) 

19. Which of these contain both polar and non-polar bonds?
(A) HCN (B) CO_2 (C) H_2O_2 (D) CH_4
Solution
(C) H–O–O–H. Here, O–H bonds is polar, O–H bonds is polar, O–O bond is non-polar. Hence, both
20. Paramagnetism is observed in
(A) N_2 (B) O_2 (C) He (D) O_2^{2-}
Solution
(B) MO configuration of O_2: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^2\ \pi^* 2p_y^1\ \pi^* 2p_z^1$
Hence, it is paramagnetic due to two unpaired electrons.
21. Which of the following species has the shortest bond length?
(A) NO^{2+} (B) NO^+ (C) NO (D) NO^-
Solution
(B) Bond order is inversely proportional to bond length:

$$\text{Bond order} \propto \frac{1}{\text{Bond length}}$$

NO^{2+}: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2 \sigma 2p_x^2\ \pi 2p_x^2\ \pi 2p_y^1$
Bond order is 2.5.
NO^+: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2 \sigma 2p_x^2\ \pi 2p_x^2\ \pi 2p_y^2$
Bond order is 3.
NO: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2 \sigma 2p_x^2\ \pi 2p_x^2\ \pi 2p_y^2\ \pi^* 2p_y^1$
Bond order is 2.5.
NO^-: $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2 \sigma 2p_x^2\ \pi 2p_x^2\ \pi 2p_y^2\ \pi^* 2p_y^1\ \pi^* 2p_z^1$
Bond order is 2.0.
So, NO^+ species has the shortest bond length.

22. Among the following, which is polar?
(A) CO_2 (B) SO_2 (C) $BeCl_2$ (D) Cl–C_6H_4–Cl (*para*)

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Solution
(B) SO_2 is a bent molecule, and has a net dipole moment; hence, it is polar.

23. The silver atom in $AgCl_2^-$ is
(A) sp^3 hybridized. (B) sp^2 hybridized. (C) *sp* hybridized. (D) unhybridized.
Solution
(C) $AgCl_2^-$ is *sp*-hybridized because it is linear in shape, and silver complexes are always linear in shape and silver is *sp*-hybridized.

24. In which pair or pairs is the stronger bond found in the first species?
(I) O_2, N_2 (II) N_2, N_2^+; (III) NO^+, NO^-
(A) I, II (B) I, III (C) I, II and III (D) II and III only
Solution
(D) $N_2 > O_2$; $N_2 > N_2^+ > NO^+ > NO^-$

25. The correct order of Cl–O bond lengths in ClO^-, ClO_2^-, ClO_3^- and ClO_4^- is
(A) $ClO^- < ClO_2^- < ClO_3^- < ClO_4^-$ (B) $ClO_4^- < ClO_3^- < ClO_2^- < ClO^-$
(C) $ClO_3^- < ClO_4^- < ClO_2^- < ClO^-$ (D) $ClO_4^- = ClO_3^- = ClO_2^- < ClO^-$
Solution
(B) Bond length is inversely proportional to bond order. Greater the delocalization of πe^- (–ve charge) shorter the bond length.
26. The C–C bond length is 1.54 Å, C=C bond length is 1.33 Å. What is the circumference of benzene ring?
(A) 8.61 Å (B) 8.82 Å (C) 8.4 Å (D) None of these
Solution
(A)
In benzene, due to the resonance C–C bond length is in between double bond and single bond. However, if we consider the six carbon atoms to be linked by three single and three double bonds, then the the circumference of benzene ring is 3 × (C=C) + 3 × (C–C) = 3 × 1.33 + 3 × 1.54 = 8.61, Å.

27. What is the effect of more electronegative atom on the strength of ionic bond?
(A) Increases (B) Decreases (C) Remains the same (D) None of these.
Solution
(A) As the electronegativity of the atom increases, the strength of the ionic bond increases because electronegative atom attracts the electron pair toward itself. If electronegativity is more, strength of ionic bond is also more.

28. Which is the weakest among the following type of bonds?
(A) Ionic bond (B) Covalent bond (C) Metallic bond (D) Hydrogen bond
Solution
(D) Hydrogen bond is the weakest bond.
29. The bond angle in ammonia molecule is
(A) 93°3' (B) 91°8' (C) 106°45' (D) 109.5°
Solution
(C) The structure of ammonia is

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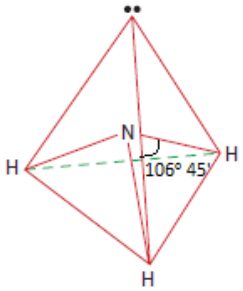
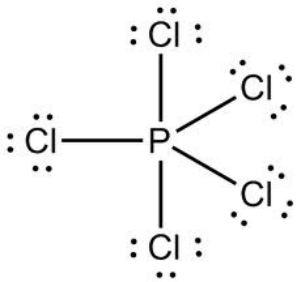


30. The correct order toward bond angle is
(A) $sp^3 < sp^2 < sp$ (B) $sp < sp^2 < sp^3$
(C) $sp < sp^3 < sp^2$ (D) bond angle does not depend on hybridization.
Solution
(A) $sp^3 < sp^2 < sp$
109.5° < 120° < 180°

31. The bond in the formation of fluorine molecule will be
(A) due to *s*–*s* overlapping. (B) due to *s*–*p* overlapping.
(C) due to *p*–*p* overlapping. (D) due to hybridization.
Solution
(C) Due to overlapping of *p*–*p* orbitals, bond exists between two fluorine atoms to form fluorine molecule.
32. Which of the following compounds does not follow the octet rule for electron distribution?
(A) H_2O (B) PH_3 (C) PCl_3 (D) PCl_5
Solution
(D) P has 10 electrons in its valence shell in PCl_5.



33. On hybridization of one *s*-orbital and one *p*-orbital, we get
(A) mutual perpendicular orbitals. (B) two orbitals at 180° to each other.
(C) four orbitals directed tetrahedrally. (D) three orbitals in one plane.
Solution
(B) *s* and *p* orbitals hybridize to form 2*sp* hybrid orbitals oriented at an angle of 180°.

34. The bond order of CO and NO is
(A) 3 and 2 (B) 3 and 2.5 (C) 3 and 1.3 (D) 3 and 3.5
Solution
(B)

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

Bond order= $\frac{\text{Number of bonding electrons (10)} - \text{Number of electrons in ABMO (4)}}{2}$ =3

NO $\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \pi 2p_x^2 = \pi 2p_y^2\ \sigma 2p_z^2 \pi^* 2p_x^1$

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

35. Which of the following has pyramidal shape?
(A) XeO_3 (B) XeF_4 (C) XeF_2 (D) XeF_8
Solution
(A) XeO_3 is sp^3 hybridized with 3 bonds in σ bonds and one lone pair and 3 π bonds. Hence, the shape is trigonal pyramid.

36. The maximum number of hydrogen bonds that can be formed by a water molecule is
(A) 2 (B) 3 (C) 4 (D) 6
Solution
(C) Each water molecule is surrounded by 4 other water molecules that are bonded by hydrogen bonds.

37. The types of bonds present in $CuSO_4 \cdot 5H_2O$ are only
(A) electrovalent and covalent. (B) electrovalent and coordinate covalent.
(C) electrovalent, covalent and coordinate covalent. (D) covalent and coordinate covalent.
Solution
(C) In $CuSO_4 \cdot 5H_2O$, electrovalent, covalent, and coordinate covalent bonds are present.

38. Shape of NH_3 is very similar to
(A) CH_4 (B) CH_3^- (C) BH_3 (D) CH_3^+
Solution
(B) CH_3^- is sp^3-hybridized with 3 bond pairs and 1 lone pair. Thus, the shape is trigonal pyramidal like NH_3.

39. Which of the following is not diamagnetic?
(A) C_2 (B) O_2^{2-} (C) Li_2 (D) N_2^+
Solution
(D) In N_2^+, there is one unpaired electron present in bonding molecular orbital; rest all have paired.

40. Which of the following is arranged in order of increasing dipole moment?
(A) $BCl_3 < NH_3 < H_2O < SO_2$ (B) $BCl_3 < NH_3 < SO_2 < H_2O$
(C) $NH_3 < SO_2 < H_2O < BCl_3$ (D) $H_2O < SO_2 < NH_3 < BCl_3$
Solution
(B) Dipole moment is defined as the product of the magnitude of charge on any one of the atoms and the distance between them.

$$BCl_3 < NH_3 < SO_2 < H_2O$$

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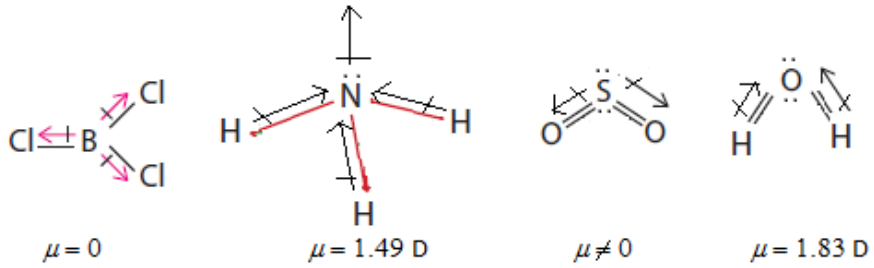

The dipole moment of BCl_3 molecule is zero. The bond dipoles of three B–Cl bonds give a net sum of zero because the resultant of any two is equal and opposite to the third (parallelogram law of forces) as shown above. In other molecules dipole moment shows increasing trend due to the presence of central atom that is electronegative in nature. As the electronegativity of the central metal atom increases (N < S < O), dipole moment of the resulting molecule also increases. Hence, this shows the above trend.

41. Which is steam volatile?

(A) *o*-Nitrophenol (B) Aniline (C) Glycerol (D) *p*-Nitrophenol

Solution

(A) *o*-Nitrophenol due to intramolecular hydrogen bonding, boiling point is low. Hence, it is steam volatile, whereas *para* has intermolecular H-bonding.

42. A solid processing high electrical and thermal conductivities is

(A) NaCl (B) Si (C) Ice (D) Li

Solution

(D) Metals have high electrical and thermal conductivity in solid state.

Multiple Correct Choice Type

1. Polarity of a molecule is calculated in terms of dipole moment pointing an arrow from +ve to –ve pole. Find out the correct statement among the following.

(A) Carbony1 group shows greater dipole moment than carbon monoxide.

(B) In carbon monoxide C atom forms – ve pole of dipole whereas O atom forms + ve pole even O is more electronegative than C.

(C) Dipole moment of CH_3Cl is greater than CH_3F.

(D) All homonuclear polyatomic molecules are non-polar.

Solution

(A, B, C)

2. Which among the following have regular geometry?

(A) BF_3 (B) NF_3 (C) PF_3 (D) BF_4^-

Solution

(A, D) BF_3 has a regular trigonal planar structure and BF_4^- has a regular tetrahedral structure. Both NF_3 and PF_3 have pyramidal structure.

3. CO_2 is isostructural with

(A) $HgCl_2$ (B) $SnCl_2$ (C) C_2H_2 (D) NO_2

Solution

(A, C) CO_2 is isostructural with $HgCl_2$ and C_2H_2.

All these are linear in structure and process *sp* hybridization.

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4. Which of the following statement is/are correct?
(A) The bond energy of B–X is more in BX_3 than in BX_3 pyridine.
(B) When BF_3 is added to Na_3AlF_6, Na_3BF_4 is formed.
(C) All alkaline earths form peroxides on heating in excess of oxygen.
(D) Magnesium cannot be extracted by the electrolysis of fused $Mg(OH)_2$.
Solution
(A, B, D)

5. Which among the following species have zero bond order?
(A) F_2^{2-} (B) Ar_2 (C) He_2^+ (D) H_2^+
Solution
(A, B) MO configuration of F_2^{2-}:
$\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \sigma 2p_x^2\ \pi^2 2p_y\ \pi^2 2p_z\ \pi^* 2p_y^2\ \pi^* 2p_z^2\ \sigma^* 2p_x^2$
Bond order = (1/2) (10 – 10) = 0
Ar_2 also has zero bond order because it is a noble gas; it exists only monoatomic in nature.

6. Which of the following molecules are polar and sp^3d^2 hybridized?
(A) IF_5 (B) XeF_4 (C) $XeOF_4$ (D) ICl_4^-
Solution
(A, C)
IF_5: H = 5 + (1/2) (7 – 5) = 5 + 1 = 6 = sp^3d^2
XeF_4: H = 4 + (1/2) (8 – 4) = 4 + 2 = sp^3d^2
$XeOF_4$: H = 5 + (1/2) (8 – 6) = sp^3d^2
ICl_4^-: H = 4 + (1/2) (7 – 4 + 1) = 5 = sp^3d
But XeF_4 is a square structure that is symmetrical in nature. As a result dipole moments cancel each other and it becomes non-polar. ICl_4^- is polar structure but it is sp^3d hybridized.

7. Which of the following combination(s) of atomic orbitals suggest non-bonding combinations of atomic orbitals?

Solution
(C, D)

Non-bonding combinations of atomic orbitals.

8. Correct order about bond angle is
(A) $H_2O > H_2S > H_2Se > H_2Te$ (B) $C_2H_2 > C_2H_4 > CH_4 > NH_3$
(C) $SF_6 < NH_3 < H_2O < OF_2$ (D) $ClO_2 > H_2O > H_2S > SF_6$
Solution
(A, B, D)
$H_2O > H_2S > H_2Se > H_2Te$
$C_2H_2 > C_2H_4 > CH_4 > NH_3$

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180° 120° 109° 107
$ClO_2 > H_2O > H_2S > SF_6$

9. Which of following energy terms are associated with the Born–Haber cycle of dissociation of $BaCl_2$ in water?
(A) Hydration enthalpy of Ba^{2+}.
(B) Lattice enthalpy of $BaCl_2$.
(C) Sublimation enthalpy of Ba.
(D) Electron gain enthalpy of Cl^-.
Solution
(A, B)

10. In which of the following, the hybrid orbitals of the central atom have the same *s* character?
(A) CH_4 (B) $Ni(CO)_4$ (C) XeO_3 (D) $[Ni(CN)_4]^{2-}$
Solution
(A, C) CH_4 and XeO_3 both are in sp^3 hybridization. So, the hybrid orbitals of both central atom have the same *s*-character.

11. Which of the following set correctly represent sp^3 hybridization and tetrahedral geometry?
(A) ClO_4^-, SO_4^{2-}, Diamond
(B) SiO_4^{4-}, $[PCl_4]^+$, SiC
(C) $Ni(CO)_4$, $SnCl_4$, Graphite
(D) Ice, $FClO_3$, $[Zn(H_2O)_4]^{+2}$
Solution
(A, B, D) $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \sigma 2p_z < \pi 2p_y = \pi 2p_x < \pi^* 2p_y = \pi^* 2p_x < \sigma^* 2p_z$
Energy order of molecular orbitals in oxygen molecule

12. Among the following bonds, which are more ionic than covalent?
(A) Si–O (B) Ba–O (C) Se–Cl (D) K–Br
Solution
(B, D) If the electronegativity difference between the two atoms is more than 1.9, the bond is taken as ionic. Here, (B) and (D) have an electronegativity difference greater than 1.9.

13. The linear structure is assumed by
(A) $SnCl_2$ (B) NCO^- (C) NO_2^+ (D) CS_2
Solution
(A, C, D) $SnCl_2$: Bent or V-shape
NCO^-: V-shape
NO_2^+: Linear shape
CS_2: Bent or V-shape

Assertion–Reasoning Type
Choose the correct option from the following:
(A) Statement 1 is True, Statement 2 is True; Statement 2 is a correct explanation for Statement 1.
(B) Statement 1 is True, Statement 2 is True; Statement 2 is not a correct explanation of Statement 1.
(C) Statement 1 is True; Statement 2 is False.
(D) Statement 1 is False; Statement 2 is True.
1. Statement 1: Magnetic moment of super oxide ion is much larger than that of peroxide ion.
Statement 2: Peroxide ion has no unpaired electrons but superoxide ion has unpaired electrons.
Solution
(A)
2. Statement 1: There are 10 valence electrons on the sulphur atom in SF_4 molecule.

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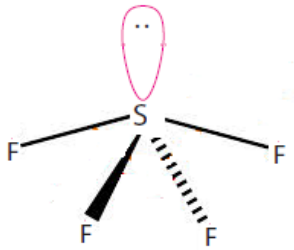
Statement 2: The structure of SF_4 molecule is based on a distorted trigonal bipyramidal.
Solution
(B) Due to sp^3d hybridization the fluorine atomic orbitals overlap with the sulphur atomic orbitals of the sulphur atom to form SF_4 molecule.

Sulphur in ground state: [↑↓] [↑↓][↑][↑] [ ][ ][ ][ ][ ]

Sulphur in excited state: [↑↓] [↑][↑][↑] [↑][ ][ ][ ][ ]

Hybridization: [↑↓] [↑×][↑×][↑×] [↑×][ ][ ][ ][ ]

Total ten electrons on sulphur atom.
The shape of SF_4 is see-saw shaped



3. Statement 1: BF_3 molecule is planar but NF_3 is pyramidal.
Statement 2: N atom is smaller than B.
Solution
(B) BF_3 molecule is planar because it undergoes sp^2 hybridization and NF_3 molecule is pyramidal in shape because in this the central atom, nitrogen undergoes sp^3 hybridization.
4. Statement 1: *o*-Nitrophenol has a higher boiling point than *p*-nitrophenol.
Statement 2: Intramolecular hydrogen bonding occurs in *p*-nitrophenol.
Solution
(D) -Nitrophenol has intramolecular hydrogen bonding which weaker than intermolecular hydrogen bonding in *p*-nitrophenol.

5. Statement 1: σ Bond is strong while π bond is a weak bond.
Statement 2: Atoms rotate freely about π bond.
Solution
(C) σ Bond is strong while π bond is weak because in σ bond formation overlapping of orbitals takes place along their internuclear axis, giving rise to maximum electron density on the axis. So σ bond is strong.

6. Statement 1: Fluorine molecule has bond order 1.
Statement 2: The number of electrons in antibonding molecular orbital is two less than that of bonding molecular orbitals.
Solution
(A) F_2 = $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \ = \pi 2p_y^2 \pi^* 2p_x^2 \ = \pi^* 2p_y^2$

Bond order = (10 – 8)/2 = 1

7. Statement 1: In tetrahedral hybridization, that is, in sp^3 hybridization all *p* orbitals are involved and no *p* orbital is left for forming π bonds.
Statement 2: Central atom cannot form double bonds in the molecules or species having sp^3 hybridization.
Solution

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(C) Because in sp^3 hybridization there are no vacant orbitals as well as unpaired electrons to form π bond.

8. Statement 1: The boiling point of NH_3 lies between that of SbH_3 and BiH_3.
Statement 2: PH_3 has much lower boiling point than NH_3, but it increases from PH_3 to AsH_3 to SbH_3 to BiH_3 due to increase in van der Waals forces.
Solution
(D) Except for ammonia the boiling point increases as we move down the group. The abnormally high boiling point of NH_3 is due to the association of molecules through hydrogen bonding and boiling point increases down the group due to increase in van der Waals forces.

PH_3	<	A SH_3	<	NH_3	<	SbH_3	<	BiH_3
–87.5°C		–62.45°C		–34.5C		–18.45°C		+16.8 5°C

9. Statement 1: Decrease in lattice enthalpy is not as sharp as increase in interionic distance.
Statement 2: Change in interionic distance affects the structure, which compensates the otherwise change in lattice enthalpy.
Solution
(A)

10. Statement 1: NH_3 has low melting point (e.g., −77 °C) in comparison to H_2O.
Statement 2: NH_3 molecules do not have intermolecular force as high as imparted by intermolecular hydrogen bonding.
Solution
(C) As hydrogen bonding in ammonia is less compared to that in water and also nitrogen is less electronegative than oxygen, so NH_3 has low melting point than water.

11. Statement 1: Both $\pi 2p_x$ and $\pi^* 2p_x$ molecular orbitals have one nodal plane each.
Statement 2: All molecular orbitals formed by sideway overlapping of $2p$ orbitals have one nodal plane.
Solution
(D)
$\pi(2p_x)$ has one nodal plane whereas $\pi^*(2p_x)$ has two nodal planes. Bonding π-molecular orbitals have one nodal plane whereas antibonding π-molecular orbitals have two nodal planes.

<H2>Comprehension Type
Read the paragraphs and answer the questions that follow.
Paragraph I
The following steps help predict the geometrical shape of species with only one central atom, without first drawing the Lewis structure:
(A) Total number of σ electron pairs = (Number of valence electrons ± Electrons (for ionic charges)/2
(B) Number of bonded electron pairs = Number of atoms – 1
(C) Number of electron pairs around central atom = Total number of electron pairs – 3 [Number of terminal atoms (except H)]
(D) Number of lone pairs = (Number of central electron pairs – Number of bond pairs)
1. Based on above method, structure of the some of the molecules have been matched. Which is the incorrect matching?
(A) ICl_2^- – Linear
(B) ClO_3^- – Pyramidal
(C) ICl_4^- – Square planar
(D) $PdCl_4^{2-}$ – Tetrahedral
Solution

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(D) The shape of $PdCl_4^{-2}$ is square planar.

2. If A is the central element of the molecule containing A and X elements and *E* is the number of lone pairs around it, then VSEPR notation AX_3E will be for the molecules:

(A) PCl_3, ICl_4^- (B) NH_3, H_2O (C) NH_3, ClO_3^- (D) ICl_4^-, ClO_3^-

Solution

(C) NH_3 and ClO_3^- have one lone pair of electrons.

3. Square planar structure is predicated for:

(A) ICl_4^-, ClO_3^- (B) XeF_4, XeO_4 (C) BrF_5, Al_2Cl_6 (D) ICl_4^-, $PtCl_4^{2-}$

Solution

(D)

Paragraph II

Valence shell electron pair repulsion (VSEPR) theory can be used to predict the approximate shape of a molecule. Electrons in bonds and in lone pairs can be thought of as "charge cloud" that repel one another and stay as far apart possible, thus causing molecules to assume specific shapes.

The repulsive interactions of electron pairs decrease in the order:

Lone pair–lone pair > Lone pair–bond pair > Bond pair–bond pair.

These repulsions effect result in deviations from idealized shapes and alteration in bond angles in molecules.

4. Molecular shape of XeF_3^+, SF_3^+, and CF_3^+ are

(A) the same with 2, 1, and 0 lone pairs of electrons, respectively.
(B) different with 2, 1, and 0 lone pairs of electrons, respectively.
(C) different with 0, 1, and 2 lone pairs of electrons, respectively.
(D) the same with 2, 0, and 1 lone pairs of electrons, respectively.

Solution

(B) XeF_3^+ : H = 3 + (1/2) (8 – 3 – 1) = sp^3d

$AB_5 \rightarrow AB_2L_2 \rightarrow \text{Two lone pairs}$

SF_3^+ = 3 + (1/2) (6 – 3 – 1) = sp^3

$AB_4 \rightarrow AB_3L \rightarrow \text{One lone pair}$

CF_3^+ : 3 + (1/2) (4 – 3 – 1) = 3 = sp^2

$AB_3 \rightarrow \text{Zero lone pair}$

5. Which of the following statement is correct with respect to bond angle?

(A) The F–S–F angle in SF_2 is more than 109.5°.
(B) The H–N–N angle in N_2H_2 is approximately 180°.
(C) The F–Kr–F angle in KrF_4 is 90°.
(D) The Cl–N–O angle in NOCl is more than 120°.

Solution

(C) The structure of KrF_4 is

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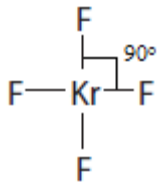
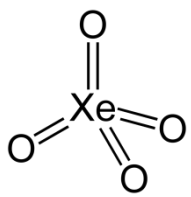


6. Which of the following statements is incorrect?
(A) In ClF_3, the axial Cl–F bond length is larger than equatorial Cl–F bond length.
(B) In SF_4, F–S–F equatorial bond angle is not 120° and 104° due to lone pair–bond pair repulsions.
(C) In ICl_4^-, bond angles is 90°.
(D) In OBr_2, the bond angle is less than OCl_2.
Solution
(D) In OBr_2, the bond angle is more than OCl_2 because Br is less electronegative element than chlorine.

Passage III
In 1962, Bartlett prepared the first real compound of noble gases. Soon after this, there was a rapid extension of the chemistry of the noble gases and in particular of xenon. Xenon reacts directly with fluorine when heated at 400°C in a sealed nickel tube. The products obtained depend on the ratio of Xe and F_2. When 1:20 mixture of Xe and F_2 is heated, it gives XeF_6, which is a white solid and undergoes slow hydrolysis in small quantity of water forming a xenon oxyfluoride (A) with Xe, O and F weight ratio 2.62: 0.32: 1.52. It undergoes complete hydrolysis in excess of water forming an oxide (B) in the same oxidation state. When the oxide (B) reacts with (A) it forms another compound (C) with formula XeO_2F_2. XeF_6 can act as a fluoride donor and forms a complex (D) with AsF_5. The complex in the crystalline form is found to contain monovalent ions in which Xe is present in the cationic part. (Xe = 131).

7. In the partial hydrolysis of XeF_6 the hybridization of Xe changes from
(A) sp^3d^2 to sp^3d (B) sp^3d^3 to sp^3 (C) sp^3d^3 to sp^3d (D) sp^3d^3 to sp^3d^2
Solution
(D) $XeF_6 + H_2O \rightarrow XeOF4 + 2HF$
sp^3d^3 sp^3d^2

8. The shape of the molecule (B) is
(A) tetrahedral. (B) pyramidal. (C) octahedral. (D) angular bent.
Solution
(B) $XeF_6 + 3H_2O \rightarrow XeO_3$ (B) + 6HF
The structure of XeO_3 is tetrahedral



9. The correct order of bond angles in the compounds A and B is
(A) A > B (B) A < B (C) A ≠ B = 90° (D) A = B = 90°
Solution
(B) In XeF_6 (A), the bond angle is 90°.

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In XeO_3 (B), the bond angle is 103°.

Paragraph IV

The average distance between the nuclei of the two atoms held by a bond is called bond length. This represents the internuclear distance corresponding to the minimum potential energy for the system. The main factors that affect the bond length are as follows:

(A) Double/triple bonds are shorter than corresponding single bonds.

(B) Sometimes, bond distances of single bonds are somewhat larger than those of double bonds of their respective covalent radii (e.g., F_2). It is due to strong repulsive interaction between the lone-pair electrons on adjacent atoms.

(C) Sometimes single bond distances are somewhat shorter than double of their respective covalent radii because bonds acquire some partial double bond character. This occurs when one atom has vacant orbital and another atom has a lone pair. The bond length may also become shorter due to high ionic character in the covalent bond.

10. Which is not true about the N–N bond length among the following species?

(I) $H_2N–NH_2$ (II) N_2 (III) $HN^+–NH_2$ (IV) N_2O

(A) N–N bond length is the shortest in II.

(B) N–N bond length in I is shorter than that of in III.

(C) N–N bond length in III is shorter than that of in I.

(D) N–N bond length IV is intermediate between I and III.

Solution

(A) Because of triple bond.

11. The bond angle is Cl_2O is more than that in F_2O because of so many reasons like

(I) $p\pi–d\pi$ overlap in Cl_2O, which is not possible in F_2O.

(II) There is no lone pair–lone pair repulsion in Cl_2O, which is present in F_2O.

(III) F is more electronegative than oxygen while Cl is less electronegative than oxygen.

The correct reason is (are)

(A) both I and II. (B) only I.

(C) only II. (D) All of these.

Solution

(D)

12. Which is correct about the C–F bond?

(A) C–F is shorter than expected because of lone pair–lone pair repulsion.

(B) C–F is longer than expected because of lone pair–lone pair repulsion.

(C) C–F is shorter than expected because of high ionic character.

(D) C–F is longer than expected because of high ionic character.

Solution

(C) In accordance with Fajans' rule.

13. The correct order about of C–O bond length in the following is

(I) CO (II) CO_2 (III) CO_3^{2-}

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

Solution

(C) The order is as follows:

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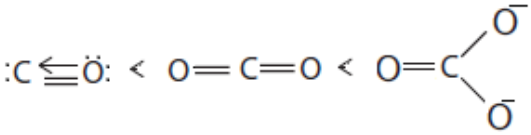
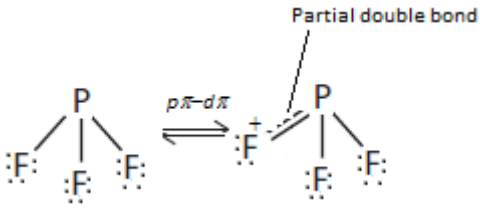


14. In which of the following cases, the central atom–F bond has partial double bond character?
(A) NF_3 (B) CF_4 (C) PF_3 (D) OF_2
Solution
(C) The PF_3 molecule has lone pair of electrons on F atom and P has vacant *d* orbitals. The small size of F atom leads to $p\pi$–$d\pi$ bonding in PF_3.



Passage V
Lattice energy ($\Delta H_{lattice}$) is given by Born–Lande equation

$$\Delta H_{lattice} = 1.39\times10^5 \text{ kJ mol}^{-1}\text{pm}\left(\frac{Z_+Z_-A}{r_0}\right)\left(1-\frac{1}{n}\right)$$

where A = Madelung constant, n = Born exponent, Z_+ and Z_- are the ionic charges, r_0 = internuclear distance between oppositely charged ions.

15. Among LiF, CaF_2, and MgS the increasing order of lattice energy (only value without sign) is
(A) MgS > CaF_2 > LiF (B) LiF > MgS > CaF_2
(C) CaF_2 > LiF > MgS (D) LiF > MgF_2 > CaS
Solution
(A) The correct order is

MgS > CaF_2 > LiF
(+2)(−2) (+2)(−1) (+1)(−1)
777 623 240 Kcal mol^{-1}

16. Among LiF, LiCl, LiBr, LiI, the lattice energy (only value without sign) decrease in the order
(A) LiI > LiBr > LiCl > LiF (B) LiF > LiCl > LiBr > LiI
(C) LiCl > LiF > LiBr > LiI (D) LiCl > LiBr > LiI > LiF
Solution
(B) LiF > LiCl > LiBr > LiI
247 203 193 180 kcal mol^{-1}

17. When $\Delta H_{hyd} = f_1\left(\frac{1}{r_+}\right) + f_2\left(\frac{1}{r_-}\right)$ then choose the correct statement
(A) LiI and CsF are the most insoluble of alkali metal halides.
(B) AgF is insoluble in water.
(C) LiF and CsI are the most insoluble alkali metal halides.
(D) $AgClO_4$ is insoluble in hot water, but soluble in benzene.

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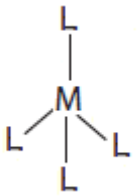
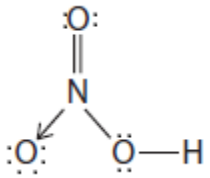
Solution
(C) LiF and CsI ($r^+ = r^-$). Lattice energy is more than hydration energy.

Integer Answer Type
The answer is a **non-negative** integer.
1. The molecule ML_x is planar with 7 pairs of electrons around M in the valence shell. The value of x is ___.
Solution
(4) The structure is



2. The number of lone pair(s) present in the structure of HNO_3 is___.
Solution
(7) The structure showing lone pairs is:



3. The number of 90° bond angles present in the molecule of SF_4 is___.
Solution
(0) SF_4 is seesaw shaped. Its bond angles are 186° and 116°.

4. The number of ions among the following having pseudo-inert gas configuration is ___.
Zn^{2+}, Cd^{2+}, Hg^{2+}, Cu^+, Ag^+, Au^+
Solution
(6) Zn^{2+}, Cd^{2+}, Hg^{2+}, Cu^+, Ag^+, Au^+ all these show pseudo-inert gas configuration.

5. In $[Ag(CN)_2]^-$, the number of π bonds is ___.
Solution
(4) $[\overset{1}{\underset{2}{N}}\equiv C-Ag-C\overset{3}{\underset{4}{\equiv}}N]^-$

6. The sum of oxidation states of N atom in the following compounds of nitrogen is ___.
Solution
(+9) Total = 1 – 3 + (–2) + (–1) + 0 + 1 + 2 + 3 + 4 + 5 = +9

$$\overset{-3}{N}H_3,\ \overset{-2}{N_2}H_4,\ \overset{-1}{N}H_2OH,\ \overset{0}{N_2},\ \overset{+1}{N_2}O,\ \overset{+2}{N}O,\ \overset{+3}{N_2}O_3,\ \overset{+4}{N}O_2,\ \overset{+5}{N_2}O_5$$

7. The number of unpaired electrons in a paramagnetic diatomic molecule of an element with atomic number 16 is ___.
Solution
(2)
16 electrons in MO configuration are filled as

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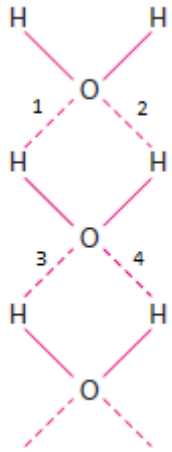
$\sigma 1s^2\ \sigma^* 1s^2\ \sigma 2s^2\ \sigma^* 2s^2\ \sigma 2p_z^2\ \pi 2p_x^2\ \pi 2p_y^2\ \pi^* 2p_x^1\ \pi^* 2p_y^1$

8. The maximum number of hydrogen bonds possible in which a water molecule can participate is ___.

Solution

(4) One water molecule can participate in 4 hydrogen bonds as shown:



Matrix–Match Type

1. Match the species with their properties.

Column I	**Column II**
(A) O_2^+	(p) Isoelectronic with N_2
(B) CO	(q) Fractional bond order
(C) NO	(r) paramagnetic
(D) He_2^+	(s) Diamagnetic

Solution

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

O_2^- : Bond order is 1.5 – paramagnetic.

CO: Bond order is 3 – diamagnetic.

NO^+: Bond order is 3 – diamagnetic.

He_2^+ : Bond order is 0.5 – paramagnetic.

2. Match the molecules with hybridization of central atom.

Column I	**Column II**
(A) $IO_2F_2^-$	(p) sp^3d
(B) F_2SeO	(q) sp^3
(C) $ClOF_3$	(r) sp^2
(D) XeF_5^+	(s) sp^3d^2

Solution

A → (p); B → (q); C → (p); D → (s)

(A) $IO_2F_2^-$: H = 4 + (1/2) (7 – 6 + 1) = 4 + 1 = 5 = sp^3d

(B) F_2SeO: H = 3 + (1/2) (6 – 4) = 4 = sp^3

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(C) $ClOF_3$: H = 4 + (1/2) (7 – 5) = 4 + 1 = 5 = sp^3d

(D) XeF_5^+ : H = 5 + (1/2) (8 – 5 – 1) = 6 = sp^3d^2

3. Match the pair of compounds with their identical property.

Column I	**Column II**
(A) PCl_3F_2, PCl_2F_3	(p) Hybridization of central atom
(B) BF_3, BCl_3	(q) Shape of molecule/ion
(C) CO_2, CN_2^{2-}	(r) μ (dipole moment)
(D) C_6H_6, $B_3N_3H_6$	(s) Total number of electrons

Solution

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

(A) Both PCl_3F_2 and PCl_2F_3 have sp^3d hybridization with trigonal bipyramidal shape.

(B) Both BF_3 and BCl_3 have sp^2 shape with trigonal planar shape and also the dipole moments are zero in both the cases (dipole moment cancels each other).

(C) Both CO_2 and CN_2^{2-} have sp hybridization with linear geometry, zero dipole moment, and also have same number of electrons, that is, 22.

(D) C_6H_6 known as benzene and $B_3N_3H_6$ known as borazine are planar molecules with aromatic character; B and C atoms undergo sp^2 hybridization, have equal dipole moments and total number of electrons is 30.

4. Match the molecules/species with the properties.

Column I	**Column II**
(A) BrF_5	(p) It has at least one angle less than 90°.
(B) ICl_3	(q) The central atom is sp^3d hybridized.
(C) H_3O^+	(r) It is non-planar.
	(s) The central atom is having only one lone pair.

Solution

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

$$\text{Bond order} = \frac{1}{2}[2-1] = \frac{1}{2}$$

Since He_2^+ has one unpaired electron, thus He_2^+ is paramagnetic in nature and it also has fractional bond order.

(A) O_2^{2-} : The electronic configuration of O is $1s^2\ 2s^2\ 2p^4$. The MO formed during combination of two O atoms occupy 18 electrons because of the presence of two extra electrons, leading to MO configuration of:

$$O_2^{2-} : \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_x^2, \pi 2p_z^2, \pi 2p_y^2, \pi^* 2p_z^2, \pi^* 2p_y^2$$

Since O_2^{2-} does not have unpaired electron, O_2^{2-} is diamagnetic.

(B) CO: It contains total 14 electrons: 6 electrons from carbon and 8 electrons from oxygen. It is isoelectronic with N_2 because N_2 also contains total 14 electrons.

MO configuration of CO: $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_x^2, \pi 2p_z^2, \pi 2p_y^2$

Since CO does not have unpaired electron, it is diamagnetic.

(C) NO: Electronic configuration of N and O are $1s^2\ 2s^2\ 2p^3$ and $1s^2\ 2s^2\ 2p^4$, respectively. The MO formed during combination of N and O atoms occupy 15 electrons leading to

MO configuration of NO: $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_x^2, \pi 2p_z^2, \pi 2p_y^2, \pi^* 2p_z^1$

Bond order of NO: $\frac{1}{2}[10-5] = \frac{5}{2}$

Since NO has one unpaired electron, NO is paramagnetic and it also has fractional bond order.

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(D) He_2^+: It contains total 3 electrons.

MO configuration of He_2^+: $\sigma 1s^2$, $\sigma^* 1s^2$

5. Match the hybrid bond orbitals with the species.

Column I	**Column II**
(A) sp^3	(p) AsF_6^-
(B) dsp^2	(q) $TeCl_4$
(C) sp^3d	(r) SiO_2
(D) sp^3d^2	(s) $[Ni(en)_2]^{2+}$

Solution

A → (r); B → (s); C → (q); D → (p)

Using the formula $H = S + (1/2)(E - V \pm C)$

(A) AsF_6^-: H = 6 + (1/2) (5 – 6 + 1) = 6 = sp^3d^2

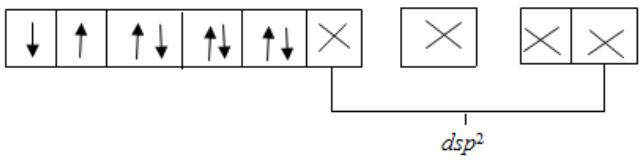
(B) $TeCl_4$: H = 4 + (1/2) (6 – 4) = 5 = sp^3d

(C) SiO_2: H = 4(as O is divalent) + (1/2) (4 – 4) = 4 = sp^3

(D) Ground state configuration of Ni = $3d^8\ 4s^2$

Electronic configuration of Ni^{2+} = $3d^8$

The hybridization in $[Ni(en)_2]^{2+}$



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