

Chapter 1

Problem 1P

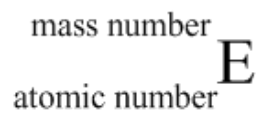
Oxygen has three isotopes with mass numbers of 16, 17, and 18. The atomic number of oxygen is eight. How many protons and neutrons does each of the isotopes have?

Step-by-step solution

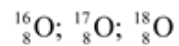
step 1 of 4

Isotopes have the same atomic number (that is, the same number of protons), but different mass numbers because they have different number of neutrons.

Representation of an element, E with its atomic number and mass number is:



The three isotopes of oxygen are:



We know that,

Number of protons in an isotope = atomic number of the element

Number of neutrons in an isotope = mass number of isotope - atomic number of isotope

step 2 of 4

These three isotopes have the same number of protons or atomic number. That is 8.

Number of neutrons in ${}^{16}_8\text{O}$ isotope,

$$= 16 - 8$$

$$= 8$$

Therefore, the number of neutrons are present in the ${}^{16}_8\text{O}$ isotope is 8.

step 3 of 4

Number of neutrons in ${}^{17}_8\text{O}$ isotope,

$$= 17 - 8$$

$$= 9$$

Therefore, the number of neutrons are present in the ${}^{17}_8\text{O}$ isotope is 9.

step 4 of 4

Number of neutrons in ${}^{18}_8\text{O}$ isotope,

$$= 18 - 8$$

$$= 10$$

Therefore, the number of neutrons are present in the $^{18}_8\text{O}$ isotope is 10.

Problem 2P

Step-by-step solution

step 1 of 5

step 2 of 5

Atoms contains sub atomic articles, they are electrons, protons and neutrons, electrons are negatively charged articles revolves round the nucleus, nucleus contains positively charged protons and neutral charged neutrons

In a neutral atom the number of protons is equals to number of electrons and this number is in turn equal to the atomic number of the atom

step 3 of 5

Ions of the atoms are formed either by the addition of electrons or removal of electrons, so the protons number will remain constant, so we can say that either in ionic or neutral state the number of protons in an atom is equal to its atomic number

step 4 of 5

(a)

1. Given element is sodium ion Na^+ , the atomic number of the sodium is 11, so in a neutral sodium atom the number of protons will be 11, as protons of an atom remain constant and which is equal to its atomic number, even though the sodium is in ionic state the number of protons in it is equal to its atomic number which is 11

2. Given element is argon Ar , the atomic number of the argon is 18, so in a neutral argon atom the number of protons will be 18

3. Given element is chlorine ion Cl^- , the atomic number of the chlorine is 17, so in a neutral chlorine atom the number of protons will be 17, as protons of an atom remain constant and which is equal to its atomic number, even though the chlorine is in ionic state the number of protons in it is equal to its atomic number which is 17

step 5 of 5

(b)

1. Given element is sodium ion Na^+ , the atomic number of the sodium is 11, so in a neutral sodium atom the number of electrons will be 11, as sodium is in ionic state the sodium ion is formed by the loss of one electron, so the number of electrons resent in Na^+ is 10

2. Given element is argon Ar , the atomic number of the argon is 18, so in a neutral argon atom the number of electrons will be 18

3. Given element is chlorine ion Cl^- , the atomic number of the chlorine is 17, so in a neutral chlorine atom the number of electrons will be 17, here chlorine is in ionic state, this ion is formed by the gain of one electron, so the number of electrons in the Cl^- will be 18

Problem 3P

Chlorine has two isotopes, ^{35}Cl and ^{37}Cl ; 75.77% of chlorine is ^{35}Cl and 24.23% is ^{37}Cl . The atomic mass of ^{35}Cl is 34.969 amu and the atomic mass of ^{37}Cl is 36.966 amu. What is the atomic weight of chlorine?

Step-by-step solution

step 1 of 1

The atomic mass of ^{35}Cl is 34.969 amu and its abundance is 75.77 %, the atomic mass of ^{37}Cl is 36.966 amu and its abundance is 24.23 %.

The atomic weight of the chlorine calculated as follows:

$$\begin{aligned}\text{The atomic weight of chlorine} &= (\text{first isotopic abundance} \times \text{its atomic mass}) \\ &\quad + (\text{second isotopic abundance} \times \text{its atomic mass}) \\ &= (0.7577 \times 34.969 \text{ amu}) + (0.2423 \times 36.966 \text{ amu}) \\ &= 26.4960 \text{ amu} + 8.9576 \\ &= 35.5 \text{ amu}\end{aligned}$$

Therefore, chlorine atomic weight is 35.45 amu

Problem 4P

How many valence electrons do the following atoms have?

a. boron b. nitrogen c. oxygen d. fluorine

Step-by-step solution

step 1 of 4

Electrons in the outer most shell are called valence electrons.

a) Electronic configuration of boron, $B - 1s^2 \underline{2s^2 2p^1}$
Hence no. of valence electrons in $B = 2 + 1$
 $= 3$

step 2 of 4

(b) Electronic configuration of nitrogen, $N - 1s^2 \underline{2s^2 2p^3}$
Hence no. of valence electrons in $N = 2 + 3$
 $= 5$

step 3 of 4

(c) Electronic configuration of oxygen, $O - 1s^2 \underline{2s^2 2p^4}$
Hence no. of valence electrons in $O = 2 + 4$
 $= 6$

step 4 of 4

(d) Electronic configuration of fluorine, $F - 1s^2 \underline{2s^2 2p^5}$
Hence no. of valence electrons in $F = 2 + 5$
 $= 7$

Problem 5P

Step-by-step solution

step 1 of 4

step 2 of 4

Atoms contains sub atomic articles, they are electrons, protons and neutrons, electrons are negatively charged articles revolves round the nucleus, nucleus contains positively charged protons and neutral charged neutrons

In a neutral atom the number of protons is equals to number of electrons and this number is in turn equal to the atomic number of the atom

The electrons in the atoms are arranged in orbitals, electrons tends to occupy the available lowest energy orbital

The rules useful for arrangement of electrons in orbital as follows

1. Aufbau Principle: electrons tend to occupai the lowest energy orbital, the atomic orbital closer the nucleus has lower energy, so $1s$ orbital has lower energy than $2s$ orbital, so the energy order will be $1s < 2s < 3s < 4s$

s-orbital has lower energy than p-orbital, similarly the energy of orbital follows the order $s < p < d < f$

so the relative energy of the orbital will be $1s < 2s < 2p < 3s < 3p < 3d$

2. Pauli exclusion principle: each atomic orbital occupies a maximum of two electrons, more than two electrons shall not be occupied y an atomic orbital, the two electrons must be in opposite spin, same sin electrons should not occupai in a single atomic orbital. s-orbital has only one sub atomic orbital so an s-orbital can occupai a maximum of two electrons, p-orbital has three degenerate(equal energy) sub atomic orbital

(p_x, p_y, p_z) , so the p-orbital can occupai a maximum of 6 electrons, d-orbital has three degenerate sub atomic orbital $(d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}, d_{z^2})$, so the d-orbital can occupai a maximum of 10 electrons

3. Hund's rule: in case of degenerate orbital the all the degenerate obitals must be half filled then only pairing of electrons takes place

step 3 of 4

1. The atomic number of chlorine is 17, so it will have 17 electrons, the electrons can be arranged according to the above rules as follows, $1s^2 2s^2 2p^6 3s^2 3p^5$, the ground state electron configuration can be denoted including spin orientation of electrons as follows

$1s \uparrow\downarrow 2s \uparrow\downarrow 2p \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow 3s \uparrow\downarrow 3p \uparrow\downarrow \uparrow\downarrow \uparrow$, the inner electron configuration of atoms is same as the electron configuration of noble gas with lower atomic number than the atom, so simplified ground state electron configuration will be $[\text{Ne}] 3s \uparrow\downarrow 3p \uparrow\downarrow \uparrow\downarrow \uparrow$

2. The atomic number of bromine is 35, so it will have 35 electrons, the electrons can be arranged according to the above rules as follows, $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$, the nearest noble gas have lower atomic number than bromine is argon, argon atomic number is 18, so the simplified electron configuration will be

$[\text{Ar}] 4s^2 3d^{10} 4p^5$ the ground state electron configuration can be denoted including spin orientation of electrons as follows

$[\text{Ar}] 4s \uparrow\downarrow 3d \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow 4p \uparrow\downarrow \uparrow\downarrow \uparrow$

3. The atomic number of iodine is 53, so it will have 53 electrons, the electrons can be arranged according to the above rules as follows, $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$, the nearest noble gas have lower atomic number than iodine is krypton, krypton atomic number is 36, so the simplified electron configuration will be

$[\text{Kr}] 5s^2 4d^{10} 5p^5$ the ground state electron configuration can be denoted including spin orientation of electrons as follows

$[\text{Kr}] 5s \uparrow\downarrow 4d \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow 5p \uparrow\downarrow \uparrow\downarrow \uparrow$

step 4 of 4

The electrons present in the outer most orbital are called valance electrons, the chemical property of the atom depends upon the valance electrons

1. The electron configuration of chlorine is $[\text{Ne}]3s\uparrow\downarrow 3p\uparrow\downarrow\uparrow\downarrow\uparrow$, the outer most shell is 3, the electrons present outermost shell are $3s^2 3p^5$, so a total of 7 valance electros present in chlorine

2. The electron configuration of bromine is $[\text{Ar}]4s\uparrow\downarrow 3d\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow 4p\uparrow\downarrow\uparrow\downarrow\uparrow$, the outer most shell is 4, the electrons present outermost shell are $4s^2 4p^5$, so a total of 7 valance electros present in bromine

3. The electron configuration of iodine is $[\text{Kr}]5s\uparrow\downarrow 4d\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow 5p\uparrow\downarrow\uparrow\downarrow\uparrow$, the outer most shell is 5, the electrons present outermost shell are $5s^2 5p^5$, so a total of 7 valance electros present in iodine

Problem 6P

Look at the relative positions of each pair of atoms listed here in the periodic table. How many core electrons does each have? How many valence electrons does each have?

- a. carbon and silicon c. nitrogen and phosphorus
b. oxygen and sulfur d. magnesium and calcium

Step-by-step solution

step 1 of 4

Electrons in inner shell are called core electrons. Electrons in outermost shell are called valance electrons.

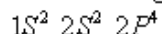
- (a) The electronic configuration of carbon is $1s^2 2s^2 2p^2$.
So, carbon has two core electrons and four valance electrons.
The electronic configuration of silicon is
 $1s^2 2s^2 2p^6 3s^2 3p^2$
So, silicon has ten core electrons and four valance electrons.

step 2 of 4

- (b) The electronic configuration of nitrogen is:
 $1s^2 2s^2 2p^3$
So, nitrogen has two core electrons and five valance electrons.
The electronic configuration of phosphorus is
 $1s^2 2s^2 2p^6 3s^2 3p^3$
So, phosphorus has ten core electrons and five valance electrons.

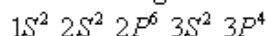
step 3 of 4

(c) The electronic configuration of oxygen is:



So, oxygen has two core electrons and six valence electrons.

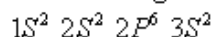
The electronic configuration of sulfur is:



So, sulfur has ten core electrons and six valence electrons.

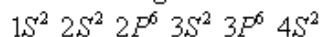
step 4 of 4

(d) The electronic configuration of magnesium is:



So, magnesium has 10 core electrons and two valence electrons.

The electronic configuration of calcium is:



So, calcium has eighteen core electrons and two valence electrons.

Problem 7P

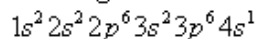
a. Find potassium (K) in the periodic table and predict how many valence electrons it has.

b. What orbital does the unpaired electron occupy?

Step-by-step solution

step 1 of 2

(A) Electrons configuration of K (atomic no 19) is-



Its outer most shell is 4 and no. of valence electrons = 1.

Hence, its position in the periodic table -4th period ,group 1.

step 2 of 2

(B) The unpaired electrons occupies s-orbital.

Problem 8P

Which bond is more polar?

a. H—CH₃ or Cl—CH₃

b. H—OH or H—H

c. H—Cl or H—F

d. Cl—Cl or Cl—CH₃

Step-by-step solution

step 1 of 4

Polarity of a bond is generated due to electro negativity difference between the two bonded atoms. Electro negativity is the ability of an atom to pull the bonding electrons of a covalent bond towards itself.

- (A) $Cl-CH_3$ is more polar than $H-CH_3$ as the electro negativity difference between $-Cl$ and $-C$ is more than that between $-H$ and $-C$.

step 2 of 4

- (B) $H-OH$ bond is more polar than $H-H$ bond. Because, there will be no electro negativity difference between the two H -atoms as $H-H$ as they are identical in all respect.

step 3 of 4

- (C) $H-F$ will be more polar than $H-Cl$ because F atoms is more electro negative than Cl atom due to its smaller size.

step 4 of 4

- (D) $Cl-CH_3$ is more polar than $Cl-Cl$ because electro negativity difference exists in $Cl-CH_3$ but it is absent in $Cl-Cl$ where two identical atoms are connected

Problem 9P

Which of the following has

- a. the most polar bond?

NaI $LiBr$ Cl_2 KCl

- b. the least polar bond?

Step-by-step solution

step 1 of 2

- (a) KCl has the most polar bond.
Reason- The electro negativity difference is highest in KCl in which K is most electropositive and Cl is the most electronegative element among the others given in the form of compounds.

step 2 of 2

- (b) Cl_2 has the least-polar bond.
Reason- Two identical chlorine atoms are connected where no electro negativity difference exist. Hence polarity of Cl_2 is zero.

Problem 11P

Use the symbols δ^+ and δ^- to show the direction of the polarity of the indicated bond in each of the following compounds:

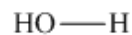
δ^+ and δ^-

Step-by-step solution

step 1 of 6

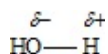
a)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this oxygen (3.5) is more electronegative than hydrogen (2.1).

Therefore, the oxygen has partial negative (δ^-) charge and hydrogen has partial positive charge (δ^+).



b)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this fluorine (4.0) is more electronegative than bromine (2.8).

Therefore, the fluorine has partial negative charge (δ^-) and bromine has partial positive charge (δ^+).



step 2 of 6

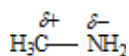
c)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this nitrogen (3.0) is more electronegative than carbon (2.5).

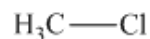
Therefore, the nitrogen has partial negative charge (δ^-) and carbon has partial positive charge (δ^+).



step 3 of 6

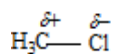
d)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this chlorine (3.0) is more electronegative than carbon (2.5).

Therefore, the chlorine has partial negative charge (δ^-) and carbon has partial positive charge (δ^+).



step 4 of 6

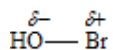
e)

The provided bond is as follows:



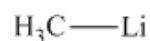
The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this oxygen (3.5) is more electronegative than bromine (2.8).

Therefore, the oxygen has partial negative charge (δ^-) and bromine has partial positive charge (δ^+).



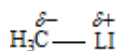
f)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this carbon (2.5) is more electronegative than lithium (1.0).

Therefore, the carbon has partial negative charge (δ^-) and lithium has partial positive charge (δ^+).



step 5 of 6

g)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this chlorine (3.0) is more electronegative than iodine (2.5).

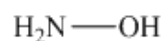
Therefore, the chlorine has partial negative charge (δ^-) and iodine has partial positive charge (δ^+).



step 6 of 6

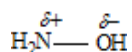
h)

The provided bond is as follows:



The more electronegative atom has partial negative charge (δ^-) and less negative atom have partial positive charge (δ^+). In this oxygen (3.5) is more electronegative than nitrogen (3.0).

Therefore, the oxygen has partial negative charge (δ^-) and nitrogen has partial positive charge (δ^+).



Problem 13P

Explain why HCl has a smaller dipole moment than HF, even though the H-Cl bond is longer than the H-F bond.

Step-by-step solution

step 1 of 1

The dipole moment (μ) of a bond is equal to the magnitude of the charge (e) on the atom (either the partial positive charge or the partial negative charge, because they have the same magnitude) times the distance between the two charges (d).

So, dipole moment = $\mu = e \times d$.

Although the $\text{H}-\text{Cl}$ bond distance is greater than $\text{H}-\text{F}$ bond distance, due to the much greater partial charge on the atoms of $\text{H}-\text{F}$, HCl has a smaller dipole moment than that of HF .

Problem 14P

After examining the potential maps for LiH, HF, and H_2 , answer the following questions:

- Which compounds are polar?
- Why does LiH have the largest hydrogen?
- Which compound has the hydrogen that would be most apt to attract a negatively charged molecule?

Step-by-step solution

step 1 of 3

- (A) LiH and HF are polar as electro negatively difference between the bonded atoms exists in both the compounds.

step 2 of 3

- (B) Hydrogen in LiH has greatest electrons density as LiH exists as $\overset{\delta+}{\text{Li}}-\overset{\delta-}{\text{H}}$. That's why LiH have the Largest hydrogen.

step 3 of 3

- (C) Hydrogen in HF would be the most *apt* to attract a negatively charged molecule because HF exists as $\overset{\delta+}{\text{H}}-\overset{\delta-}{\text{F}}$ where electrons density is least on the hydrogen atom making the hydrogen smallest in size.

Problem 15P

The formal charge does not necessarily indicate that the atom has greater or less electron density than other atoms in the molecule without formal charges. You can see this by examining the potential

maps for H_2O , H_3O^+ , and HO^- .

a. Which atom bears the formal negative charge in the hydroxide ion?

b. Which atom has the greater electron density in the hydroxide ion?

c. Which atom bears the formal positive charge in the hydronium ion?

d. Which atom has the least electron density in the hydronium ion?

Step-by-step solution

step 1 of 4

$$\text{Formal on an atom} = \text{no. of valence electrons} - \left(\frac{\text{no. of lone pair electrons} +}{2} \times \text{no. of bonding electrons} \right)$$

(A) Formal charge on H atoms in $\text{HO}^- = \left[1 - \left(0 + \frac{1}{2} \times 2 \right) \right] = [1 - 1] = 0$

$$\text{Formal charge on O atoms in } \text{HO}^- = \left[6 - \left(3 \times 2 + \frac{1}{2} \times 2 \right) \right] = [6 - 7] = -1$$

So, oxygen atoms bear's formal negative charge in hydroxide ion.

step 2 of 4

(B) Oxygen atom has the greater electron density in hydroxide ion as it is more Electronegative with a negative charge which support the higher electron density.

step 3 of 4

(C) Formal charge on H atoms in $\text{H}_3\text{O}^+ = \left[1 - \left(0 + \frac{1}{2} \times 2 \right) \right] = [1 - 1] = 0$

$$\text{Formal charge on O atoms in } \text{H}_3\text{O}^+ = \left[6 - \left(1 \times 2 + \frac{1}{2} \times 6 \right) \right] = [6 - 5] = +1$$

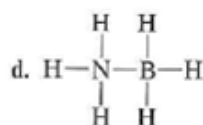
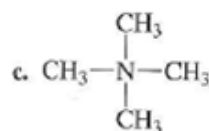
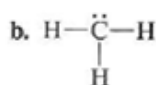
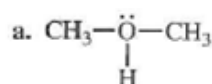
So, oxygen atoms bears formal positive charge in hydroxium ion.

step 4 of 4

(D) Oxygen atom has the learnt electron density as it forms co-ordinate co-relent bonds with hydrogen and thereby electron density is reduced.

Problem 16P

Give each atom the appropriate formal charge:

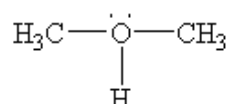


Step-by-step solution

step 1 of 4

Formal charge on an atom = no. of valence electrons - (no. of lone pair of electrons + $\frac{1}{2} \times$ no. of bonding electrons)

(A) In



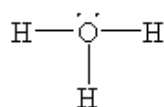
$$\text{Formal charge on each H atoms} = \left[1 - \left(0 + \frac{1}{2} \times 2 \right) \right] = [1 - 1] = 0$$

$$\text{Formal charge on each C atom} = \left[4 - \left(0 + \frac{1}{2} \times 8 \right) \right] = [4 - 4] = 0$$

$$\text{Formal charge on O atoms} = \left[6 - \left(2 + \frac{1}{2} \times 6 \right) \right] = [6 - 5] = +1$$

step 2 of 4

(B) In

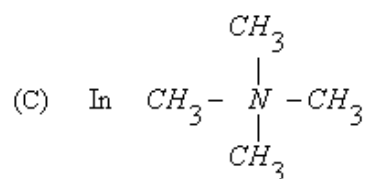


figure

$$\text{Formal charge on each H atoms} = \left[1 - \left(0 + \frac{1}{2} \times 2 \right) \right] = [1 - 1] = 0$$

$$\text{Formal charge on C atom} = \left[4 - \left(2 + \frac{1}{2} \times 6 \right) \right] = [4 - 5] = -1$$

step 3 of 4

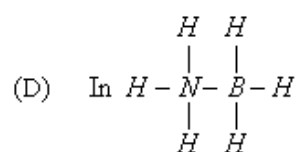


$$\text{Formal charge on each H atoms} = \left[1 - \left(0 + \frac{1}{2} \times 2 \right) \right] = [1 - 1] = 0$$

$$\text{Formal charge on each C atoms} = \left[4 - \left(0 + \frac{1}{2} \times 8 \right) \right] = [4 - 4] = 0$$

$$\text{Formal charge on N atom} = \left[5 - \left(0 + \frac{1}{2} \times 8 \right) \right] = [5 - 4] = +1$$

step 4 of 4



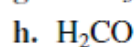
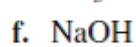
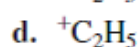
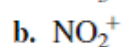
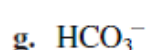
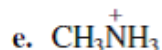
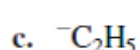
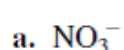
$$\text{Formal charge on each H atoms} = \left[1 - \left(0 + \frac{1}{2} \times 2 \right) \right] = [1 - 1] = 0$$

$$\text{Formal charge on N atom} = \left[5 - \left(0 + \frac{1}{2} \times 8 \right) \right] = [5 - 4] = +1$$

$$\text{Formal charge on B atom} = \left[3 - \left(0 + \frac{1}{2} \times 8 \right) \right] = [3 - 4] = -1$$

Problem 17P

Draw the Lewis structure for each of the following:



Step-by-step solution

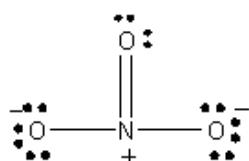
step 1 of 9

While drawing the structure, two main things are to be considered-

- Valency of C, N and O are 4, 3 and 2 respectively. These valences are to be filled in by drawing bonds.
- Lone pair of electrons on N and O atoms are 1 and 2 pairs respectively.

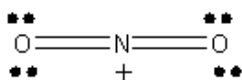
step 2 of 9

a)



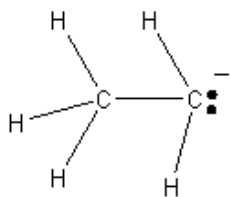
step 3 of 9

b)



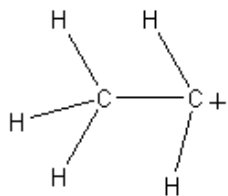
step 4 of 9

c)



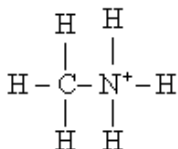
step 5 of 9

d)



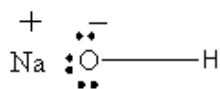
step 6 of 9

e)



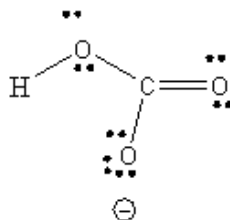
step 7 of 9

f)



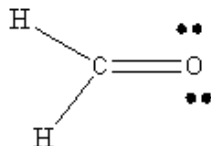
step 8 of 9

g)



step 9 of 9

h)



Problem 18P

a. Draw two Lewis structures for C_2H_6O .

b. Draw three Lewis structures for C_3H_8O .

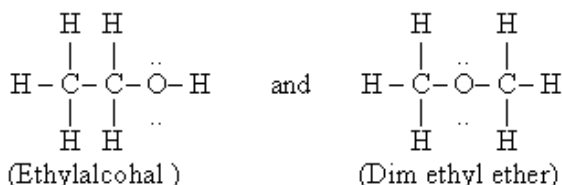
(Hint: The two Lewis structures in part a are constitutional isomers; they have the same atoms, but differ in the way the atoms are connected; see page 143. The three Lewis structures in part b are also constitutional isomers.)

Step-by-step solution

step 1 of 2

The compounds with same chemical formula but different in structure are called constitutional isomers.

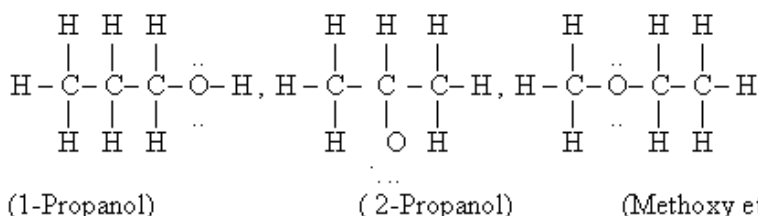
(A) The two Lewis structures for C_2H_6O –



These two are functional isomers as the functional groups are $-OH$ i.e., alcohol and $-O-$ i.e., ether respectively.

step 2 of 2

(B) The three Lewis structure for C_3H_8O –



The first two are metamers i.e., they have same $(-OH)$ functional group but the positions are different.

The 3rd one is having functional isomeric relationship with the first two as the 3rd one has a different functional group $(-O-)$ i.e., ether.

Problem 19P

Draw the lone-pair electrons that are not shown in the following condensed structures:

- | | | |
|-------------------|-----------------|-----------------|
| a. $CH_3CH_2NH_2$ | c. CH_3CH_2OH | e. CH_3CH_2Cl |
| b. CH_3NHCH_3 | d. CH_3OCH_3 | f. $HONH_2$ |

Step-by-step solution

step 1 of 6

Usually in most of the cases lone pairs on nitrogen oxygen and halogen atoms are 1,2 and 3 respectively.

- (A) $H_3CCH_2\ddot{N}H_2$ - Here valency of N atom is 3, hence it contains one lone pair of electron to Satisfy octate.

step 2 of 6

- (B) $H_3C\ddot{N}HCH_3$ - Again valency of N atom is 3, so it contains one lone pair of electrons to satisfy octate.

step 3 of 6

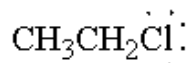
- (C) $H_3CCH_2\ddot{O}H$ - Here valency of O atom is 2, it contains two lone pair of elections to obtain octate configuration.

step 4 of 6

- (D) $H_3C\ddot{O}CH_3$ - Here O atom is connected with two bonds, so it contains two lone pair of electrons to satisfy octate.

step 5 of 6

- (E)



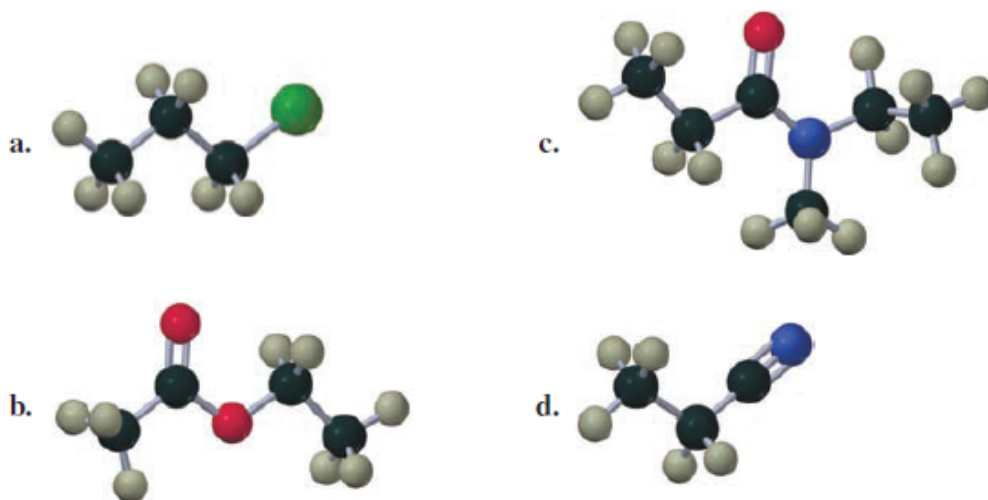
Here the Cl atoms form only one bonds, hence it contains there lone pair of electrons to have octate.

step 6 of 6

- (F) $H\ddot{O}\ddot{N}H_2$ - Here O and N atoms have two and one lone pair of elections receptively.

Problem 20P

Draw condensed structures for the compounds represented by the following models (black = C, gray = H, red = O, blue = N, and green = Cl):



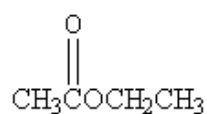
Step-by-step solution

step 1 of 4

- (a) The structures are simplified by omitting some of the covalent bonds and listing the atoms bonded to a particular carbon (or nitrogen or oxygen) next to it with subscripts as necessary. These structures are called condensed structures.
 $\text{CH}_3\text{CH}_2\text{CH}_2\text{Cl}$

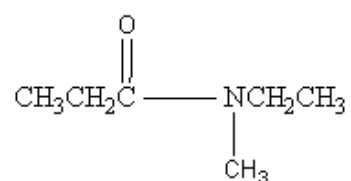
step 2 of 4

(b)



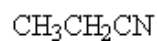
step 3 of 4

(c)



step 4 of 4

(d)



Problem 21P

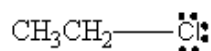
Which of the atoms in the molecular models in Problem 19 have

a. three lone pairs? b. two lone pairs? c. one lone pair? d. no lone pairs?

Step-by-step solution

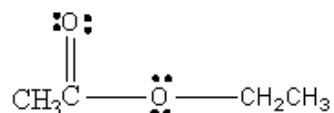
step 1 of 4

- (a) Chlorine atom has three lone Pairs.



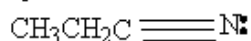
step 2 of 4

- (b) Oxygen atom has two lone Pairs.



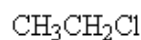
step 3 of 4

- (c) Nitrogen atom has one lone pairs.



step 4 of 4

- (d) Carbon atom has no lone Pairs.



Problem 22P

Step-by-step solution

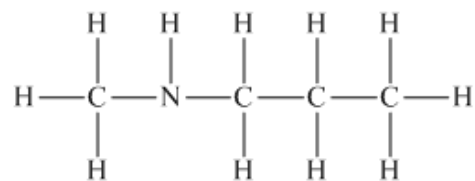
step 1 of 5

step 2 of 5

In a kekule structure the bonds between atoms denoted by straight lines similar to Lewis structure, in case of kekule structure representation of loan air of electrons is omitted

a)

The kekule structure for $\text{CH}_3\text{NH}(\text{CH}_2)_2\text{CH}_3$ can be drawn as follows, the repeated CH_2 groups are denoted in parentheses

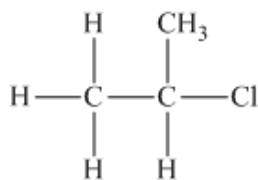


N-methylpropan-1-amine

step 3 of 5

b)

The kekule structure for $(\text{CH}_3)_2\text{CHCl}$ can be drawn as follows, the identical CH_3 groups on terminal carbon of the chain are denoted in parentheses

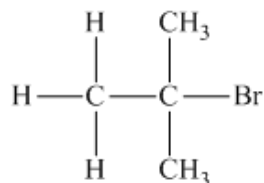


2-chloropropane

step 4 of 5

c)

The kekulé structure for $(\text{CH}_3)_3\text{CBr}$ can be drawn as follows, the identical CH_3 groups on terminal carbon of the chain are denoted in parentheses

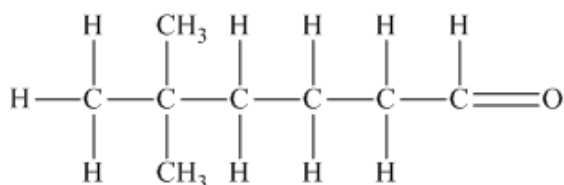


2-bromo-2-methylpropane

step 5 of 5

d)

The kekulé structure for $(\text{CH}_3)_3\text{C}(\text{CH}_2)_3\text{CHO}$ can be drawn as follows, the repeated CH_2 groups are denoted in parentheses, , the identical CH_3 groups on terminal carbon of the chain are denoted in parentheses



5,5-dimethylhexanal

Problem 23P

Step-by-step solution

step 1 of 4

In skeletal structure we need to draw the molecule in vertex model, no need to represent the hydrogen atoms bonded to carbon atom, need not to represent the lone pair of electrons present on the atoms

a)

let us draw the condensed structure for the model, from the condensed structure we can draw the skeletal structure

The condensed structure for model is $\text{CH}_3(\text{CH}_2)_2\text{Cl}$, now skeletal structure can be drawn as follows, the repeated CH_2 groups are denoted in parentheses



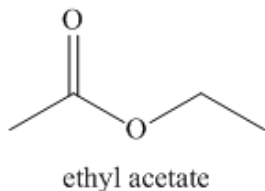
1-chloropropane

step 2 of 4

b)

let us draw the condensed structure for the model, from the condensed structure we can draw the skeletal structure

The condensed structure for model is $\text{CH}_3\text{CO}_2\text{CH}_2\text{CH}_3$, now skeletal structure can be drawn as follows, the group CO_2 present in the structure is ester group

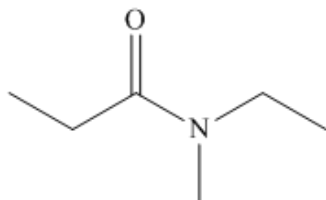


step 3 of 4

c)

let us draw the condensed structure for the model, from the condensed structure we can draw the skeletal structure

The condensed structure for model is $\text{CH}_3\text{CH}_2\text{CON}(\text{CH}_3)\text{CH}_2\text{CH}_3$, now skeletal structure can be drawn as follows, the group CH_3 attached to nitrogen is represented in parentheses



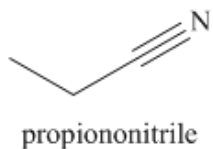
N-ethyl-*N*-methylpropionamide

step 4 of 4

d)

let us draw the condensed structure for the model, from the condensed structure we can draw the skeletal structure

The condensed structure for model is $\text{CH}_3\text{CH}_2\text{CN}$, now skeletal structure can be drawn as follows, there is triple bond between terminal carbon and nitrogen



Problem 24P

Draw a picture of

a. a 3s orbital. b. a 4s orbital. c. a 3p orbital.

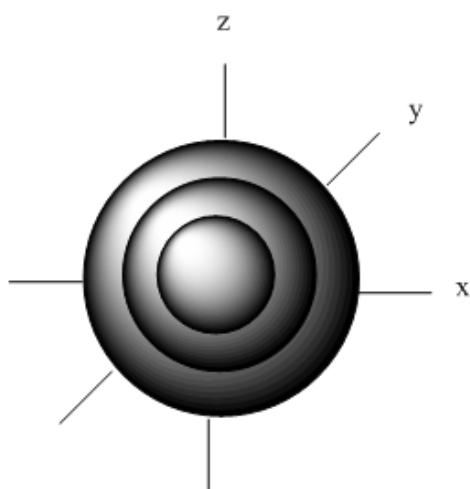
Step-by-step solution

step 1 of 3

a)

The **3s** orbital have two nodes.

The picture of $3s$ with nodes is as follows:

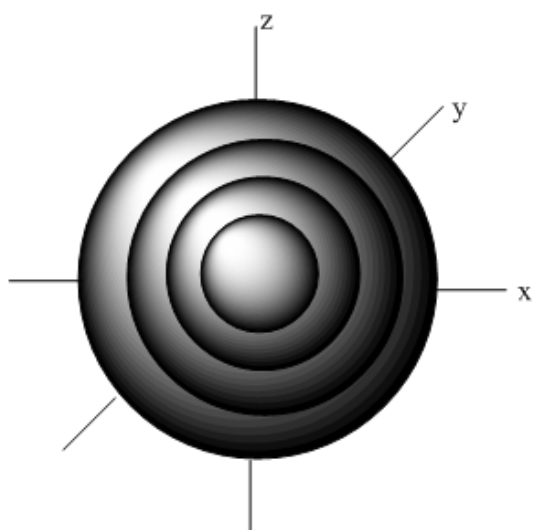


step 2 of 3

b)

The $4s$ orbital have two nodes.

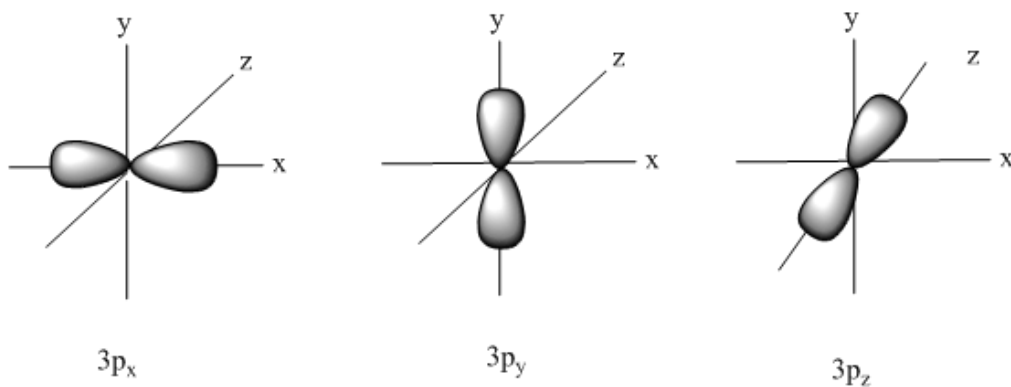
The picture of $3s$ with nodes is as follows:



step 3 of 3

c)

The picture of $3p$ orbital is as follows:



Step-by-step solution

Problem 25P

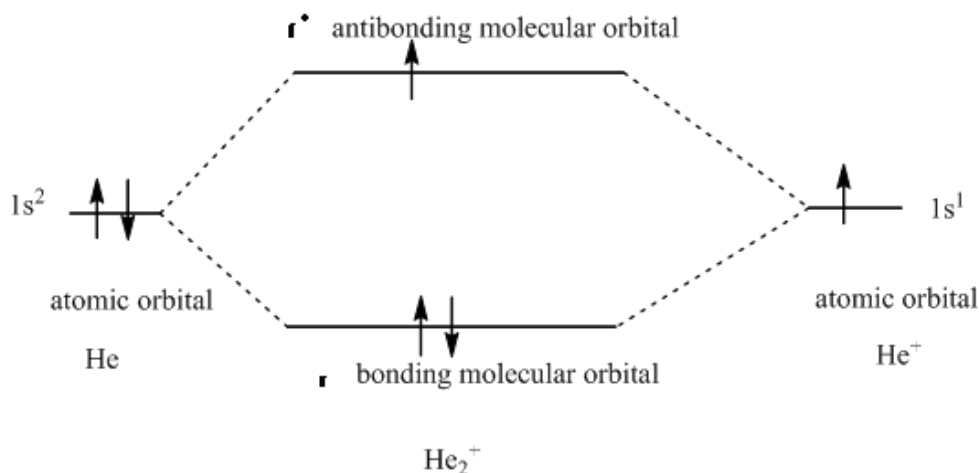
step 1 of 1

Molecular orbitals are formed by the overlap of atomic orbitals, due to this overlapping the energy of the atomic orbitals decreases, so stability of the molecule increases, the number of atomic orbitals overlapping is equal to the number of

F molecular orbitals formed, constructive overlapping of orbitals gives bonding molecular orbital, destructive overlapping of orbitals gives antibonding orbitals,

Bonding orbitals have lower energy than antibonding orbitals, the electrons present in the bonding orbitals stabilise the molecule, whereas the electrons present in the antibonding orbitals destabilise the molecule, so the molecule contains more electrons in bonding orbitals when compared to antibonding orbitals will be stable

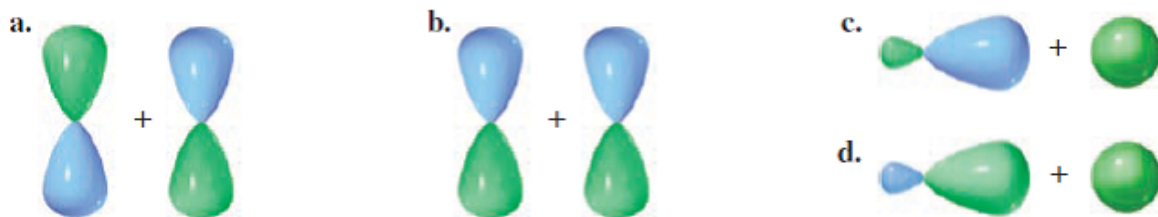
Let us draw the molecular orbital diagram for the He_2^+



There are two electrons in bonding molecular orbitals and one electron in antibonding molecular orbital, so the molecule has stability so there is a possibility of existing He_2^+

Problem 26P

Indicate the kind of molecular orbital (σ , σ^* , π , or π^*) that results when atomic orbitals are combined as indicated:



Step-by-step solution

step 1 of 5

When two orbitals overlap end-on, the result will be σ and σ^* molecular orbitals.

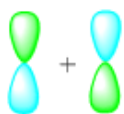
When two orbitals overlap side-to-side, the result would be π and π^* molecular orbitals.

When spherically symmetrical s-orbitals combine with any other orbital then the result would be σ and σ^* molecular orbitals.

step 2 of 5

(a)

The orbitals are as follows:



Here, the green colored lobe of p-orbital overlaps with the blue colored lobe of another p-orbital and results in formation of π^* antibonding orbital.

step 3 of 5

(b)

The orbitals are as follows:



Here, the blue colored lobe p-orbital overlaps with the green colored lobe of s-orbital and results in forming the antibonding σ^* molecular orbital.

step 4 of 5

(c)

The orbitals are as follows:

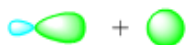


The two p-orbitals overlap end-on forming σ and σ^* molecular orbitals. Here, the green colored lobe of p-orbital overlaps with the blue colored lobe of another p-orbital and results in formation of σ^* antibonding orbital.

step 5 of 5

(d)

The orbitals are as follows:



In this figure, the green color lobe of the p-orbital overlaps with the green color lobe of s-orbital and forms σ bonding orbital.

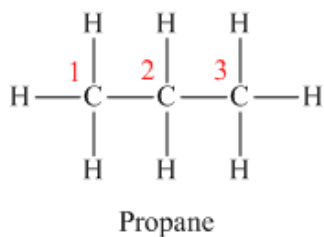
Problem 27P

What orbitals are used to form the 10 covalent bonds in propane ($\text{CH}_3\text{CH}_2\text{CH}_3$)?

Step-by-step solution

step 1 of 2

The three carbon atoms in propane ($\text{CH}_3\text{CH}_2\text{CH}_3$) are tetrahedral. Each carbon uses four sp^3 orbitals to form four covalent bonds.

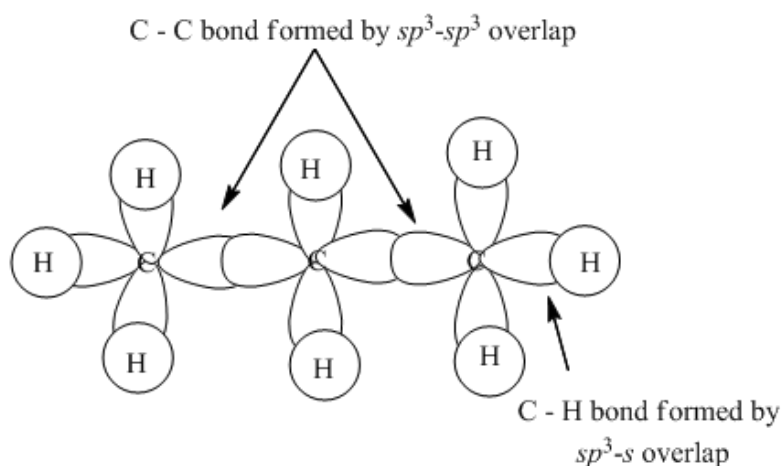


One sp^3 orbital of one carbon overlaps an sp^3 orbital of other carbon to form the $C-C$ bond. Each of the remaining three sp^3 orbitals of each carbon overlaps the s orbital of hydrogen to form a $C-H$ bond. Thus, the $C-C$ bond is formed by sp^3-sp^3 overlap, and each $C-H$ bond is formed by sp^3-s overlap.

step 2 of 2

All the bonds in propane are sigma (σ) bonds because they are all formed by the end-on overlap of atomic orbitals.

The orbital diagram of propane is:



Therefore, orbitals are used form the 10 covalent bonds in propane are sp^3-sp^3 and sp^3-s orbitals.

Problem 28P

Explain why a sigma bond formed by overlap of an s orbital with an sp^3 orbital of carbon is stronger than a sigma bond formed by overlap of an s orbital with a p orbital of carbon.

Step-by-step solution

step 1 of 2

When two atoms approach each other, their atomic orbital's overlap which is the main factor for the formation of covalent bonds in homo nuclear or hetero nuclear diatomic molecules.

Sigma bond forms by the end to end overlap of atomic orbitals. Its strength depends on the extent of overlap of orbitals. If more is the extent of overlapping more will be the bond strength. Overlapping of pure atomic orbitals is not effective in bond formation.

step 2 of 2

A sigma bond formed by the overlapping of pure s and p orbital of C is not strong, because s orbital is spherical and p orbital is dumb-bell shape. So effective overlapping do not take place therefore, lesser is the bond strength.

But in sp^3 hybridization four sp^3 orbitals are equivalent in energy and shape, so hybrid orbitals are more

effective in bond formation. Thus, sigma bond formed by the overlap of s orbital with sp^3 orbital of carbon atom is stronger than that with p orbital.

Problem 30P

Put a number in each of the blanks:

- a. ___ s orbital and ___ p orbitals form ____ sp^3 orbitals.
- b. ___ s orbital and ___ p orbitals form ____ sp^2 orbitals.
- c. ___ s orbital and ___ p orbitals form ____ sp orbitals.

Step-by-step solution

step 1 of 3

a)

Atomic orbitals are merging to form hybrid orbitals. Hybridization is nothing but the merging of atomic orbitals.

One s orbital combines with three p orbitals to form a four sp^3 hybrid orbitals.

step 2 of 3

b)

Atomic orbitals are merging to form hybrid orbitals. Hybridization is nothing but the merging of atomic orbitals.

One s orbital combines with two p orbitals to form a three sp^2 hybrid orbitals.

step 3 of 3

c)

Atomic orbitals are merging to form hybrid orbitals. Hybridization is nothing but the merging of atomic orbitals.

One s orbital combines with one p orbital to form a two sp hybrid orbitals.

Problem 32P

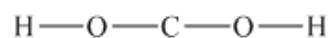
Step-by-step solution

step 1 of 9

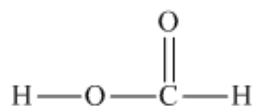
a)

1. The molecule is H_2CO_2

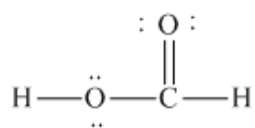
Arrange the atoms in the molecule with single bonds.



The bonds for carbon is not satisfied, make double bond with one of the oxygen, make bond between carbon and hydrogen.



Arrange the loan pair of electrons on the oxygen atoms, so the structure can be drawn as follows:

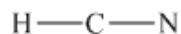


formic acid

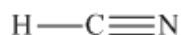
step 2 of 9

2. The molecule is HCN

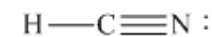
Arrange the atoms in the molecule with single bonds.



The bonds for carbon, nitrogen are not satisfied; make triple bond between carbon and nitrogen.



Arrange the loan pair of electrons on the nitrogen atom, so the structure can be drawn as follows:

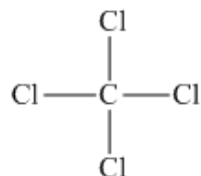


hydrogen cyanide

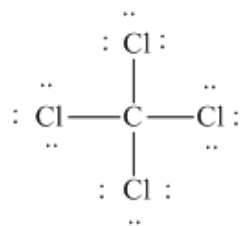
step 3 of 9

3. The molecule is CCl_4

Arrange the atoms in the molecule with single bonds.



Arrange the loan pair of electrons on the chlorine atom, so the structure can be drawn as follows:



tetrachloromethane

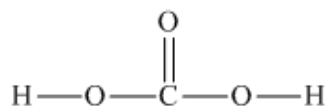
step 4 of 9

4. The molecule is H_2CO_3

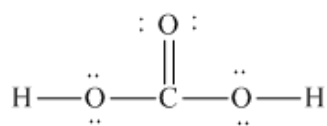
Arrange the atoms in the molecule with single bonds.



The bonds for carbon is not satisfied, make double bond with one of the oxygen.



Arrange the lone pair of electrons on the oxygen atoms, so the structure can be drawn as follows:

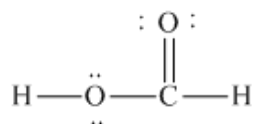


carbonic acid

step 5 of 9

b)

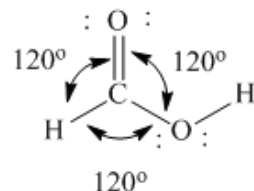
1. The molecule is H_2CO_2 , the Lewis structure is as follows:



formic acid

step 6 of 9

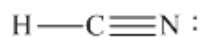
In formic acid carbon forms three sigma bonds and one pi bond, so the carbon is in sp^2 hybridisation, so the bond angle will be 120° , so the structure will be as follows:



formic acid

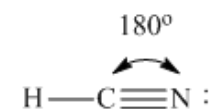
step 7 of 9

2. The molecule is HCN , the structure can be drawn as follows



hydrogen cyanide

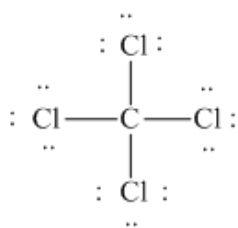
Carbon forms two sigma bonds and two pi bonds, the carbon involves in sp hybridization, so the bond angle will be 180° ; now the structure is as follows:



hydrogen cyanide

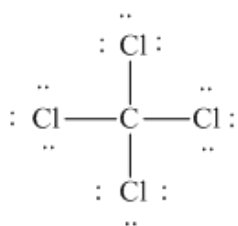
step 8 of 9

3. The molecule is CCl_4 , the structure can be drawn as follows:



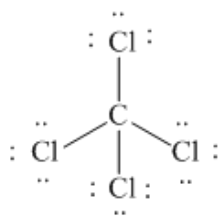
tetrachloromethane

Carbon forms four sigma bonds, there are no pi bonds, so carbon involves in sp^3 hybridisation, so the bond angle will be 90° , now the structure is as follows:



carbon tetra chloride

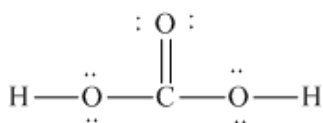
or



carbon tetra chloride

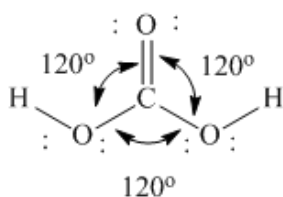
step 9 of 9

4. The molecule is H_2CO_3 , the structure can be drawn as follows:



carbonic acid

Carbon forms three sigma bonds and one pi bond, so the carbon involves in sp^2 hybridization, its bond angle is 120° , so the structure can be drawn as follows:



carbonic acid

Problem 33P

Predict the approximate bond angles in

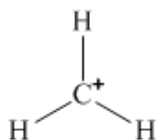
a. the methyl cation. b. the methyl radical. c. the methyl anion.

Step-by-step solution

step 1 of 3

a)

The structure of the methyl cation is as follows:



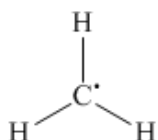
The hybridization of the methyl cation is sp^2 because it is formed by combination of one s-orbital and two p-orbitals. The three sp^2 orbitals are separated by 120° .

Hence, the bond angle in the methyl cation is 120°

step 2 of 3

b)

The structure of the methyl radical is as follows:



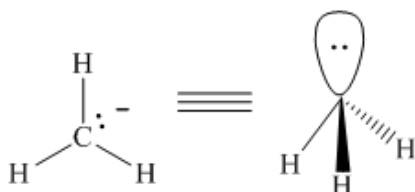
The hybridization of the methyl radical is sp^2 because it is formed by combination of one s-orbital and two p-orbitals. The three sp^2 orbitals are separated by 120° .

Hence, the bond angle in the methyl radical is 120°

step 3 of 3

c)

The structure of the methyl carbanion is as follows:



The hybridization of the methyl carbanion is sp^3 . It has four sp^3 orbitals. Among these four orbitals three sp^3 orbitals are overlap with 1s orbital of the hydrogen and fourth orbital have lone pair of electrons. These four sp^3 orbitals are separated by 109.5°

Hence, the bond angle in the methyl carbanion is 109.5° .

Problem 34P

According to the potential map for the ammonium ion, which atom has the greatest electron density?

Step-by-step solution

step 1 of 1

The colours in the potential map indicate the degree to which a molecule or atom in a molecule attracts charged particles. Red, signifying the most negative electrostatic potential, is used for regions that attract positively charged molecules most strongly. Blue is used for areas with the most positive electrostatic potential, regions that attract negatively charged molecules most strongly. Other colour indicate intermediate level of attraction

attracts positive charge \Rightarrow RedorangeyellowgreenBlue \Leftarrow attracts negative charge

So, according to the potential map of ammonium ion, nitrogen has the greatest electron density.

Step-by-step solution

step 1 of 3

1. The molecule is methane CH_4 , in methane molecule all the carbon and hydrogen bonds are placed at equal distance, there are no lone pair of electrons in the molecule, so the carbon involves in sp^3 hybridization, the structure of the molecule will be tetrahedral, the shape of the electrostatic potential map is also in tetrahedral, hence the methane molecule is a nonpolar molecule.

step 2 of 3

2. The molecule is ammonia NH_3 , in ammonia molecule all the nitrogen and hydrogen bonds are placed at equal distance, there are no lone pair of electrons in the molecule, so the nitrogen involves in sp^3 hybridization, the structure of the molecule should be trigonal planar with 120° , due to the presence of lone pair of electrons on nitrogen, the lone pair of electrons repulses the sigma bonds, so the bond angle in the molecule became 107.3° , more electric charge exist at lone pair of electrons, so ammonia molecule is slightly polar.

step 3 of 3

3. The molecule is water H_2O , there are two lone pair of electrons on oxygen atom, so the repulses the oxygen hydrogen bonds, hence the bond angle in water molecule is 104.5° , most electric charge located at lone pair of electrons, hence water molecule is a polar molecule.

Of the given three molecules water H_2O is most polar molecule, and methane CH_4 , is the least polar molecule.

Step-by-step solution

step 1 of 2

a)

The extent of overlapping of the hybrid orbitals decides the bond angle and bond strength, chlorine is smaller atom when compared to bromine, the overlapping of hybrid orbitals is more in chlorine, than in bromine, so the bond length in chlorine will be smaller than bond length in bromine

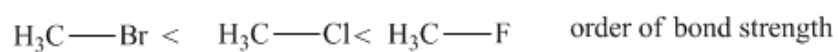
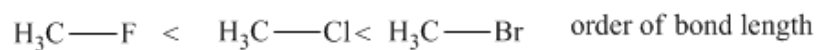
Length of the bond is inversely proportional to the strength of the bond, so the bond strength is more for chlorine molecule when compared to bromine molecule

step 2 of 2

b)

The carbon-halogen bond length increases with increase in size of the halogen, the carbon-halogen bond strength decreases with increase in size of the halogen

The order of the size of halogens is $F < Cl < Br$, so the order of bond lengths and strengths can be drawn as follows



Problem 38P

a. Which bond would be longer?

b. Which bond would be stronger?

1. C-Cl or C-I 2. C-C or C-Cl 3. H-Cl or H-F

Step-by-step solution

step 1 of 3

1. **C-Cl or C-I**:

Both Cl and I are halogens. The I atom has a much larger size than the Cl atom. When the atoms bonding are larger in size, the bond between them is longer. A longer bond is also a weaker bond. Hence the **C-I** bond is longer. As a result, the **C-Cl** bond is stronger.

step 2 of 3

2. **C-C or C-Cl**:

The atom Cl is larger than the atom C. Hence the **C-Cl** bond is longer. So the **C-C** bond is stronger.

step 3 of 3

3. **H-Cl or H-H**:

The atom Cl is much larger than the atom H. Hence the **H-Cl** bond is longer. So the **H-H** bond is stronger.

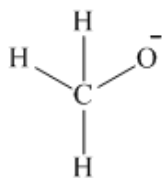
Step-by-step solution

step 1 of 5

a)

Carbon has four valance electrons, four valance orbitals (three p-orbitals and one s-orbital) each containing one electron will overlap to produce four sp^3 orbitals,

Three sp^3 orbitals overlaps with s-orbitals hydrogen atoms to form sigma bonds, another sp^3 orbital overlaps with p-orbital of oxide ion to form sigma bond, so the carbon involves in sp^3 hybridisation and forms four sigma bonds the shape of the molecule is tetrahedral, and bond angle is 109°



methanolate

step 2 of 5

b)

Carbon has four valence electrons, two valence orbitals (one p-orbital and one s-orbital) each containing one electron will overlap to produce two sp orbitals,

two sp orbitals overlaps with p-orbital oxygen atoms to form sigma bonds, remaining two p-orbitals of carbon involves in sidewise overlapping with p-orbital of each oxygen atom to form two pi bonds, so the carbon involves in sp hybridisation and forms two sigma bonds and two pi bonds, the shape of the molecule is linear, and bond angle is 180°

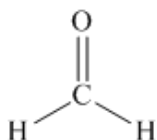


carbon dioxide

step 3 of 5

c)

carbon has four valence electrons, four valence orbitals (two p-orbitals and one s-orbital) each containing one electron will overlap to produce three sp^2 orbitals, two sp^2 orbitals overlaps with s-orbitals hydrogen atoms to form sigma bonds, another sp^2 orbital overlaps with p-orbital of oxygen to form sigma bond, remaining p-orbital of carbon involves in sidewise overlapping with p-orbital of oxygen atom to form a pi bond so the carbon involves in sp^2 hybridisation and forms three sigma bonds and one pi bond, the shape of the molecule is trigonal planar with bond angle 120°



formaldehyde

step 4 of 5

d)

in each nitrogen atom one s-orbital and one p-orbital hybridise to form two sp orbitals remaining two p-orbitals of each nitrogen involve in sidewise overlapping to form two pi-bonds, in each of the nitrogen atoms one of the sp orbital is filled with lone pair of electrons, so each nitrogen contains one sigma bond and two pi bonds, the shape of the molecule is linear, bond angle is 180°



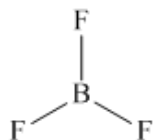
dinitrogen

step 5 of 5

e)

in boron atom one s-orbital and two p-orbitals involves in hybridization to form three sp^2 hybrid orbital, each hybrid orbital of boron overlaps with p-orbital of fluorine to form three sigma bonds, so the boron

trifluoride is in trigonal planar structure with bond angle 120°



boron trifluoride

Problem 41P

Which of the bonds in a carbon-oxygen double bond has more effective orbital-orbital overlap, the sigma bond or the pi bond?

Step-by-step solution

step 1 of 1



Problem 42P

Would you expect a C-C sigma bond formed by sp^2-sp^2 overlap to be stronger or weaker than a C-C sigma bond formed by sp^3-sp^3 overlap?

Step-by-step solution

step 1 of 1

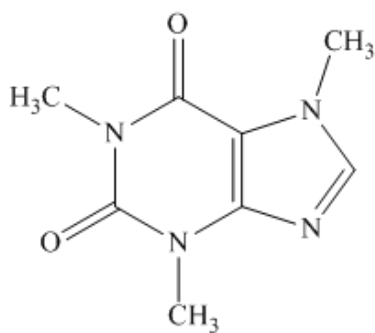
C-C σ bond formed by sp^2-sp^2 overlap is stronger than that formed by sp^3-sp^3 overlap. It has been estimated that C-C σ bond strength for sp^2-sp^2 overlap is around 112 kcal / mol where as same for sp^3-sp^3 overlap is almost 90 kcal / mol.

Reason: - More the s-character of the orbital's closer it will be to the nucleus. This makes bond length shorter and stronger. In sp^2 orbital, the s-character is 33.3% but in sp^3 orbital, it's only 25%. That's why sp^2-sp^2 overlap makes the bonds shorter and stronger compared to that for sp^3-sp^3 overlap.

Step-by-step solution

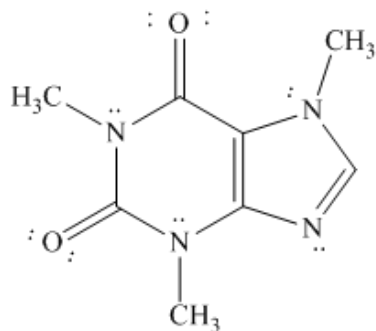
step 1 of 1

The structure of the caffeine molecule is



caffeine

Each nitrogen atom forms three bonds, and it contains one lone pair of electrons on it, each oxygen atom forms two bonds and it contains two lone pairs of electrons. Add one lone pair of electron on each nitrogen atom, two lone pairs of electrons on each oxygen atom, now the perfect structure can be drawn as follows



caffeine

Step-by-step solution

step 1 of 4

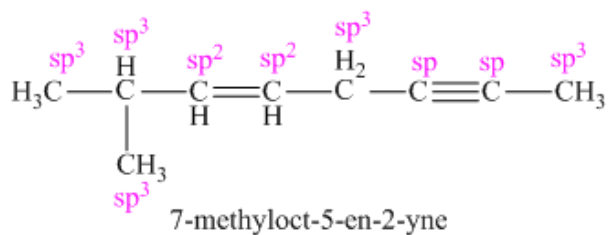
Hybridization: The process of mixing the atomic orbitals in an atom (usually the central atom) to generate a set of new atomic orbital is called hybridization. We can decide the hybridization of the central atom based on the number of electron groups around the central atom.

Number of electron groups (b.p + l.p)	Hybridization	Geometry
2	sp	Linear
3	sp^2	Trigonal planar
4	sp^3	Tetrahedral
5	sp^3d	Trigonal bipyramidal
6	sp^3d^2	Octahedral
7	sp^3d^3	Pentagonal bipyramidal

step 2 of 4

a)

The hybridisation of each carbon is denoted in pink color.

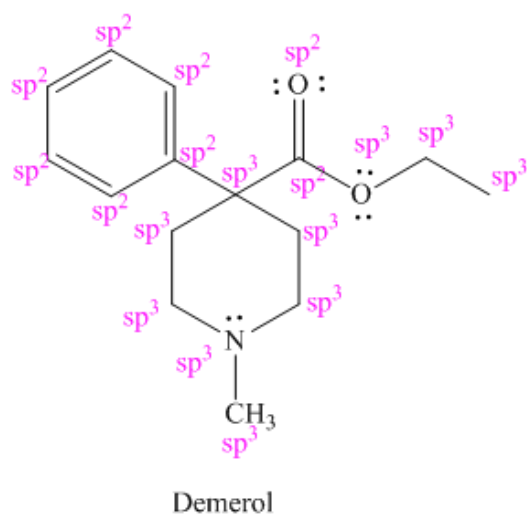


step 3 of 4

b)

The atom with four sigma bonds is sp^3 hybridised, the atom with three sigma bonds sp^2 hybridised, the atom with two sigma bonds is sp hybridised. Nitrogen has three bonding pairs and one lone pair. So, the hybridisation of nitrogen is sp^3 . One of the oxygen has one bond pair and two lone pair so its hybridisation is sp^2 . Another oxygen atom has two bond pairs and two lone pairs so its hybridisation is sp^3 .

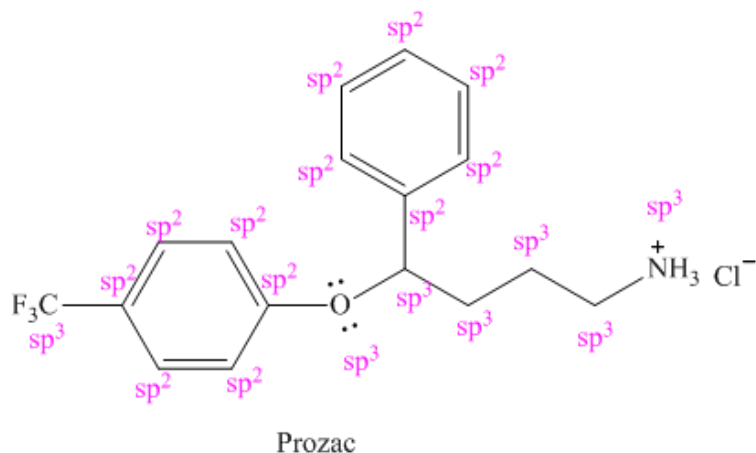
The hybridisation of each carbon, oxygen and nitrogen is denoted in pink color.



step 4 of 4

The atom with four sigma bonds is sp^3 hybridised, the atom with three sigma bonds sp^2 hybridised, the atom with two sigma bonds is sp hybridised. Nitrogen has four bond pairs. So, the hybridisation on nitrogen is sp^3 . Oxygen has two bond pair and two lone pair so its hybridisation is sp^3 . Another oxygen atom has two bond pairs and two lone pairs so its hybridisation is sp^3 .

The hybridisation of each carbon, oxygen and nitrogen is denoted in pink color.



Problem 45P

Predict the approximate bond angles for

- a. the C-N-C bond angle in $(\text{CH}_3)_2\text{N}^+\text{H}_2$. c. the H-C-N bond angle in $(\text{CH}_3)_2\text{NH}$.
- b. the C-N-H bond angle in $\text{CH}_3\text{CH}_2\text{NH}_2$. d. the H-C-O bond angle in CH_3OCH_3 .

Step-by-step solution

step 1 of 5

a.

In $(\text{CH}_3)_2\text{NH}_2^+$, N is the central atom and forms no π bonds. Therefore, it is sp^3 hybridized. The lone pair on the N atom is lost and a positive charge is formed that will reduce the lone pair repulsions and the bond angle is expected to remain the same. The bond angle between C-N-C is 109.5° .

step 2 of 5

b.

step 3 of 5

In $\text{CH}_3\text{CH}_2\text{NH}_2$, N is the central atom and forms only single bonds. So, it is said to be sp^3 hybridized. The uncharged N atom has a lone pair that tends to affect the bond angle slightly. So, the expected C-N-H bond angle is 107.5° .

step 4 of 5

c.

In $(\text{CH}_3)_2\text{NH}$, N is the central atom and forms only single bonds. So, it is said to be sp^3 hybridized. The uncharged N atom has a lone pair that tends to affect the bond angle slightly. So, the expected C-N-H bond angle is 107.5° .

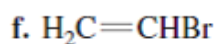
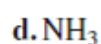
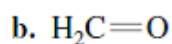
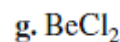
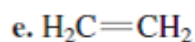
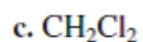
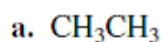
step 5 of 5

d.

In CH_3OCH_3 , O is the central atom and forms only single bonds. So, it is said to be sp^3 hybridized. The uncharged O atom has two lone pairs that tends to affect the bond angle slightly to a greater extent.. So, the expected C-O-H bond angle is 105° .

Problem 46P

Which of the following molecules would you expect to have a dipole moment of zero? To answer parts g and h, you may need to review the Problem Solving Strategy on page 42.



Step-by-step solution

step 1 of 10

The reasons for polarity of a molecule-

- (1) Electron negatively difference

step 2 of 10

- (2) Dissymmetry in its structure.

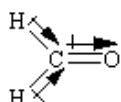
Direction of moment will always be towards the electro negative element.

step 3 of 10

- (A) CH_3CH_3 is a symmetric molecules, hence its dipole moment is zero.

step 4 of 10

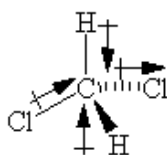
- (B)



The compound is not symmetric and all the moments are acting to a same direction resulting intense resultant dipole moment.

step 5 of 10

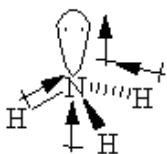
- (C)



The compound being a symmetric one, one moment nullifies other leading to resultant dipole moment zero.

step 6 of 10

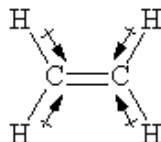
- (D)



NH_3 is dissymmetrical and vectroial addition of all the moments are acting to a same direction. So it has got high value of dipole moment.

step 7 of 10

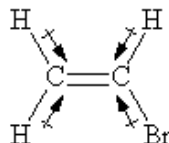
(E)



The compound being a symmetrical one, resultant dipole moment is zero.

step 8 of 10

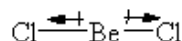
(F)



The compound is not symmetrical. So the resultant dipole moment is not zero at all.

step 9 of 10

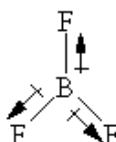
(G)



Since BeCl_2 is symmetrical, one moment nullifies the other one, making its dipole moment zero.

step 10 of 10

(H)



Dipole moment of BF_3 is zero following law of triangle.

Problem 47P

Account for the difference in the shape and color of the potential maps for ammonia and the ammonium ion in Section 1.11.

Step-by-step solution

step 1 of 2

The difference in the shape and color of the potential maps for $\ddot{\text{N}}\text{H}_3$ ion are due to the following reasons-

- (i) Both are having sp^3 hybridized N atom but bond angle in NH_4^+ is 109.5° where as that in $\ddot{\text{N}}\text{H}_3$ is 107.3° . This is due to lone pair-bond pair repulsion in $\ddot{\text{N}}\text{H}_3$.

step 2 of 2

- (ii) Electron density is highest on N atom in $\ddot{N}H_3$ but it is least on N atom in $\overset{+}{N}H_4$ ion. Because N reduces electron density by forming co-ordinate bond with H^+ ion in $\overset{+}{N}H_4$.

Problem 48P

If the dipole moment of CH_3F is 1.847 D and the dipole moment of CD_3F is 1.858 D, which is more electronegative, hydrogen or deuterium?

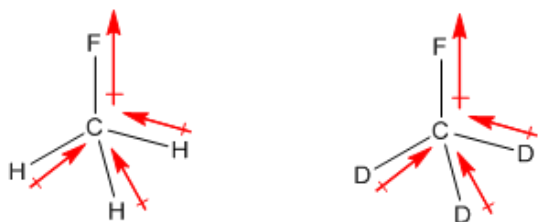
Step-by-step solution

step 1 of 3

Dipole moment is the product of magnitude of charge and the distance of separation between the charges. The polar covalent bond contains a dipole.

step 2 of 3

The dipole moment representations of CH_3F and CD_3F are as follows:



Here, the head of the arrow represents the negative end of the bond and the tail of the arrow represents the positive end of the bond. The negative end is more electronegative than the positive end.

step 3 of 3

From the provided data, the dipole moment of CH_3F is 1.847 D and for CD_3F is 1.858 D. So, the dipole moment of CD_3F is greater than CH_3F . Thus, CD_3F is more polar covalent than CH_3F , which represents that the electro negativity difference between C-D is greater than C-H.

Therefore, the hydrogen is more electronegative than deuterium.

Problem 49P

Step-by-step solution

step 1 of 5

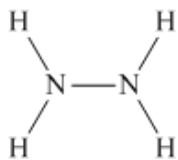
a)

The molecule is N_2H_4

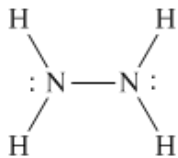
Arrange the atoms in the molecule with single bonds



The bonds for nitrogen and hydrogen atoms are not satisfied make sure that each nitrogen atom bonded with two hydrogen atoms



Arrange the loan pair of electrons on the nitrogen atoms, so the structure can be drawn as follows



hydrazine

step 2 of 5

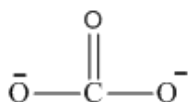
b)

The molecule is CO_3^{2-}

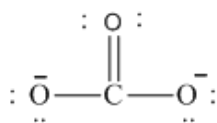
Arrange the atoms in the molecule with single bonds



The bonds for carbon are not satisfied; make double bond between carbon and middle oxygen



Arrange the loan air of electrons on the oxygen atom, so the structure can be drawn as follows



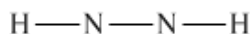
carbonate

step 3 of 5

c)

The molecule is N_2H_2

Arrange the atoms in the molecule with single bonds



The bonds for nitrogen atoms is not satisfied



Arrange the loan pair of electrons on the nitrogen atoms, so the structure can be drawn as follows



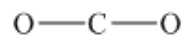
diazene

step 4 of 5

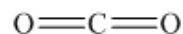
d)

The molecule is CO_2

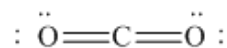
Arrange the atoms in the molecule with single bonds



The bonds for carbon are not satisfied; make double bond between carbon and both oxygen atoms



Arrange the lone pair of electrons on the oxygen atom, so the structure can be drawn as follows



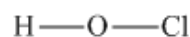
carbon dioxide

step 5 of 5

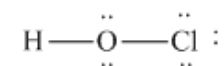
e)

The molecule is **HOCl**

Arrange the atoms in the molecule with single bonds



Arrange the lone pair of electrons on the oxygen atom and chlorine atom, so the structure can be drawn as follows

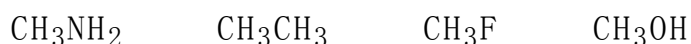


hypochlorous acid

Problem 50P

a. Which of the following has a nonpolar covalent bond?

b. Which of the following has a bond closest to the ionic end of the bond spectrum?



Step-by-step solution

step 1 of 2

(a)

The bond which has the electronegativity difference is zero is the responsible for non-polar nature.

The electronegativity difference of **C-C** is zero in the compound **$\text{CH}_3\text{-CH}_3$** . Thus, it is non-covalent bond.

step 2 of 2

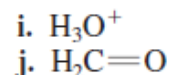
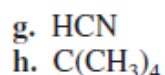
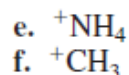
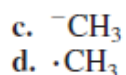
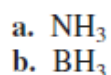
(b)

The bond which is more electronegativity difference, it is the responsible for more ionic nature. In the given compounds the more electronegativity difference is for **C-F** bond so **CH_3F** is more ionic.

Problem 51P

What is the hybridization of all the atoms (other than hydrogen) in

each of the following species? What are the bond angles around the central atom?



Step-by-step solution

step 1 of 11

Hybridization: The process of mixing the atomic orbitals in an atom (usually the central atom) to generate a set of new atomic orbital is called hybridization.

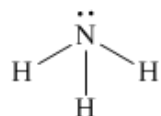
We can decide the hybridization of the central atom based on the number of electron groups around the central atom.

Number of electron groups (b.p + l.p)	Hybridization	Geometry
2	sp	Linear
3	sp^2	Trigonal planar
4	sp^3	Tetrahedral
5	sp^3d	Trigonal bipyramidal
6	sp^3d^2	Octahedral
7	sp^3d^3	Pentagonal bipyramidal

step 2 of 11

(a)

The Lewis structure of NH_3 molecule is as follows:



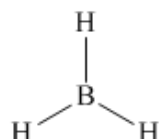
Here, the central nitrogen atom is attached to the four electron groups (3 bonded pairs + 1 lone pair). Hence, the hybridization of nitrogen atom in NH_3 is sp^3 .

Due to bond pair and lone pair repulsions, the regular bond angle of $109^\circ 28'$ reduces to 107.3° .

step 3 of 11

(b)

The Lewis structure of BH_3 molecule is as follows:

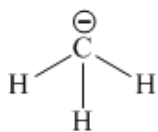


Here, the central boron atom is attached to the three electron groups (3 bonded pairs). Hence, the hybridization of boron atom in BH_3 is sp^2 . The bond angle is 120° .

step 4 of 11

(c)

The Lewis structure of CH_3^- molecule is as follows:

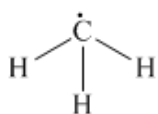


Here, the central carbon atom is attached to the four electron groups (3 bonded pairs + 1 lone pair). Hence, the hybridization of carbon atom in CH_3^- is sp^3 . Due to bond pair and negative charge repulsions, the regular bond angle of $109^\circ 28'$ reduces to 107.3° .

step 5 of 11

(d)

The Lewis structure of CH_3^\bullet molecule is as follows:

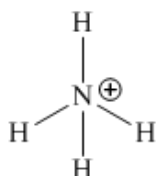


Here, the central carbon atom is attached to the three electron groups (3 bonded pairs). Hence, the hybridization of carbon atom in CH_3^\bullet is sp^2 . The bond angle is 120° .

step 6 of 11

(e)

The Lewis structure of NH_4^+ molecule is as follows:

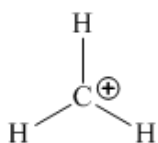


Here, the central nitrogen atom is attached to the four electron groups (4 bonded pairs). Hence, the hybridization of nitrogen atom in NH_4^+ is sp^3 . The bond angle is $109^\circ 28'$.

step 7 of 11

(f)

The Lewis structure of CH_3^+ molecule is as follows:

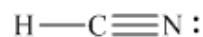


Here, the central carbon atom is attached to the three electron groups (3 bonded pairs). Hence, the hybridization of carbon atom in CH_3^+ is sp^2 . The bond angle is 120° .

step 8 of 11

(g)

The Lewis structure of HCN is as follows:



Here, the central carbon atom is attached to the two electron groups (2 bonded pairs). Hence, the hybridization of carbon atom in HCN is sp . The bond angle is 120° .

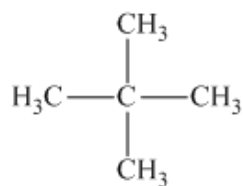
The nitrogen atom is attached to the two electron groups (1 bond pair + 1 lone pair). Hence, the hybridization of nitrogen atom in HCN is sp .

Due to bond pair and lone pair repulsions, the regular bond angle of 180° reduces to $<180^\circ$.

step 9 of 11

(h)

The Lewis structure of $\text{C}(\text{CH}_3)_4$ is as follows:

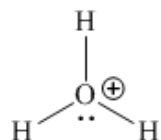


Here, the central carbon atom is attached to the four electron groups (4 bonded pairs). Hence, the hybridization of carbon atom in $\text{C}(\text{CH}_3)_4$ is sp^3 . The bond angle is $109^\circ 28'$.

step 10 of 11

(i)

The Lewis structure of H_3O^+ molecule is as follows:



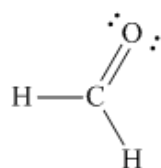
Here, the central oxygen atom is attached to the four electron groups (3 bonded pairs + one lone pair). Hence, the hybridization of oxygen atom in H_3O^+ is sp^3 .

Due to bond pair and lone pair repulsions, the regular bond angle of 109° reduces to 105° .

step 11 of 11

(j)

The Lewis structure of H_2CO is as follows:



Here, the central carbon atom is attached to the three electron groups (3 bonded pairs). Hence, the hybridization of carbon atom in H_2CO is sp^2 . The bond angle is 120° .

The oxygen atom is attached to the three electron groups (3 bonded pairs). Hence, the hybridization of oxygen

atom in H_2CO is $\boxed{sp^2}$. Due to bond pair and lone pair repulsions, the regular bond angle of 120° reduces to $\boxed{<120^\circ}$.

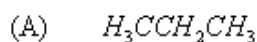
Problem 52P

Draw the condensed structure of a compound that contains only carbon and hydrogen atoms and that has

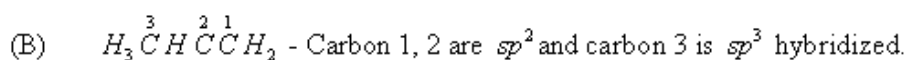
- three sp^3 carbons.
- one sp^3 carbon and two sp^2 carbons.
- two sp^3 carbons and two sp carbons.

Step-by-step solution

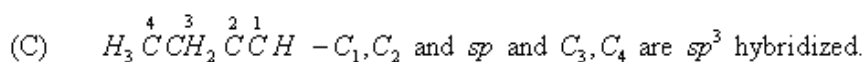
step 1 of 3



step 2 of 3



step 3 of 3



Problem 53P

Predict the approximate bond angles:

- the C-N-C bond angle in $(\text{CH}_3)_2^+\text{NH}_2$
- the C-N-H bond angle in $(\text{CH}_3)_2\text{NH}$
- the C-O-H bond angle in CH_3OH
- the C-N-C bond angle in $(\text{CH}_3)_2\text{NH}$

Step-by-step solution

step 1 of 5

a.

In $(\text{CH}_3)_2\text{NH}_2^+$, N is the central atom and forms no π bonds. Therefore, it is sp^3 hybridized. The lone pair on the N atom is lost and a positive charge is formed that will reduce the lone pair repulsions and the bond angle is expected to remain the same. The bond angle between $\text{C}-\text{N}-\text{C}$ is 109.5° .

step 2 of 5

b.

step 3 of 5

In CH_3OH , C is the central atom and forms only single bonds. So, it is said to be sp^3 hybridized. The uncharged O atom has two lone pairs that tend to affect the bond angle slightly to a greater extent. So, the expected $\text{C}-\text{O}-\text{H}$ bond angle is 105° .

step 4 of 5

c.

In $(\text{CH}_3)_2\text{NH}$, N is the central atom and forms only single bonds. So, it is said to be sp^3 hybridized. The uncharged N atom has a lone pair that tends to affect the bond angle slightly. So, the expected $\text{C}-\text{N}-\text{H}$ bond angle is 107.5° .

step 5 of 5

d.

In $(\text{CH}_3)_2\text{NH}$, N is the central atom and forms only single bonds. So, it is said to be sp^3 hybridized. The uncharged N atom has a lone pair that tends to affect the bond angle slightly. So, the expected $\text{C}-\text{N}-\text{C}$ bond angle is 104.5° due to the methyl group repulsions.

Problem 54P

Draw the ground-state electronic configuration for the following:

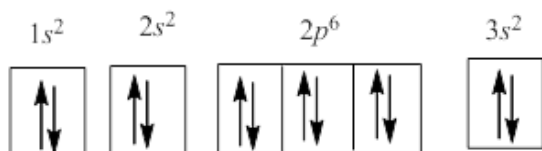
a. Mg b. Ca^{2+} c. Ar d. Mg^{2+}

Step-by-step solution

step 1 of 4

(a)

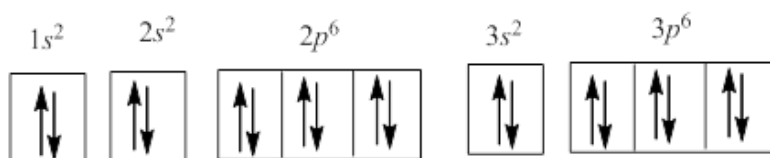
The atomic number of magnesium (Mg) is 12. So, it has 12 electrons. The ground state electronic configuration of Mg is as follows:



step 2 of 4

(b)

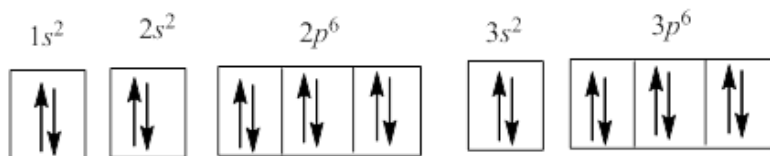
The atomic number of calcium (Ca) is 20. So, it has 20 electrons. Calcium (Ca) loses two electrons to form its cation Ca^{2+} . This cation has only 18 electrons. The ground state electronic configuration of Ca^{2+} is as follows:



step 3 of 4

(c)

The atomic number of argon (**Ar**) is 18. So, it has 18 electrons. The ground state electronic configuration of **Ar** is as follows:



step 4 of 4

(d)

The atomic number of magnesium (**Mg**) is 12. So, it has 12 electrons. Magnesium (**Mg**) loses two electrons to form its cation **Mg²⁺**. This cation has only 10 electrons. The ground state electronic configuration of **Mg²⁺** is as follows:

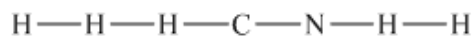
Step-by-step solution

step 1 of 5

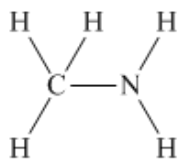
a)

The molecule is **CH₃NH₂**

Arrange the atoms in the molecule with single bonds

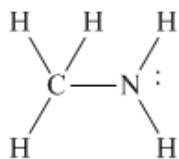


The bonds for carbon, nitrogen and hydrogen atoms are not satisfied make sure that carbon bonded to three hydrogen atoms and nitrogen atom bonded with two hydrogen atoms



methanamine

Arrange the lone pair of electrons on the nitrogen atom, so the structure can be drawn as follows



methanamine

step 2 of 5

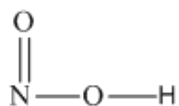
b)

The molecule is **HNO₂**

Arrange the atoms in the molecule with single bonds

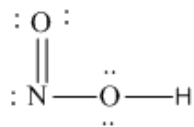


The bonds for nitrogen and oxygen are not satisfied; attach the hydrogen to one of the oxygen atom; remaining oxygen will make the double bond with nitrogen



Arrange the lone pair of electrons on the oxygen atoms and nitrogen atom, so the structure can be drawn as follows

nitrous acid

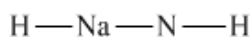


step 3 of 5

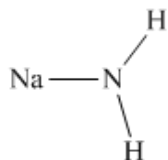
c)

The molecule is NaNH_2

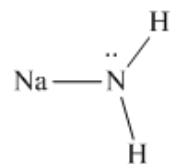
Arrange the atoms in the molecule with single bonds



The bonds for nitrogen and sodium atoms are not satisfied, attach two hydrogen atoms to nitrogen



Arrange the lone pair of electrons on the nitrogen atoms, so the structure can be drawn as follows



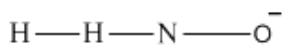
sodium amide

step 4 of 5

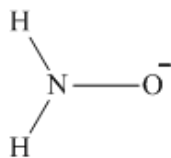
d)

The molecule is NH_2O^-

Arrange the atoms in the molecule with single bonds

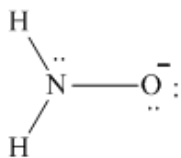


The bonds for hydrogen and nitrogen are not satisfied; attach two hydrogen atoms to nitrogen atom



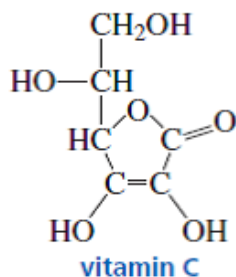
step 5 of 5

Arrange the lone pair of electrons on the oxygen and nitrogen atoms, so the structure can be drawn as follows



Problem 56P

What is the hybridization of each of the carbon and oxygen atoms in vitamin C?



Step-by-step solution

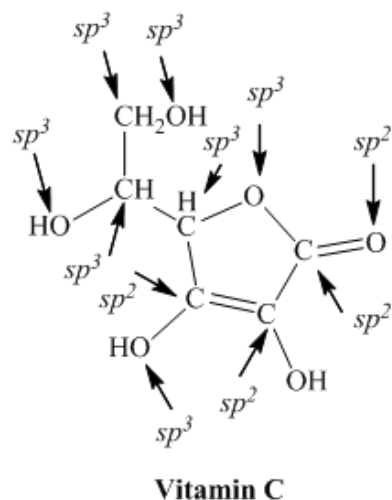
step 1 of 2

The hybridization of the atoms in any molecules indicated by the following rules:

- (1) If the atom is not surrounding by π bonds, it will be sp^3 hybridized.
- (2) If the atom is surrounded by one π bond, it will be sp^2 hybridized.
- (3) If the atom is surrounded by two π bonds, it will be sp hybridized.

step 2 of 2

The hybridization of each of the carbon and oxygen atoms in vitamin C as shown below:



Step-by-step solution

step 1 of 4

In a chemical bond the electron pair is shared by two atoms, the bond pair electrons has to lie at the midpoint between the two bonded atoms, but the more electronegative atom will attracts the bond pair of electrons towards itself, so bond pair of electrons lie close to the more electronegative atom participating

in bonding this is nothing but polarity

a)

Given chemical bonds are C—O , C—F C—N

in these three chemical bonds one of the bonding atom carbon is same, so the polarity of the bond depends upon the remaining atom participating in the bond, more electronegative atom drags the bond pair of electrons close towards itself, this results in more polar bond, so polarity depends upon the electro negativity of the atom, as the electro negativity of the atom increases the polarity of the bond also increases

The electro negativity order of the oxygen, fluorine, and nitrogen is

$\text{N} < \text{O} < \text{F}$ electro negativity order

Let us give ranking based on polarity order of the bonds:

1. C—F
2. C—O
3. C—N

So more polar bond is C—F , and least polar bond is C—N

step 2 of 4

b)

Given chemical bonds are C—Cl , C—I C—Br

in these three chemical bonds one of the bonding atom carbon is same, so the polarity of the bond depends upon the remaining atom participating in the bond, more electronegative atom drags the bond pair of electrons close towards itself, this results in more polar bond, so polarity depends upon the electro negativity of the atom, as the electro negativity of the atom increases the polarity of the bond also increases

The electro negativity order of the chlorine, iodine and bromine is

$\text{I} < \text{Br} < \text{Cl}$ electro negativity order

Let us give ranking based on polarity order of the bonds:

1. C—Cl
2. C—Br
3. C—I

So more polar bond is C—Cl , and least polar bond is C—I

step 3 of 4

c)

Given chemical bonds are H—O , H—N H—C

in these three chemical bonds one of the bonding atom hydrogen is same, so the polarity of the bond depends upon the remaining atom participating in the bond, more electronegative atom drags the bond pair of electrons close towards itself, this results in more polar bond, so polarity depends upon the electro negativity of the atom, as the electro negativity of the atom increases the polarity of the bond also increases

The electro negativity order of the oxygen, nitrogen and carbon is

$\text{C} < \text{N} < \text{O}$ electro negativity order

Let us give ranking based on polarity order of the bonds:

1. H—O
2. H—N
3. H—C

So more polar bond is $\text{H}-\text{O}$, and least polar bond is $\text{H}-\text{C}$

step 4 of 4

d)

Given chemical bonds are $\text{C}-\text{H}$, $\text{C}-\text{C}$, $\text{C}-\text{N}$

in these three chemical bonds one of the bonding atom carbon is same, so the polarity of the bond depends upon the remaining atom participating in the bond, more electronegative atom drags the bond pair of electrons close towards itself, this results in more polar bond, so polarity depends upon the electro negativity of the atom, as the electro negativity of the atom increases the polarity of the bond also increases, here in these bonds carbon-carbon bond is non polar because, the atoms participating in the bond are same so the bond air of electrons lie at the centre of the bond, hence no polarity

The electro negativity order of the hydrogen, carbon and nitrogen is

$\text{H} < \text{C} < \text{N}$ electro negativity order

Let us give ranking based on polarity order of the bonds:

1. $\text{C}-\text{N}$
2. $\text{C}-\text{H}$
3. $\text{C}-\text{C}$

So more polar bond is $\text{C}-\text{N}$, and least polar bond is $\text{C}-\text{C}$

Problem 58P

Draw the Lewis structure for each of the following compounds:

a. CH_3CHO b. CH_3OCH_3 c. CH_3COOH

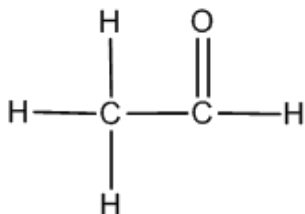
Step-by-step solution

step 1 of 3

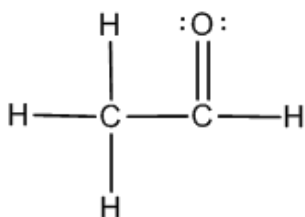
a)

The number of valence electrons in C is 4, H is 1, and O is 6. So, CH_3CHO have 18 valence electrons.

Carbon valency is four so it should have four bonds, similarly H has one bond, and O has two bonds. So, draw the possible arrangement of atoms in CH_3CHO .



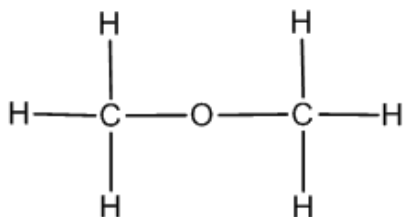
The above structure contains 14 electrons; add the remaining 4 electrons as lone pairs on O atom. This gives the Lewis structure of CH_3CHO .



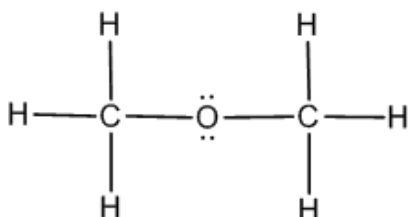
step 2 of 3

b)

The number of valence electrons in C is 4, H is 1, and O is 6. So, CH_3OCH_3 have 20 valence electrons. Carbon valency is four so it should have four bonds, similarly H has one bond, and O has two bonds. So, draw the possible arrangement of atoms in CH_3OCH_3 .



The above structure contains 16 electrons; add the remaining 4 electrons as lone pairs on O atom. This gives the Lewis structure of CH_3OCH_3 .

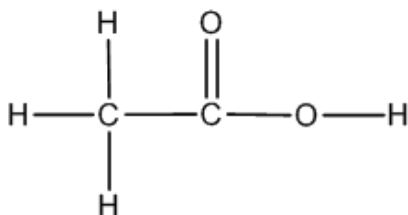


step 3 of 3

c)

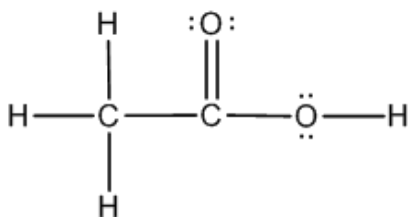
The number of valence electrons in C is 4, H is 1, and O is 6. So, CH_3COOH have 24 valence electrons. Carbon valency is four so it should have four bonds, similarly H has one bond, and O has two bonds.

So, draw the possible arrangement of atoms (avoid $\text{O}-\text{O}$ bonds) in CH_3COOH .



The above structure contains 16 electrons; add the remaining 8 electrons as two lone pairs on each O atom.

This gives the Lewis structure of CH_3COOH .



Step-by-step solution

step 1 of 4

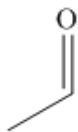
step 2 of 4

In skeletal structure we need to draw the molecule in vertex model, no need to represent the hydrogen atoms bonded to carbon atom, need not to represent the lone air of electrons present on the atoms, each vertex in

the structure represents the carbon atom in the molecule with satisfied number of hydrogen atoms

a)

Given compound is CH_3CHO , the given compound is an aldehyde, now skeletal structure can be drawn as follows,

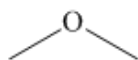


acetaldehyde

step 3 of 4

b)

Given compound is CH_3OCH_3 , presence of oxygen between two methyl groups indicates that the compound is an ether, now skeletal structure can be drawn as follows

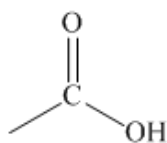


methoxymethane

step 4 of 4

c)

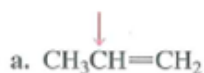
Given compound is CH_3COOH , presence of COOH group indicates that the compound is a carboxylic acid, now skeletal structure can be drawn as follows



acetic acid

Problem 60P

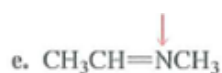
What is the hybridization of the indicated atom in each of the following compounds?





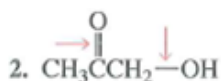
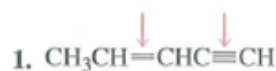






a. Which of the indicated bonds in each compound is shorter?

b. Indicate the hybridization of the C, O, and N atoms in each of the compounds.



Step-by-step solution

step 1 of 6

(A) sp^2 hybridization as the C atom forms a π bond.

step 2 of 6

(B) sp^2 hybridization as the O atom forms a π bond with carbon atom.

step 3 of 6

(C) sp^3 hybridization as the two sp^3 orbital's form covalent bond with carbon hydrogen and the other two sp^3 orbital contains two lone pair of electrons.

step 4 of 6

(D) N atom is sp hybridized as it forms two π bonds with carbon.

step 5 of 6

(E) N atom is sp^2 hybridized as it forms π bond with carbon.

step 6 of 6

(F) O atom is sp^3 hybridized as the two sp^3 orbital's combine with two carbon atoms and other two sp^3 orbital accommodate the two long pairs.

Step-by-step solution

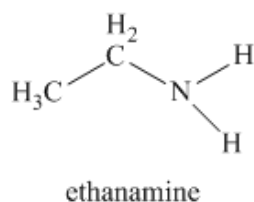
step 1 of 4

a)

The molecule is $\text{CH}_3\text{CH}_2\text{NH}_2$

let us find out the hybridisation of nitrogen atom in the molecule from that the shape of the molecule can be obtained, from the shape we can determine the bond angle, the nitrogen contains three valance orbitals with unpaired electrons (one s-orbital and two p-orbitals) now these three valance orbitals overlaps to produce

three sp^2 hybrid orbitals, these three sp^2 orbitals forms sigma bonds, one lone pair of electrons occupies the p-orbital of nitrogen, so the structure of the molecule at nitrogen centre is trigonal pyramidal, the expected bond angle is 109° , but the bulky ethyl group repulses hydrogen hence the $\text{C}-\text{N}-\text{H}$ bond angle will more than 109° , so bond angle lies between 109° to 120°

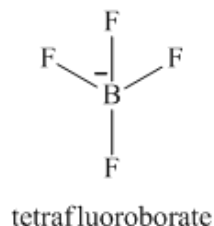


step 2 of 4

b)

The molecule is BF_4^-

let us find out the hybridisation of boron atom in the molecule from that the shape of the molecule can be obtained, from the shape we can determine the bond angle, after exiting an electron from the s-orbital to -orbital the boron contains three valance orbitals with unpaired electrons (one s-orbital and two p-orbitals), here boron receives an electron to form boron ion, now the boron ion contains four valance orbitals (one s-orbital and two p-orbitals), now these four valance orbitals overlaps to produce four sp^3 hybrid orbitals, these four sp^3 orbitals forms four sigma bonds with p-orbital of four fluorine atoms, so the hybridization on boron is sp^3 , its structure is tetrahedral and the $\text{F}-\text{B}-\text{F}$ bond angle is 109°

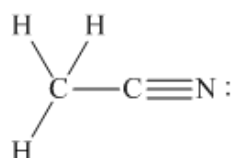


step 3 of 4

c)

The molecule is $\text{CH}_3\text{C}\equiv\text{N}$

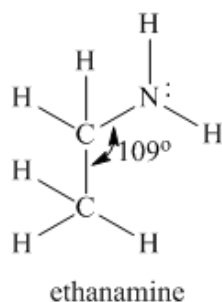
let us find out the hybridisation of carbon atom in the molecule, from that the shape of the molecule can be obtained, from the shape we can determine the bond angle, the carbon contains four valance orbitals with unpaired electrons, but one s-orbital and one p-orbital overlaps to produce two sp hybrid orbitals, these two sp hybrid orbitals forms sigma bonds, remaining two p-orbitals involves in sidewise overlapping with p-orbitals of nitrogen and results in two pi bonds, so the central carbon contains two sigma bonds and two pi bonds, it involves in sp hybridisation and its shape is linear, its bond angle is 180° , hence the $\text{C}-\text{C}-\text{N}$ bond angle will be 180°



d)

The molecule is $\text{CH}_3\text{CH}_2\text{NH}_2$

let us find out the hybridisation of carbon atom in the molecule, from that the shape of the molecule can be obtained, from the shape we can determine the bond angle, the carbon contains four valence orbitals with unpaired electrons, (one s-orbital and three p-orbitals) these four valence orbitals overlap to produce four sp^3 hybrid orbitals, these four sp^3 hybrid orbitals form sigma bonds, so the central carbon contains four sigma bonds, it involves in sp^3 hybridisation and its shape is tetrahedral, its bond angle is 109° , hence the $\text{C}-\text{C}-\text{N}$ bond angle will be 109°



Problem 62P

Show the direction of the dipole moment in each of the following bonds (use the electronegativities given in Table 1.3):

a. $\text{H}_3\text{C}-\text{Br}$ b. $\text{H}_3\text{C}-\text{Li}$ c. $\text{HO}-\text{NH}_2$ d. $\text{I}-\text{Br}$ e. $\text{H}_3\text{C}-\text{OH}$ f. $(\text{CH}_3)_2\text{N}-\text{H}$

Reference: Table 1.3:

Table 1.3 The Electronegativities of Selected Elements^a

1A	2A		3A	4A	5A	6A	7A
H 2.1							
Li 1.0	Be 1.5		B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2		Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0						Br 2.8
							I 2.5

increasing electronegativity →

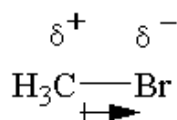
↑ increasing electronegativity

^aElectronegativity values are relative, not absolute. As a result, there are several scales of electronegativities. The electronegativities listed here are from the scale devised by Linus Pauling.

Step-by-step solution

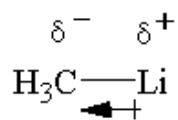
step 1 of 6

(a)



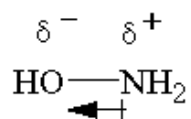
step 2 of 6

(b)



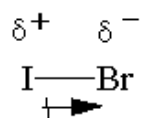
step 3 of 6

(c)



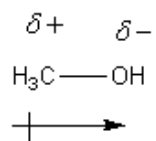
step 4 of 6

(d)



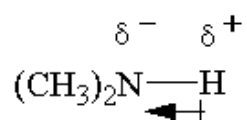
step 5 of 6

(e)



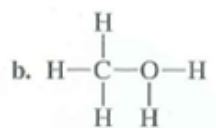
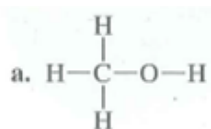
step 6 of 6

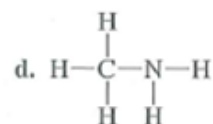
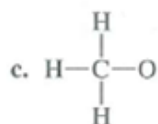
(f)



Problem 63P

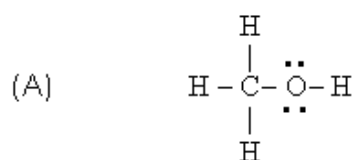
Draw the missing lone pairs and assign the missing formal charges.





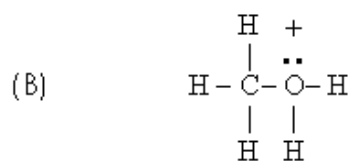
Step-by-step solution

step 1 of 4



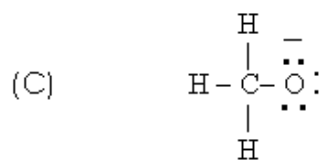
Formal charge = [No. of valence electrons - (no. of lone pair of electrons + $\frac{1}{2} \times$ no. of bonding electrons)]

step 2 of 4



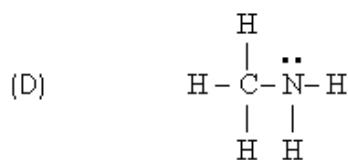
$$\begin{aligned} \text{Formal charge on oxygen} &= \left[6 - \left(2 + \frac{1}{2} \times 6 \right) \right] \\ &= [6 - 5] \\ &= +1 \end{aligned}$$

step 3 of 4



$$\begin{aligned} \text{Formal charge on oxygen} &= \left[6 - \left(6 + \frac{1}{2} \times 2 \right) \right] \\ &= [6 - 7] \\ &= -1 \end{aligned}$$

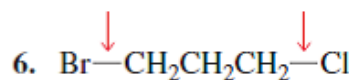
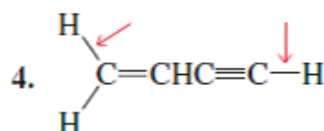
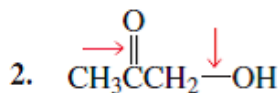
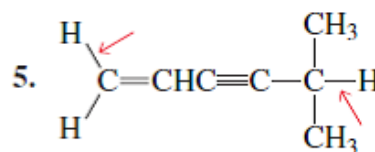
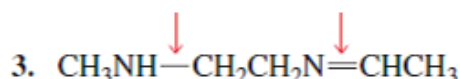
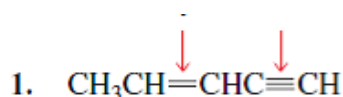
step 4 of 4



$$\begin{aligned}
 \text{Formal charge on nitrogen} &= \left[5 - \left(2 + \frac{1}{2} \times 6 \right) \right] \\
 &= [5 - 5] \\
 &= 0
 \end{aligned}$$

Problem 64P

- Which of the indicated bonds in each molecule is shorter?
- Indicate the hybridization of the C, O, and N atoms in each of the molecules.



Step-by-step solution

step 1 of 12

(a)

(1) In $\text{C}\equiv\text{C}$ bond length is shorter than $\text{C}=\text{C}$ because sp hybridization of carbon percentage of s character is maximum. The more s character in the orbital used by carbon in bond formation, the shorter is the bond length because s orbital is closer to the nucleus than a p-orbital.

step 2 of 12

(2) In $\text{C}=\text{O}$ bond length is shorter than $\text{C}-\text{O}$ because sp^2 hybridization of carbon and oxygen percentage of s character is maximum. The more s character in the orbital used by carbon and oxygen in bond formation, the shorter is the bond length because s orbital is closer to the nucleus than a p-orbital.

step 3 of 12

(3) In $\text{N}=\text{C}$ bond length is shorter than $\text{N}-\text{C}$ because sp^2 hybridization of carbon and nitrogen percentage of s character is maximum. The more s character in the orbital used by carbon and nitrogen in bond formation, the shorter is the bond length because s orbital is closer to the nucleus than a p-orbital.

step 4 of 12

(4) In $\text{H}-\text{C}\equiv$ bond length is shorter than $\text{H}-\text{C}=\text{}$ because sp hybridization of carbon percentage of s character is maximum. The more s character in the orbital used by carbon in bond formation, the shorter is the bond length because s orbital is closer to the nucleus than a p-orbital.

step 5 of 12

(5) In $\text{H}-\text{C}=\text{}$ bond length is shorter than $\text{H}-\text{C}-$ because sp hybridization of carbon percentage of s character is maximum. The more s character in the orbital used by carbon in bond formation, the shorter is the bond length because s orbital is closer to the nucleus than a p-orbital.

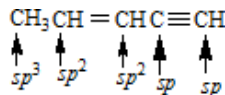
step 6 of 12

(6) In $\text{CH}_2\text{-Cl}$ bond length is shorter than $\text{CH}_2\text{-Br}$ because the average bond distance increases with increase in size of the atom.

step 7 of 12

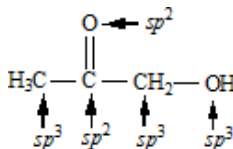
(b)

(1) Hybridization of C in the given molecule is as follows:



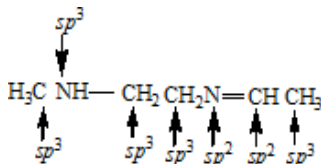
step 8 of 12

(2) Hybridization of C, O atoms in the given molecule is



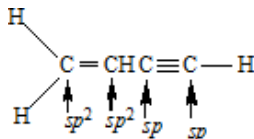
step 9 of 12

(3) Hybridization of C, N atoms in the given molecule is



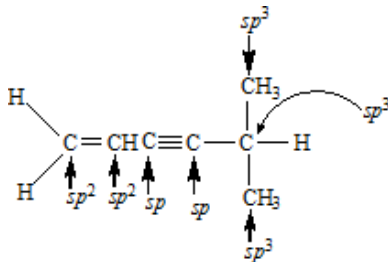
step 10 of 12

(4) Hybridization of C in the given molecule is



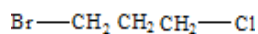
step 11 of 12

(5) Hybridization of C in the given molecule is as follows:



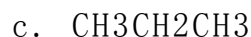
step 12 of 12

(6) Hybridization of C in the given molecule is



Problem 65P

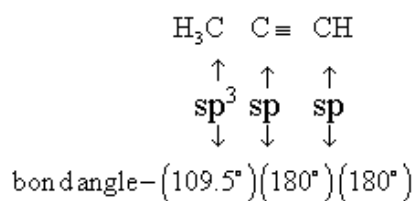
For each of the following compounds, indicate the hybridization of each carbon atom and give the approximate values of all the bond angles:



Step-by-step solution

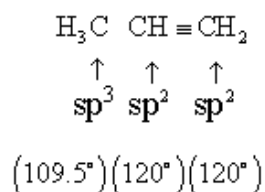
step 1 of 4

(A)



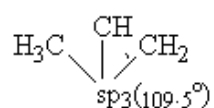
step 2 of 4

(B)



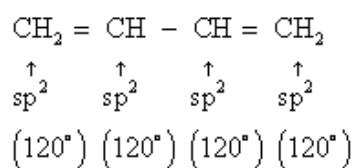
step 3 of 4

(C)



step 4 of 4

(D)



Problem 66P

Draw the Lewis structure for each of the following compounds:

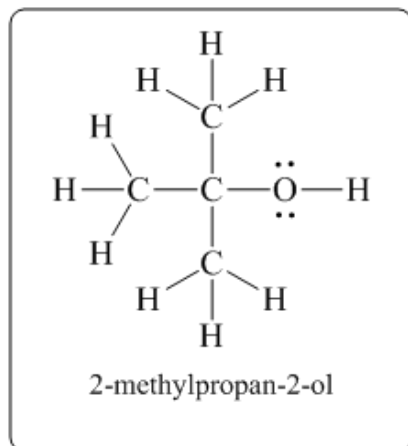
a. $(\text{CH}_3)_3\text{COH}$ b. $\text{CH}_3\text{CH}(\text{OH})\text{CH}_2\text{CN}$ c. $(\text{CH}_3)_2\text{CHCH}(\text{CH}_3)\text{CH}_2\text{C}(\text{CH}_3)_3$

Step-by-step solution

step 1 of 3

a)

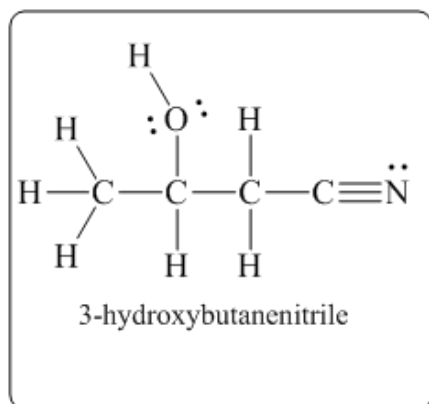
The Lewis structure of the compound $(\text{CH}_3)_3\text{COH}$ is shown below:



step 2 of 3

b)

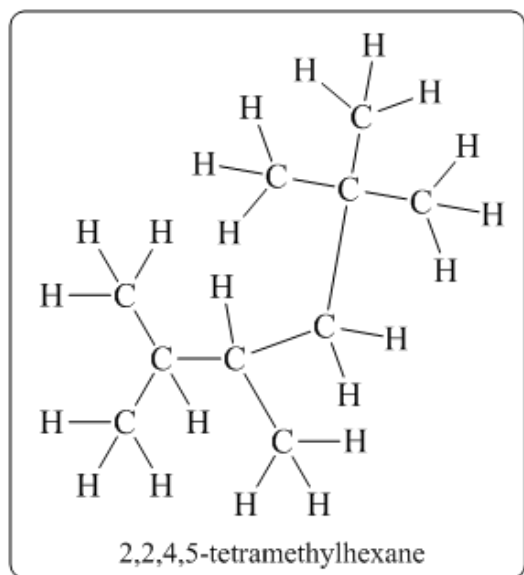
The Lewis structure of the compound $\text{CH}_3\text{CH}(\text{OH})\text{CH}_2\text{CN}$ is shown below:



step 3 of 3

c)

The Lewis structure of the compound $(\text{CH}_3)_2\text{CHCH}(\text{CH}_3)\text{CH}_2\text{C}(\text{CH}_3)_3$ is shown below:



Problem 67P

Step-by-step solution

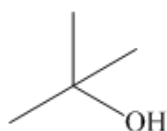
step 1 of 4

step 2 of 4

In skeletal structure we need to draw the molecule in vertex model, no need to represent the hydrogen atoms bonded to carbon atom, need not to represent the lone pair of electrons present on the atoms, each vertex in the structure represents the carbon atom in the molecule with satisfied number of hydrogen atoms

a)

Given compound is $(\text{CH}_3)_3\text{COH}$, the methyl CH_3 groups attached to same terminal carbon are denoted in parentheses, now skeletal structure can be drawn as follows,

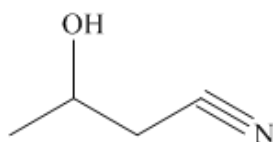


2-methylpropan-2-ol

step 3 of 4

b)

Given compound is $\text{CH}_3\text{CH}(\text{OH})\text{CH}_2\text{CN}$, the hydroxyl OH group attached to third carbon are denoted in parentheses, now skeletal structure can be drawn as follows,



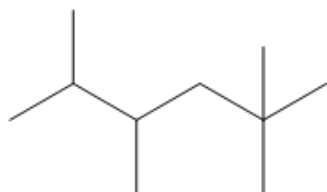
3-hydroxybutanenitrile

step 4 of 4

c)

Given compound is $(\text{CH}_3)_2\text{CHCH}(\text{CH}_3)\text{CH}_2\text{C}(\text{CH}_3)_3$, the methyl CH_3 groups attached to same terminal carbon

are denoted in parentheses, one side chine attachment of CH_3 group to fourth carbon is also denoted in parentheses, now skeletal structure can be drawn as follows,



2,2,4,5-tetramethylhexane

Problem 68P

Step-by-step solution

step 1 of 6

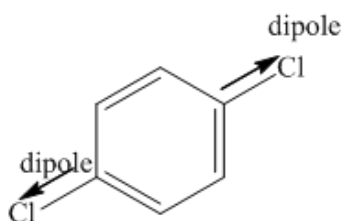
In a chemical bond the electron pair is shared by two atoms, the bond pair electrons has to lie at the midpoint between the two bonded atoms, but the more electronegative atom will attracts the bond pair of electrons towards itself, so bond pair of electrons lie close to the more electronegative atom participating in bonding this is nothing but polarity.

The dipole moment of a molecule depends on the magnitude and the direction of the individual dipoles.

step 2 of 6

1)

Given molecule is 1,4 dichlorobenzene, the chlorine is an electronegative element, its bonding with the benzene ring acts as dipole.



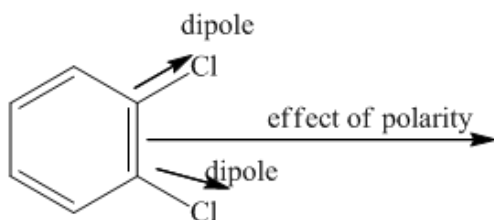
1,4-dichlorobenzene

But here the two dipoles resent in opposite direction, so the polarity is cancelled by each other, hence 1,4 dichlorobenzene is nonpolar molecule.

step 3 of 6

2)

Given molecule is 1,2 dichlorobenzene, the chlorine is an electronegative element, its bonding with the benzene ring acts as dipole.



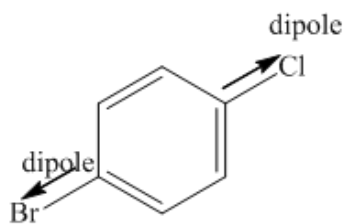
1,2-dichlorobenzene

Here the two dipoles present in nearly same direction, so the polarity is enhanced, hence 1,2 dichlorobenzene is highly polar molecule, the polarity lies in the effective direction.

step 4 of 6

3)

Given molecule is 1-bromo-4-chlorobenzene; both bromine and chlorine are electronegative elements, bonding of each atom with benzene ring acts as dipole.



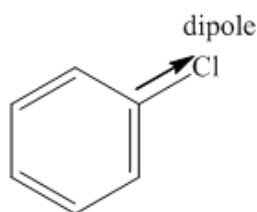
1-bromo-4-chlorobenzene

But here the two dipoles present in opposite direction, so the polarity is cancelled by each other, here chlorine is more electro negative than bromine, hence chlorine bond is more polar than bromine bond, hence the molecule is slightly polar towards chlorine bond.

step 5 of 6

4)

Given molecule is chlorobenzene, the chlorine is an electronegative element, its bonding with the benzene ring acts as dipole.

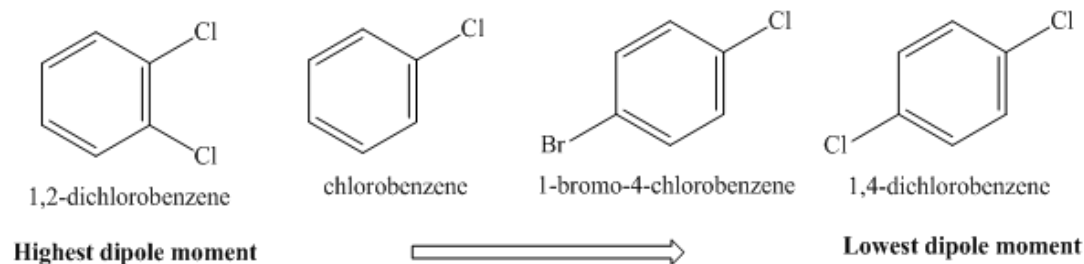


chlorobenzene

Chlorine bond is polar bond, hence chlorobenzene is polar molecule.

step 6 of 6

After analysing all four molecules we can say the order of molecules from highest dipole moment to lowest dipole moment is as follows:

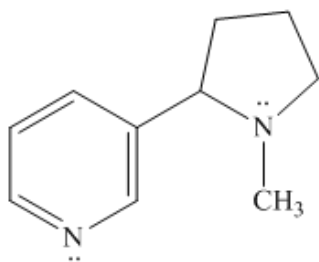


Problem 69P

Step-by-step solution

step 1 of 1

The structure of nicotine can be drawn as follows:



3-(1-methylpyrrolidin-2-yl)pyridine

or

nicotine

The nitrogen in the pyridine ring:

Nitrogen has four valence orbitals, they are one s-orbital and three p-orbitals, here one s-orbital and one p-orbital involves in hybridization to form two sp hybrid orbital, each hybrid orbital of nitrogen forms sigma bond with adjacent carbon atom, in the remaining two p-orbitals one of the p-orbital involves in sidewise overlap with p-orbital of adjacent carbon to form pi bond, remaining lone pair of electrons occupies in another p-orbital

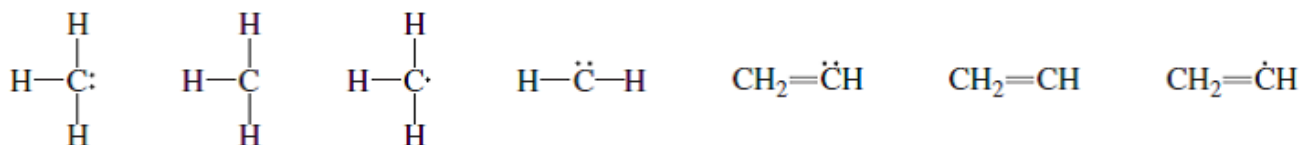
The nitrogen in the pyrrolidine ring:

Nitrogen has four valence orbitals, they are one s-orbital and three p-orbitals, here one s-orbital and two p-orbitals involves in hybridization to form three sp^2 hybrid orbital, each hybrid orbital of nitrogen forms sigma bond with adjacent carbon atom, the lone pair of electrons occupies in another p-orbital

So it is clear that the lone pair of electrons in nicotine exist in p-orbitals of the nitrogen atoms

Problem 70P

Indicate the formal charge on each carbon that has one. All lone pairs are shown.



Step-by-step solution

step 1 of 8

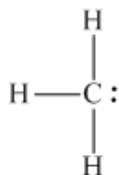
The formal charges can be calculated by using the below formula:

Formal charge = Number of valence electrons

$$-\left(\text{number of lone-pair electrons} + \frac{\text{number of bonding electrons}}{2}\right)$$

step 2 of 8

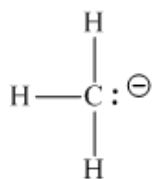
The given compound is as follows:



Calculate the formal charge for the above compound:

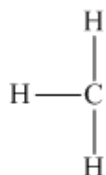
$$\text{Formal charge on C} = 4 - \left(2 + \left(\frac{6}{2} \right) \right) \\ = -1$$

So, the given compound with charge is as follows:



step 3 of 8

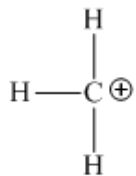
The given compound is as follows:



Calculate the formal charge for the above compound:

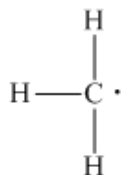
$$\text{Formal charge on C} = 4 - \left(0 + \left(\frac{6}{2} \right) \right) \\ = +1$$

So, the given compound with charge is as follows:



step 4 of 8

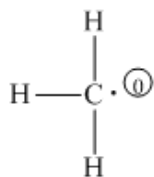
The given radical is as follows:



Calculate the formal charge for the above compound:

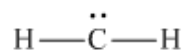
$$\text{Formal charge on C} = 4 - \left(1 + \left(\frac{6}{2} \right) \right) \\ = 0$$

The formal charge of the given radical is zero. Therefore, the given compound does not change.



step 5 of 8

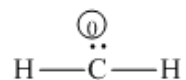
The given compound is as follows:



Calculate the formal charge for the above compound:

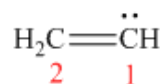
$$\begin{aligned}\text{Formal charge on C} &= 4 - \left(2 + \left(\frac{4}{2} \right) \right) \\ &= 0\end{aligned}$$

The formal charge of the given compound is zero. Therefore, the given compound does not change.



step 6 of 8

The given compound is as follows:

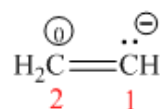


Calculate the formal charges for the above compound:

$$\begin{aligned}\text{Formal charge on C (1)} &= 4 - \left(2 + \left(\frac{6}{2} \right) \right) \\ &= -1\end{aligned}$$

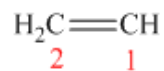
$$\begin{aligned}\text{Formal charge on C (2)} &= 4 - \left(0 + \left(\frac{8}{2} \right) \right) \\ &= 0\end{aligned}$$

So, the given compound with charges is as follows:



step 7 of 8

The given compound is as follows:

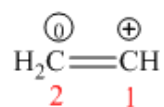


Calculate the formal charges for the above compound:

$$\begin{aligned}\text{Formal charge on C (1)} &= 4 - \left(0 + \left(\frac{6}{2} \right) \right) \\ &= +1\end{aligned}$$

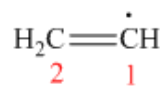
$$\begin{aligned}\text{Formal charge on C (2)} &= 4 - \left(0 + \left(\frac{8}{2} \right) \right) \\ &= 0\end{aligned}$$

So, the given compound with formal charges is as follows:



step 8 of 8

The given compound is as follows:



Calculate the formal charges for the above compound:

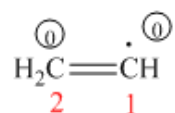
$$\text{Formal charge on C (1)} = 4 - \left(1 + \left(\frac{6}{2} \right) \right)$$

$= 0$

$$\text{Formal charge on C (2)} = 4 - \left(0 + \left(\frac{8}{2} \right) \right)$$

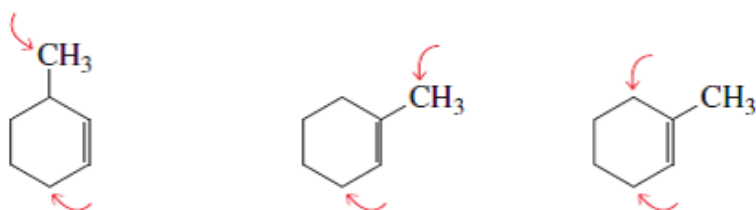
$= 0$

So, the given compound with formal charges is as follows:



Problem 71P

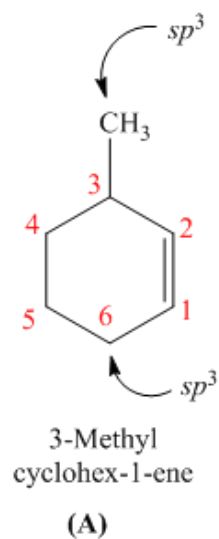
Do the sp^2 carbons and the indicated sp^3 carbons lie in the same plane?



Step-by-step solution

step 1 of 3

The given structure is drawn as follows:

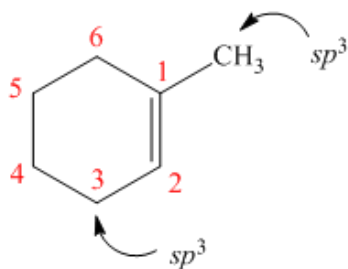


The carbons attached to the double bonds are sp^2 hybridized. The atoms attached to each of the sp^2 hybridized carbons all lie in the same plane and the others do not.

Therefore, the methyl group and the 6th carbon do not lie in the same plane.

step 2 of 3

The given structure is drawn as follows:



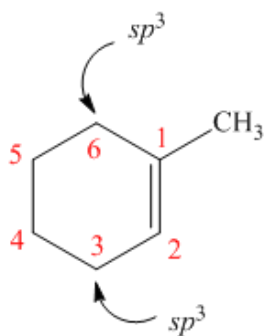
1-Methyl
cyclohex-1-ene
(B)

The carbons attached to the double bonds are sp^2 hybridized. The atoms attached to each of the sp^2 hybridized carbons all lie in the same plane and the others do not.

Therefore, the methyl group and the 3rd carbon lie in the same plane.

step 3 of 3

The given structure is drawn as follows:



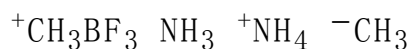
1-Methyl
cyclohex-1-ene
(C)

The carbons attached to the double bonds are sp^2 hybridized. The atoms attached to each of the sp^2 hybridized carbons all lie in the same plane and the others do not.

Therefore, the 3rd carbon and the 6th carbon lie in the same plane.

Problem 72P

- Which of the species have bond angles of 109.5 degree?
- Which of the species have bond angles of 120 degree? H_2O H_3O^+



Step-by-step solution

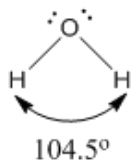
step 1 of 9

If the central atom has sp^3 hybridization and absence of lone pairs, then the bond angle will be 109.5° .

If the central atom involves in sp^2 hybridization, then the bond angle will be 120° .

Water (H_2O):

In water, the central atom is oxygen. Oxygen is attached to the two sigma bonds and has two lone pairs. So, total 4 hybrid orbitals are formed, which means the hybridization of oxygen is sp^3 .

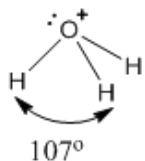


Due to presence of two lone pairs (i.e. lone pair lone pair repulsion), the bond angle in water molecule is 104.5° .

step 2 of 9

Hydronium ion (H_3O^+):

In hydronium ion, the central atom is oxygen. Oxygen is attached to the three sigma bonds and has a one lone pair. So, total 4 hybrid orbitals are formed, which means the hybridization of oxygen is sp^3 .

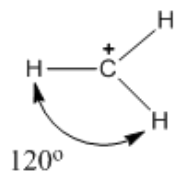


Due to presence of one lone pair (i.e. lone pair bond pair repulsion), the bond angle in water molecule is 107° .

step 3 of 9

Methyl cation (CH_3^+):

In methyl cation, the central atom is carbon. Carbon is attached to the three sigma bonds and has zero lone pairs. So, total 3 hybrid orbitals are formed, which means the hybridization of carbon is sp^2 .

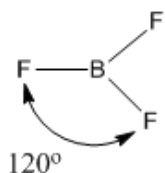


Thus, the bond angle in methyl cation is 120° .

step 4 of 9

Boron trifluoride (BF_3):

In boron trifluoride, the central atom is boron. Boron is attached to the three sigma bonds and has zero lone pairs. So, total 3 hybrid orbitals are formed, which means the hybridization of boron is sp^2 .

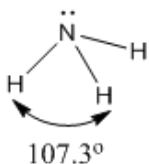


Thus, the bond angle in boron trifluoride is 120° .

step 5 of 9

Ammonia (NH_3):

In ammonia, the central atom is nitrogen. Nitrogen is attached to the three sigma bonds and has a one lone pair. So, total 4 hybrid orbitals are formed, which means the hybridization of nitrogen is sp^3 .

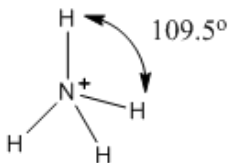


Due to presence of one lone pair (i.e. lone pair bond pair repulsion), the bond angle in water molecule is 107.3°

step 6 of 9

Ammonium ion (NH_4^+):

In ammonium ion, the central atom is nitrogen. Nitrogen is attached to the four sigma bonds and has zero lone pairs. So, total 4 hybrid orbitals are formed, which means the hybridization of nitrogen is sp^3 .



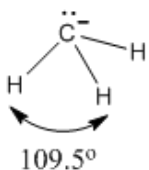
step 7 of 9

Thus, the bond angle in ammonium ion is 109.5° .

step 8 of 9

Methyl anion ($^-\text{CH}_3$):

In methyl anion, the central atom is carbon. Carbon is attached to the three sigma bonds and has one lone pair. So, total 4 hybrid orbitals are formed, which means the hybridization of carbon is sp^3 .



Thus, the bond angle in methyl anion is 109.5° .

step 9 of 9

(a)

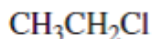
Thus, the species NH_4^+ and CH_3^- have bond angle of 109.5°

(b)

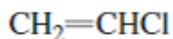
Thus, the species BF_3 and CH_3^+ have bond angle of 120° .

Problem 73P

Which compound has a longer C—Cl bond?



at one time it was used as a
refrigerant, an anesthetic, and a
propellant for aerosol sprays



used as the starting material
for the synthesis of a plastic that
is used to make bottles, flooring,
and clear packaging for food

Step-by-step solution

step 1 of 1

The length of $\text{C}-\text{Cl}$ bond depends on the hybridization of the carbon to which the chlorine is attached. The more s character in the orbital used by carbon to form the bond, the shorter and stronger is the bond and s orbital is closer to the nucleus than is a p orbital. Thus a $\text{C}-\text{Cl}$ bond formed by an sp^2 carbon is shorter than $\text{C}-\text{Cl}$ bond formed by an sp^3 carbon.

Therefore, in $\text{CH}_3\text{CH}_2\text{Cl}$, Cl is attached to sp^3 hybridized carbon and in $\text{CH}_2=\text{CHCl}$, Cl is attached to sp^2 carbon. Hence, $\text{C}-\text{Cl}$ bond in $\text{CH}_3\text{CH}_2\text{Cl}$ is longer.

Problem 74P

Which compound has a larger dipole moment, CHCl_3 or CH_2Cl_2 ?

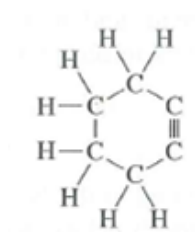
Step-by-step solution

step 1 of 1

In CHCl_3 , more Cl atoms are present, due to high electronegativity difference between $\text{C}-\text{Cl}$ and more bonds dipoles are present. Thus, dipole moment is higher than CH_2Cl_2 because it contains two $\text{C}-\text{Cl}$ bonds.

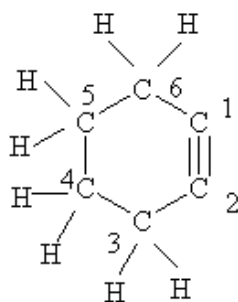
Problem 75P

Explain why the following compound is not stable:



Step-by-step solution

step 1 of 1



In the compound C_1 and C_2 are sp hybridized whereas the other carbons are sp^3 hybridized. So, in C_1 and C_2 bond arrangements must be linear but in other carbons, the bond arrangements have to be tetrahedral. Thus huge strain will be generated in the compound making it unstable.

Problem 76P

Explain why CH_3Cl has a greater dipole moment than CH_3F even though F is more electronegative than Cl.

Step-by-step solution

step 1 of 1

Dipole moment is the product of magnitude of charge and the distance of separation between the charges. In CH_3Cl because of large size of Cl the distance between two charges is more than the distance between charges in CH_3F , so distance is the dominating factor than electronegativity in these compounds. Therefore CH_3Cl has more dipole moment even though F is more electronegative in CH_3F .

Problem 77P

Step-by-step solution

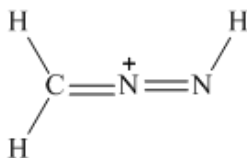
step 1 of 5

1. The molecule is CH_3N_2^+

Arrange the atoms in the molecule with single bonds

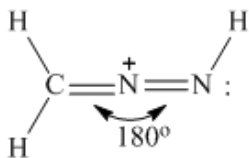


The bonds for carbon, middle hydrogen and nitrogen atoms are not satisfied, make double bond with one of the oxygen, attach two hydrogen atoms to carbon and hydrogen to the terminal nitrogen atom, and make a double bond between two nitrogen atoms, another double bond between carbon and middle nitrogen, place one unit of positive charge on middle nitrogen



1-methylenediazen-1-ium

Arrange a lone pair of electrons on the terminal nitrogen atoms, so the structure can be drawn as follows



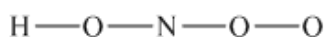
1-methylenediazen-1-ium

The middle nitrogen atom undergoes sp hybridization, so its structure is linear, and approximate bond angle is 180° , this molecule is a polar molecule

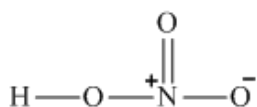
step 2 of 5

2. The molecule is HNO_3

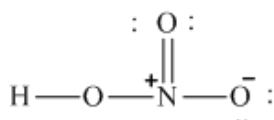
Arrange the atoms in the molecule with single bonds



The bonds for nitrogen, terminal oxygen atoms are not satisfied, after arranging the atoms the structure can be drawn as follows



Arrange two lone pairs of electrons on each oxygen atom, no need to place lone pairs of electrons on nitrogen, so the structure can be drawn as follows



nitric acid

The nitrogen atom undergoes sp^2 hybridization, so its structure is trigonal planar, and approximate bond angle is 120° , this molecule is a polar molecule

step 3 of 5

3. The molecule is N_3^-

Arrange the atoms in the molecule with single bonds



The bonds for nitrogen atoms are not satisfied, after arranging the atoms the structure can be drawn as follows



azide

Caution: A net charge appears to be present

Arrange lone pairs of electrons on terminal nitrogen atoms, no need to place lone pairs of electrons on middle nitrogen, so the structure can be drawn as follows



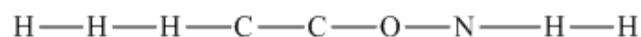
azide

The central nitrogen undergoes sp hybridization, so its structure is linear, and approximate bond angle is 180° , this molecule is a nonpolar molecule, because the central atom has same groups on both sides

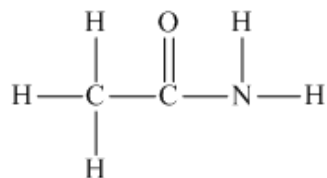
step 4 of 5

4. The molecule is CH_3CONH_2

Arrange the atoms in the molecule with single bonds

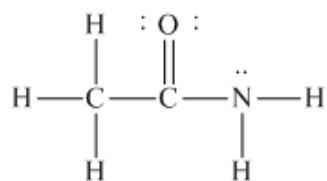


The bonds for nitrogen, carbon and hydrogen atoms are not satisfied, after arranging the atoms the structure can be drawn as follows



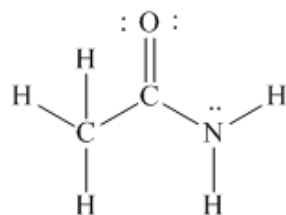
step 5 of 5

Arrange two lone pair of electrons on oxygen and nitrogen atoms, so the structure can be drawn as follows



acetamide

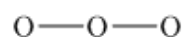
The middle carbon undergoes sp^2 hybridization, so its structure is trigonal planar, and approximate bond angle is 120° , this molecule is a polar molecule



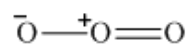
acetamide

5. The molecule is O_3

Arrange the atoms in the molecule with single bonds

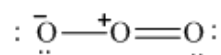


The bonds for terminal oxygen atoms are not satisfied, after arranging the atoms the structure can be drawn as follows



ozone

Arrange lone pair of electrons on terminal oxygen atoms, no need to place lone pair of electrons on middle oxygen, so the structure can be drawn as follows



ozone

The centre oxygen undergoes sp hybridization, so its structure is linear, and approximate bond angle is 180° , this molecule is polar molecule

Problem 78P

Step-by-step solution

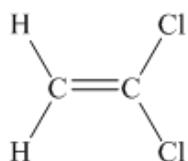
In a chemical bond the electron pair is shared by two atoms, the bond pair electrons has to lie at the midpoint between the two bonded atoms, but the more electronegative atom will attracts the bond pair of electrons towards itself, so bond pair of electrons lie close to the more electronegative atom participating in bonding this is nothing but polarity

The dipole moment of a molecule depends on the magnitude and the direction of the individual dipoles

Isomers have same chemical formula but differ in structure and properties, let us write the three constitutional isomers for $\text{C}_2\text{H}_2\text{Cl}_2$

1. the molecule is $\text{C}_2\text{H}_2\text{Cl}_2$

The structure can be drawn as follows

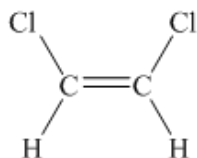


1,1-dichloroethene

The carbon chlorine bonds are more polar than carbon hydrogen bonds, the chlorine bonds attached to same carbon, hence these two chlorine bonds attracts the electron cloud in the double bond towards themselves, hence the molecule 1,1-dichloroethene is a polar molecule

2. the molecule is $\text{C}_2\text{H}_2\text{Cl}_2$

The structure can be drawn as follows

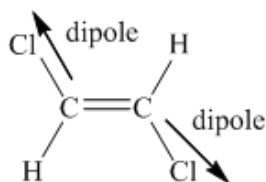


(Z)-1,2-dichloroethene

The carbon chlorine bonds are more polar than carbon hydrogen bonds, the chlorine bonds attached to same side means they are in cis isomeric condition, in cis isomers the dipole moments will not cancel, hence the molecule (Z)-1,2-dichloroethene is a polar molecule

3. the molecule is $\text{C}_2\text{H}_2\text{Cl}_2$

The structure can be drawn as follows



(E)-1,2-dichloroethene

The carbon chlorine bonds are more polar than carbon hydrogen bonds, the chlorine bonds attached to two carbon atoms and they are in opposite direction, the two dipoles have equal strength and they are in opposite direction, hence the dipoles get cancel, results in nonpolar molecule, so the molecule (E)-1,2-dichloroethene is nonpolar molecule