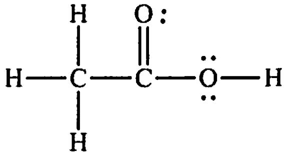
8

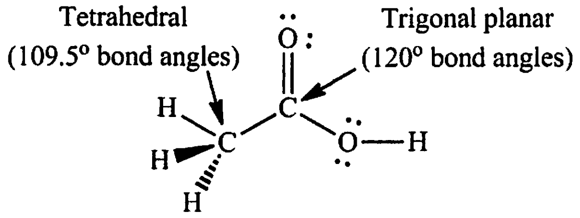
Covalent Compounds: Bonding Theories and Molecular Structure

8.1 ICl_4^-, 6 charge clouds, octahedral charge cloud arrangement, 4 bonding pairs, 2 lone pairs, square planar molecular geometry

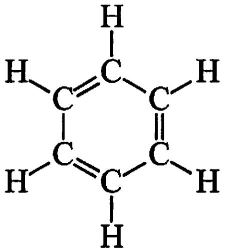
8.2 (a) 4 charge clouds, tetrahedral charge cloud arrangement, tetrahedral molecular geometry
(b) 5 charge clouds, trigonal bipyramidal charge cloud arrangement, seesaw molecular geometry.

8.3





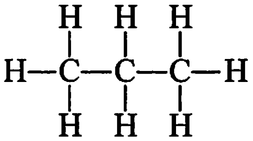
8.4



The bond angle around every carbon is 120° (trigonal planar). Benzene is a planar hexagon.

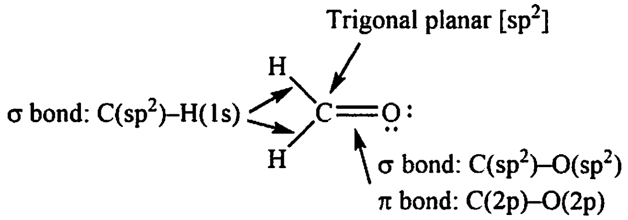
8.5 CH_2Cl_2; The C is sp^3 hybridized. The C–H bonds are formed by the overlap of one singly occupied sp^3 orbital on C with a singly occupied H 1s orbital. The C–Cl bonds are formed by the overlap of one singly occupied sp^3 orbital on C with a singly occupied Cl 2p orbital.

8.6

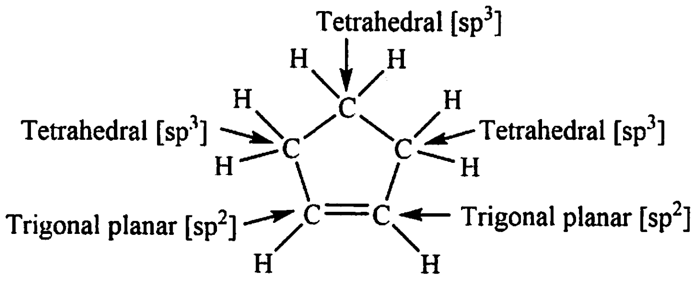


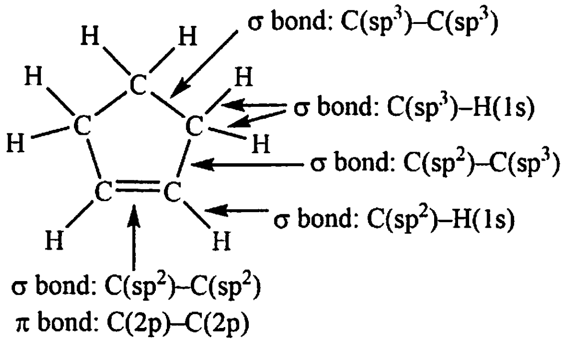
Each C is sp^3 hybridized. The C–C bonds are formed by the overlap of one singly occupied sp^3 hybrid orbital from each C. The C–H bonds are formed by the overlap of one singly occupied sp^3 orbital on C with a singly occupied H 1s orbital.

8.7



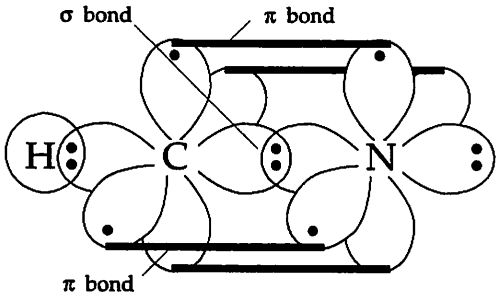
8.8





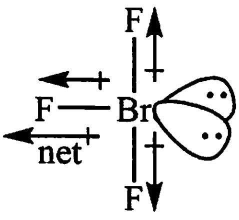
8.9 Both σ_1 and σ_2 $O(sp^2)–C(sp)$
π_1 $O(2p_x)–C(2p_x)$
π_2 $O(2p_y)–C(2p_y)$

8.10



In HCN the carbon is sp hybridized.

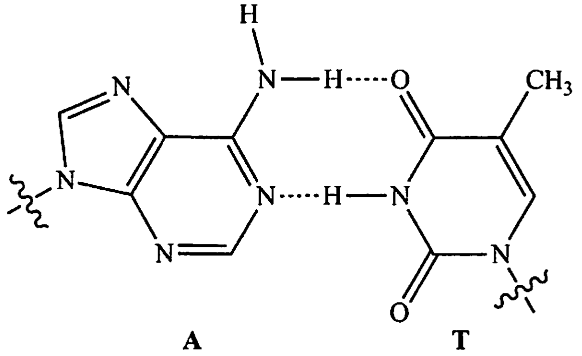
8.11 Only (c) BrF_3 has dipole moment.

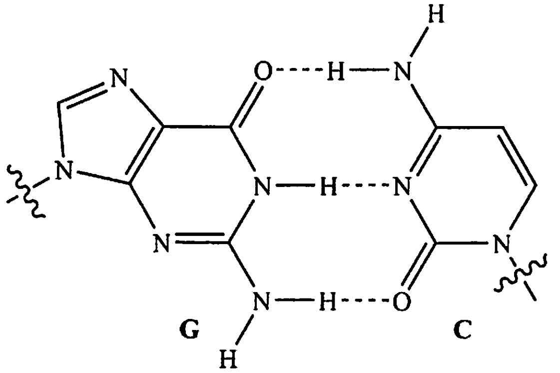


8.12 (a) CF_4 (b) CH_2F_2 (c) CHF_3 (d) CH_3F

8.13 (a) is the incorrect depiction of hydrogen bonding. A hydrogen bonded to carbon cannot hydrogen bond.

8.14 Hydrogen bonds are shown as dashed lines.





Because of the three hydrogen bonds, DNA regions that are high in G–C pairs would have the higher melting point.

8.15

Cl_2	dispersion
CCl_4	dispersion
CH_3F	dipole-dipole, dispersion
HF	hydrogen bonding, dipole-dipole, dispersion

(a) is false; (b), (c) & (d) are true

8.16

H_2S	dipole-dipole, dispersion
CH_3OH	hydrogen bonding, dipole-dipole, dispersion
C_2H_6	dispersion
Ar	dispersion

$Ar < C_2H_6 < H_2S < CH_3OH$

8.17 For He_2^+

σ^*_{1s} $\underline{\uparrow}$

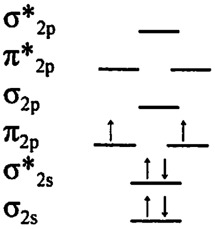
σ_{1s} $\underline{\uparrow\downarrow}$

$$He_2^+ \text{ Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2} = \frac{2-1}{2} = 1/2$$

He_2^+ should be stable with a bond order of 1/2.

8.18 The bond order in He_2^{2+} is 1, which is greater than the bond order of 1/2 in He_2^{+}; therefore He_2^{2+} is predicted to have a stronger bond and be a more stable species

8.19 For B_2

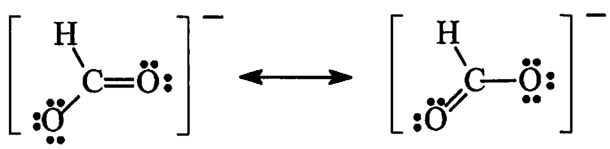


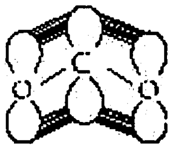
$$B_2 \text{ Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2} = \frac{4-2}{2} = 1$$

B_2 is paramagnetic because it has two unpaired electrons in the π_{2p} molecular orbitals.

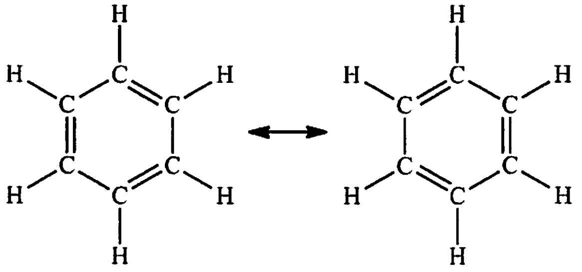
8.20 The bond orders are: $O_2^{2-} = 1$, $O_2^{-} = 1.5$, $O_2 = 2$, $O_2^{+} = 2.5$, $O_2^{2+} = 3$.
The order from weakest to strongest bond is: $O_2^{2-} < O_2^{-} < O_2 < O_2^{+} < O_2^{2+}$.
The order from shortest to longest bond is: $O_2^{2+} < O_2^{+} < O_2 < O_2^{-} < O_2^{2-}$.

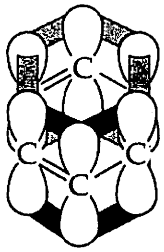
8.21





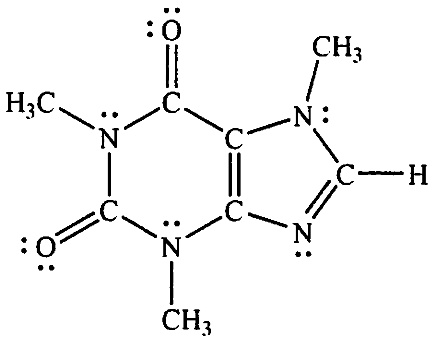
8.22





8.23 (a) C_a ~109°; C_b ~109°; C_c ~120°
(b) $C_a(sp^3)$; $C_b(sp^3)$; $C_c(sp^2)$
(c) σ bond: $C(sp^2)$–$C(sp^2)$; π bond: C(2p)–C(2p)

8.24 Only interactions (1) and (2) are hydrogen bonds. For interaction (3), a H bonded to a C cannot hydrogen bond. For interaction (4), a H cannot hydrogen bond with another H.

8.25 

(a) 8 lone pairs, 4 pi bonds (b) dispersion and dipole-dipole

(c) solid, large molecules with dipole-dipole and dispersion forces are usually solid.

8.26 The carbons in Ring 1 are all sp^2 hybridized and has a π molecular orbital that delocalizes electron density. The atoms in Ring 2 are all sp^3 hybridized with no delocalization.

Conceptual Problems

8.27 (a) square pyramidal (b) trigonal pyramidal
(c) square planar (d) trigonal planar

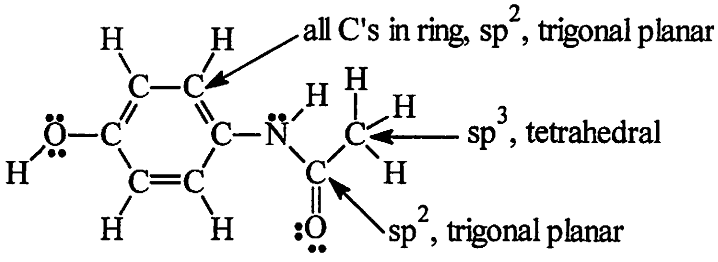
8.28 (a) trigonal bipyramidal (b) tetrahedral
(c) square pyramidal (4 ligands in the horizontal plane, including one hidden)

8.29 Molecular model (c) does not have a tetrahedral central atom. It is square planar.

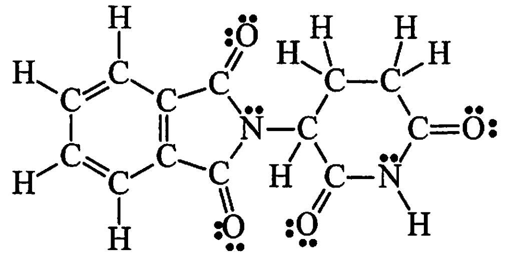
8.30 (a) sp^2 (b) sp (c) sp^3

8.31 The expected hybridizations of C and N in urea are sp^2 and sp^3, respectively. The expected bond angles are (i) N–C–O and N–C–N, ~120°, and (ii) C–N–H and H–N–H, ~109°. Based on the molecular model, the C and N are both sp^2 hybridized and all bond angles are ~120°.

8.32 (a) $C_8H_9NO_2$
(b), (c), and (d)



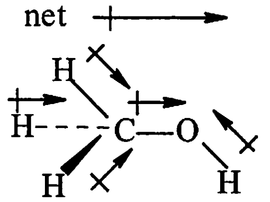
8.33 (a) $C_{13}H_{10}N_2O_4$
(b), (c), and (d)



All carbons that have only single bonds are sp^3 hybridized and have tetrahedral geometries. All carbons that have double bonds are sp^2 hybridized and have trigonal planar geometries.

8.34 The electronegative O atoms are electron rich (red), while the rest of the molecule is electron poor (blue).

8.35



8.36 The N atom is electron rich (red) because of its high electronegativity. The C and H atoms are electron poor (blue) because they are less electronegative.

8.37 (a) (i) is *trans* 1,2 dichloroethylene and (ii) is *cis* 1,2 dichloroethylene
(b) *cis* 1,2 dichloroethylene has a dipole moment of 2.39 D and *trans* 1,2 dichloroethylene has a dipole moment of 0.00 D.
(c) The polar molecule, *cis* 1,2 dichloroethylene, has the higher boiling point.

Section Problems
The VSEPR Model (Section 8.1)

8.38 From data in Table 8.1:
(a) trigonal planar (b) trigonal bipyramidal (c) linear (d) octahedral

8.39 From data in Table 8.1: (a) T shaped (b) bent (c) square planar

8.40 From data in Table 8.1:
(a) tetrahedral, 4 (b) octahedral, 6 (c) bent, 3 or 4
(d) linear, 2 or 5 (e) square pyramidal, 6 (f) trigonal pyramidal, 4

8.41 From data in Table 8.1:
(a) seesaw, 5 (b) square planar, 6 (c) trigonal bipyramidal, 5
(d) T shaped, 5 (e) trigonal planar, 3 (f) linear, 2 or 5

8.42

	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) H_2Se	2	2	bent
(b) $TiCl_4$	4	0	tetrahedral
(c) O_3	2	1	bent
(d) GaH_3	3	0	trigonal planar

8.43

	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) XeO_4	4	0	tetrahedral
(b) SO_2Cl_2	4	0	tetrahedral
(c) OsO_4	4	0	tetrahedral
(d) SeO_2	2	1	bent

8.44

	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) SbF_5	5	0	trigonal bipyramidal
(b) IF_4^+	4	1	see saw
(c) SeO_3^{2-}	3	1	trigonal pyramidal
(d) CrO_4^{2-}	4	0	tetrahedral

8.45

	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) NO_3^-	3	0	trigonal planar
(b) NO_2^+	2	0	linear
(c) NO_2^-	2	1	bent

8.46

	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) PO_4^{3-}	4	0	tetrahedral
(b) MnO_4^-	4	0	tetrahedral
(c) SO_4^{2-}	4	0	tetrahedral
(d) SO_3^{2-}	3	1	trigonal pyramidal
(e) ClO_4^-	4	0	tetrahedral
(f) SCN^- (C is the central atom)	2	0	linear

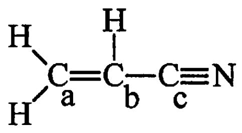
8.47

	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) XeF_3^+	3	2	T shaped
(b) SF_3^+	3	1	trigonal pyramidal
(c) ClF_2^+	2	2	bent
(d) CH_3^+	3	0	trigonal planar

8.48 (a) In SF_2 the sulfur is bound to two fluorines and contains two lone pairs of electrons. SF_2 is bent and the F–S–F bond angle is approximately 109°.
(b) In N_2H_2 each nitrogen is bound to the other nitrogen and one hydrogen. Each nitrogen has one lone pair of electrons. The H–N–N bond angle is approximately 120°.
(c) In KrF_4 the krypton is bound to four fluorines and contains two lone pairs of electrons. KrF_4 is square planar, and the F–Kr–F bond angle is 90°.
(d) In NOCl the nitrogen is bound to one oxygen and one chlorine and contains one lone pair of electrons. NOCl is bent, and the Cl–N–O bond angle is approximately 120°.

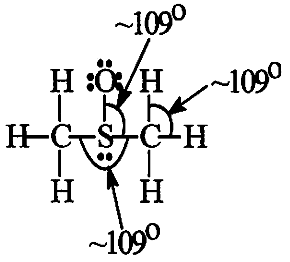
8.49 (a) In PCl_6^- the phosphorus is bound to six chlorines. There are no lone pairs of electrons on the phosphorus. PCl_6^- is octahedral, and the Cl–P–Cl bond angle is 90°.
(b) In ICl_2^- the iodine is bound to two chlorines and contains three lone pairs of electrons. ICl_2^- is linear, and the Cl–I–Cl bond angle is 180°.
(c) In SO_4^{2-} the sulfur is bound to four oxygens. There are no lone pairs of electrons on the sulfur. SO_4^{2-} is tetrahedral, and the O–S–O bond angle is 109.5°.
(d) In BO_3^{3-} the boron is bound to three oxygens. There are no lone pairs of electrons on the boron. BO_3^{3-} is trigonal planar, and the O–B–O bond angle is 120°.

8.50

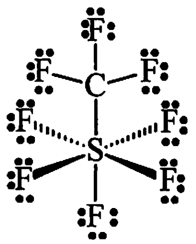


$H-C_a-H$	~ 120°	C_b-C_c-N	180°
$H-C_a-C_b$	~ 120°	C_a-C_b-H	~ 120°
$C_a-C_b-C_c$	~ 120°	$H-C_b-C_c$	~ 120°

8.51

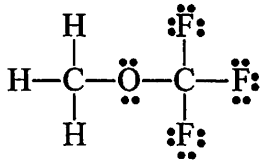


8.52



The bond angles are 109.5° around carbon and 90° around S.

8.53



The geometry is tetrahedral around both carbons and bent around the oxygen.

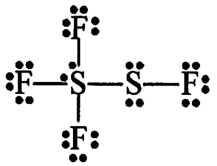
8.54 All six carbons in cyclohexane are bonded to two other carbons and two hydrogens (i.e., four charge clouds). The geometry about each carbon is tetrahedral with a C–C–C bond angle of approximately 109°. Because the geometry about each carbon is tetrahedral, the cyclohexane ring cannot be flat.

8.55 All six carbon atoms are sp^2 hybridized and the bond angles are ~120°. The geometry about each carbon is trigonal planar.

8.56

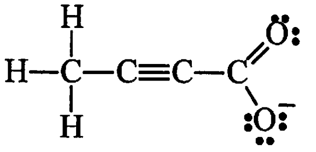
	Number of Bonded Atoms	Number of Lone Pairs	Shape
(a) BF_3	3	0	trigonal planar (120°)
PF_3	3	1	trigonal pyramidal (~107°)
PF_3 has the smaller F–X–F angles.			
(b) PCl_4^+	4	0	tetrahedral (109.5°)
ICl_2^-	2	3	linear (180°)
PCl_4^+ has the smaller Cl–X–Cl angles.			
(c) CCl_3^-	3	1	trigonal pyramidal (~107°)
PCl_6^-	6	0	octahedral (90°)
PCl_6^- has the smaller Cl–X–Cl angles.			

8.57 (a)



The left S has 5 electron clouds (4 bonding, 1 lone pair). The geometry about this S is seesaw. The right S has 4 electron clouds (2 bonding, 2 lone pairs). The geometry about this S is bent.

(b)



The left C has 4 electron clouds (4 bonding, 0 lone pairs). The geometry about this C is tetrahedral. The right C has 3 electron clouds (3 bonding, 0 lone pairs). The geometry about this C is trigonal planar. The central two C's have 2 electron clouds (2 bonding, 0 lone pairs). The geometry about these two C's is linear.

Valence Bond Theory and Hybridization (Sections 8.2–8.4)

8.58 In a π bond, the shared electrons occupy a region above and below a line connecting the two nuclei. A σ bond has its shared electrons located along the axis between the two nuclei.

8.59 Using the data here, the difference in energy between a carbon-carbon double bond and a carbon-carbon single bond is (728 kJ/mol – 350 kJ/mol) = 378 kJ/mol. This represents the energy of a π bond, which from this data would indicate that a π bond is stronger than a σ bond. However, numerous other sources list the average carbon-carbon double bond energy as ~615 kJ/mol, this would lead to a π bond energy of (615 kJ/mol – 350 kJ/mol) = 265 kJ/mol. This data would indicate that a σ bond is stronger than a π bond.

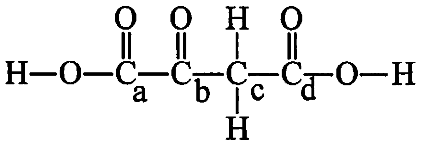
8.60 See Table 8.2. (a) sp (b) sp^2 (c) sp^3

8.61 See Table 8.2. (a) tetrahedral (b) trigonal planar (c) linear

8.62 (a) sp^2 (b) sp^2 (c) sp^3 (d) sp^2

8.63 (a) sp^3 (b) sp^2 (c) sp^2 (d) sp^3

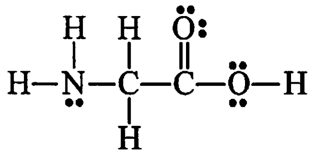
8.64



Carbons a, b, and d are sp^2 hybridized and carbon c is sp^3 hybridized.

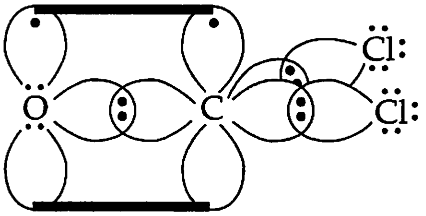
The bond angles around carbons a, b, and d are ~120°. The bond angles around carbon c are ~109°. The terminal H–O–C bond angles are ~109°.

8.65 (a)



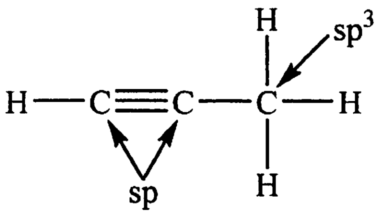
(b) H–C–H, ~109°; O–C–O, ~120°; H–N–H, ~107°
(c) N, sp^3; left C, sp^3; right C, sp^2

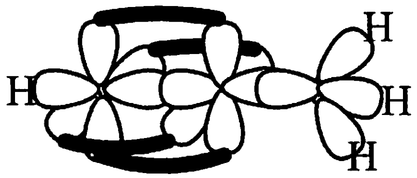
8.66



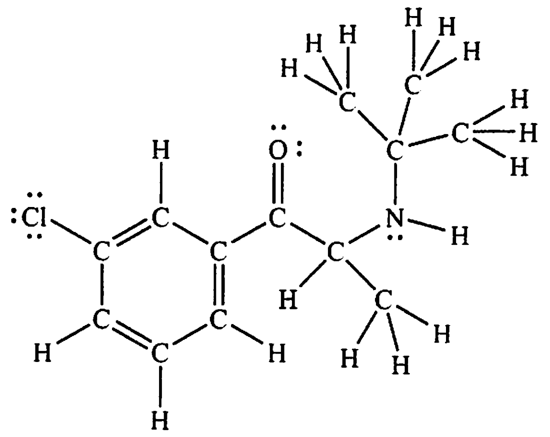
In Cl_2CO the carbon is sp2 hybridized.

8.67





8.68

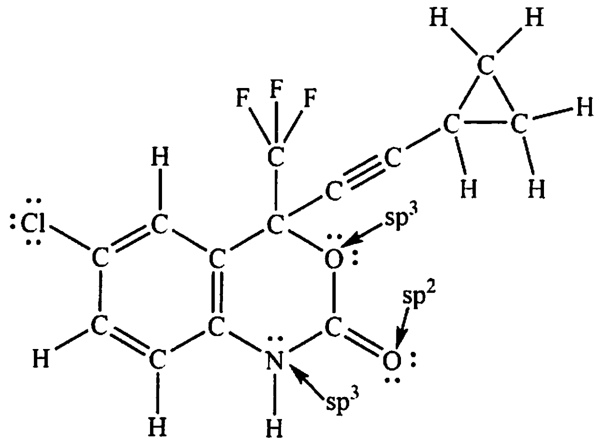


(a) There are 34 σ bonds and 4 π bonds.

(b) and (c) Each C with four single bonds is sp^3 hybridized with bond angles of 109.5°. Each C with a double bond is sp^2 hybridized with bond angles of 120°.

(d) The nitrogen is sp^3 hybridized.

8.69



(a) There are 32 σ bonds and 6 π bonds.

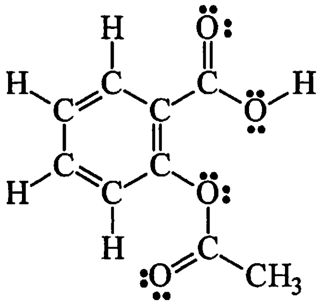
(b) Each C with four single bonds is sp^3 hybridized. Each C with a double bond is sp^2 hybridized. Each C with a triple bond is sp hybridized.

(c) and (d) see figure

8.70 Both the B and N are sp^2 hybridized. All bond angles are ~120°. The overall geometry of the molecule is planar.

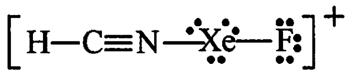
8.71 The triply bonded carbon atoms are sp hybridized. The theoretical bond angle for $\text{C–C}\equiv\text{C}$ is 180°. Benzyne is so reactive because the $\text{C–C}\equiv\text{C}$ bond angle is closer to 120° and is very strained.

8.72



21 σ bonds
5 π bonds
Each C with a double bond is sp^2 hybridized.
The $-CH_3$ carbon is sp^3 hybridized.

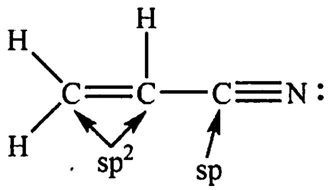
8.73



$$[\mathrm{H{-}C{\equiv}N{-}Xe{-}F}]^+$$

Both the carbon and nitrogen are sp hybridized.

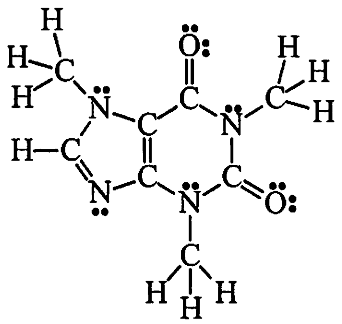
8.74 (a)



(b) There are 6 σ bonds and 3 π bonds.
(c) see figure
(d) The shortest bond is the C–N triple bond.

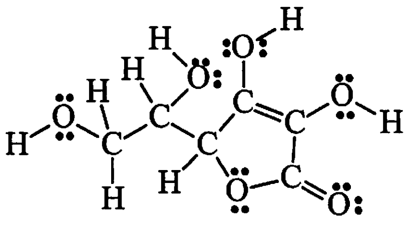
8.75 Every carbon is sp^2 hybridized. There are 18 σ bonds and 5 π bonds.

8.76



Each C with four single bonds is sp^3 hybridized.
Each C with a double bond is sp^2 hybridized.

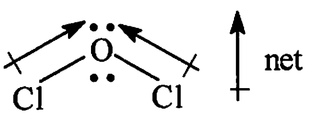
8.77



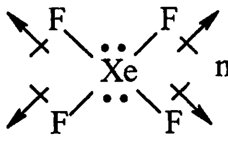
Each C with four single bonds is sp^3 hybridized.
Each C with a double bond is sp^2 hybridized.

Dipole Moments (Section 8.5)

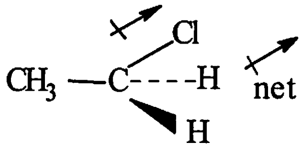
8.78 (a)

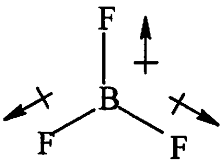


(b)

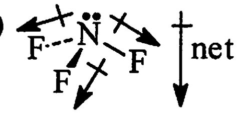


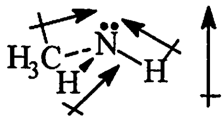
net dipole moment = 0

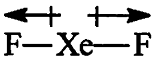
(c)


(d)


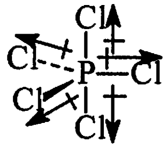
net dipole moment = 0

8.79 (a)


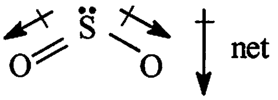
(b)


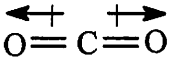
(c)


net dipole moment = 0

(d)


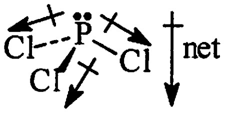
net dipole moment = 0

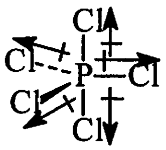
8.80




net dipole moment = 0

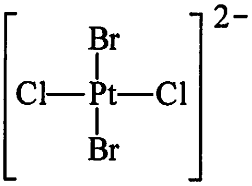
SO_2 is bent and the individual bond dipole moments add to give the molecule a net dipole moment.
CO_2 is linear and the individual bond dipole moments point in opposite directions to cancel each other out. CO_2 has no net dipole moment.

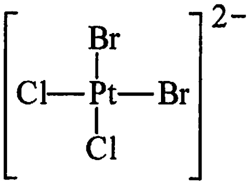
8.81




net dipole moment = 0

In both PCl_3 and PCl_5 the P–Cl bond is polar covalent. PCl_3 is trigonal pyramidal and the bond dipoles add to give the molecule a net dipole moment. PCl_5 is trigonal bipyramidal and the bond dipoles cancel. PCl_5 has no dipole moment.

8.82 (a)


(b)


8.83 SiF_4 is tetrahedral and nonpolar. SF_4 has one lone pair of electrons, adopts a see saw geometry, and is polar.

8.84 If a molecule has polar covalent bonds, the molecular shape (and location of lone pairs of electrons) determines whether the bond dipoles cancel and thus whether the molecule has a dipole moment.

8.85 Because chlorine is larger than fluorine, the charge separation is larger in CH_3Cl compared to CH_3F, resulting in CH_3Cl having a slightly larger dipole moment.

Intermolecular Forces (Section 8.6)

8.86 Dipole-dipole forces arise between molecules that have permanent dipole moments. London dispersion forces arise between molecules as a result of induced temporary dipoles.

8.87 (a) C_8H_{18} has the larger dispersion forces because of its longer hydrocarbon chain.
(b) HI has the larger dispersion forces because of the larger, more polarizable iodine.
(c) H_2Se has the larger dispersion forces because of the more polarizable and less electronegative Se.

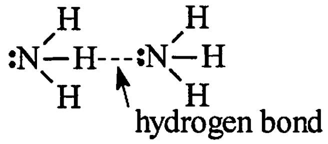
8.88 (a) $CHCl_3$ has a permanent dipole moment. Dipole-dipole intermolecular forces are important. London dispersion forces are also present.
(b) O_2 has no dipole moment. London dispersion intermolecular forces are important.
(c) Polyethylene, C_nH_{2n+2}. London dispersion intermolecular forces are important.
(d) CH_3OH has a permanent dipole moment. Dipole-dipole intermolecular forces and hydrogen bonding are important. London dispersion forces are also present.

8.89 (a) Xe has no dipole-dipole forces
(b) HF has the largest hydrogen bond forces
(c) Xe has the largest dispersion forces

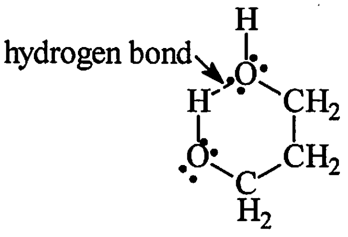
8.90 For CH_3OH and CH_4, dispersion forces are small. CH_3OH can hydrogen bond; CH_4 cannot. This accounts for the large difference in boiling points.
For 1-decanol and decane, dispersion forces are comparable and relatively large along the C–H chain. 1-decanol can hydrogen bond; decane cannot. This accounts for the 57 °C higher boiling point for 1-decanol.

8.91 Al_2O_3, ionic (greater lattice energy than NaCl because of higher ion charges);
F_2, dispersion; H_2O, H–bonding, dipole-dipole; Br_2, dispersion (larger and more polarizable than F_2), ICl, dipole-dipole, NaCl, ionic
rank according to normal boiling points: $F_2 < Br_2 < ICl < H_2O < NaCl < Al_2O_3$

8.92



8.93



8.94 Illustrations (ii) and (iii) depict the hydrogen bonding that occurs between methylamine and water.

8.95 (a) No.
(b) Illustration (ii) depicts the hydrogen bonding that occurs between dimethyl ether and water.

Molecular Orbital Theory (Sections 8.7– 8.9)

8.96 Electrons in a bonding molecular orbital spend most of their time in the region between the two nuclei, helping to bond the atoms together. Electrons in an antibonding molecular orbital cannot occupy the central region between the nuclei and cannot contribute to bonding.

8.97 The additive combination of two 2s orbitals is lower in energy than the two isolated 2s orbitals and is called a bonding molecular orbital. The subtractive combination of two 2s orbitals is higher in energy than the two isolated 2s orbitals and is called an antibonding molecular orbital.

8.98

	O_2^+	O_2	O_2^-
σ^*_{2p}	—	—	—
π^*_{2p}	↑ —	↑ ↑	↑↓ ↑
π_{2p}	↑↓ ↑↓	↑↓ ↑↓	↑↓ ↑↓
σ_{2p}	↑↓	↑↓	↑↓
σ^*_{2s}	↑↓	↑↓	↑↓
σ_{2s}	↑↓	↑↓	↑↓

$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

O_2^+ bond order $= \frac{8-3}{2} = 2.5$ $\qquad$ O_2 bond order $= \frac{8-4}{2} = 2$

O_2^- bond order $= \frac{8-5}{2} = 1.5$

All are stable with bond orders between 1.5 and 2.5. All have unpaired electrons.

8.99

	N_2^+	N_2	N_2^-
σ^*_{2p}	—	—	—
π^*_{2p}	— —	— —	↑ —
σ_{2p}	↑	↑↓	↑↓
π_{2p}	↑↓ ↑↓	↑↓ ↑↓	↑↓ ↑↓
σ^*_{2s}	↑↓	↑↓	↑↓
σ_{2s}	↑↓	↑↓	↑↓

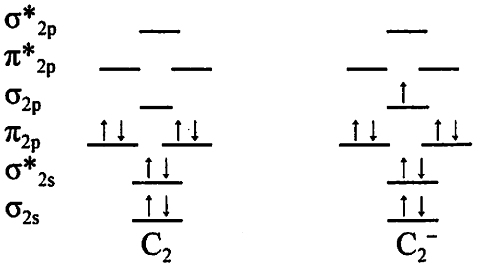
$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

N_2^+ bond order $= \frac{7-2}{2} = 2.5$ $\qquad$ N_2 bond order $= \frac{8-2}{2} = 3$

N_2^- bond order $= \dfrac{8 - 3}{2} = 2.5$

All are stable with bond orders of either 3 or 2.5. N_2^+ and N_2^- contain unpaired electrons.

8.100



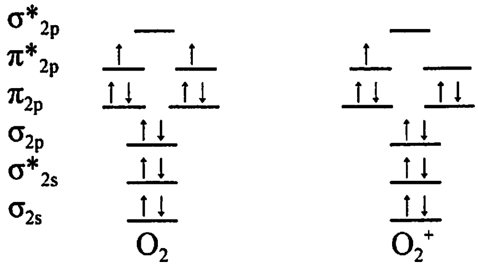
$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

(a) C_2 bond order $= \dfrac{6 - 2}{2} = 2$

(b) Add one electron because it will go into a bonding molecular orbital.

(c) C_2^- bond order $= \dfrac{7 - 2}{2} = 2.5$

8.101



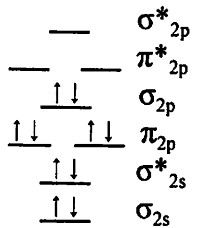
$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

(a) O_2 bond order $= \dfrac{8 - 4}{2} = 2$

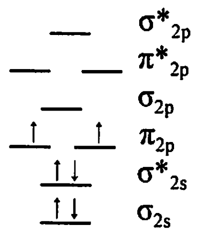
(b) Remove one electron because it will come out of an anti bonding molecular orbital.

(c) O_2^+ bond order $= \dfrac{8 - 3}{2} = 2.5$

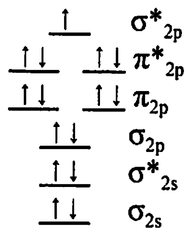
8.102



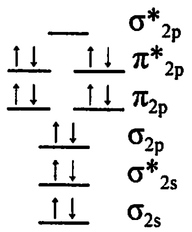
(a) C_2^{2-}
diamagnetic



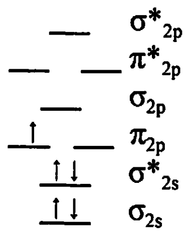
(b) C_2^{2+}
paramagnetic



(c) F_2^-
paramagnetic

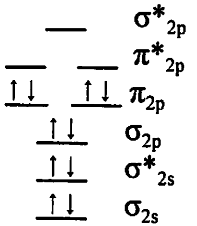


(d) Cl_2
diamagnetic

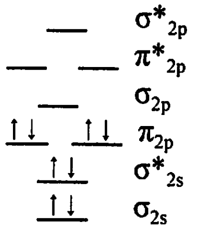


(e) Li_2^+
paramagnetic

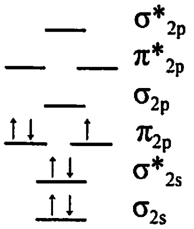
8.103



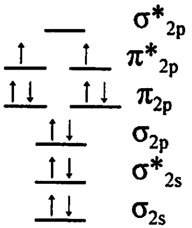
(a) O_2^{2+}
diamagnetic



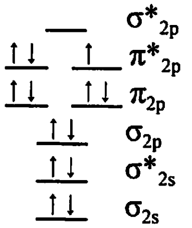
(b) N_2^{2+}
diamagnetic



(c) C_2^+
paramagnetic

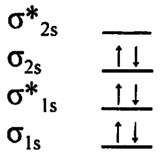


(d) F_2^{2+}
paramagnetic



(e) Cl_2^+
paramagnetic

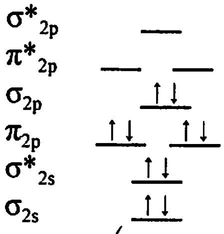
8.104 Li_2



$$Li_2 \text{ Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2} = \frac{4-2}{2} = 1$$

The bond order for Li_2 is 1, and the molecule is likely to be stable.

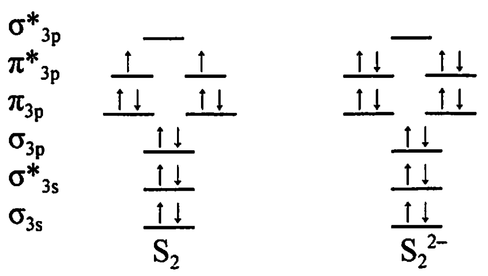
8.105 C_2^{2-}



$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

C_2^{2-} bond order $= \frac{8-2}{2} = 3$; there is a triple bond between the two carbons.

8.106 (a)



(b) S_2 would be paramagnetic with two unpaired electrons in the π^*_{3p} MOs.

(c)

$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

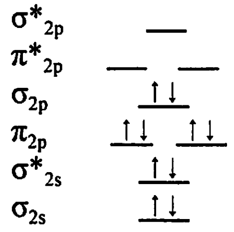
$$S_2 \quad \text{bond order} = \frac{8-4}{2} = 2$$

(d)

$$S_2^{2-} \quad \text{bond order} = \frac{8-6}{2} = 1$$

The two added electrons go into the antibonding π^*_{3p} MOs, the bond order drops from 2 to 1, and the bond length in S_2^{2-} should be longer than the bond length in S_2.

8.107 (a) CO



(b) All electrons are paired, CO is diamagnetic.

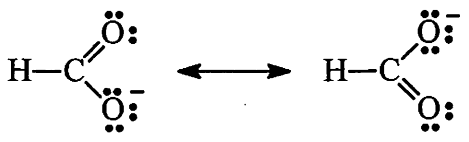
(c)

$$\text{Bond order} = \frac{\begin{pmatrix}\text{number of}\\ \text{bonding electrons}\end{pmatrix} - \begin{pmatrix}\text{number of}\\ \text{antibonding electrons}\end{pmatrix}}{2}$$

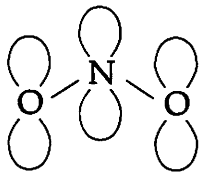
$$\text{CO} \quad \text{bond order} = \frac{8-2}{2} = 3$$

The bond order here matches that predicted by the electron-dot structure (:C≡O:).

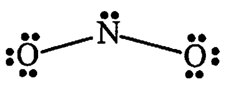
(d)



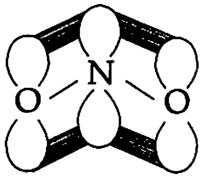
8.108



p orbitals in NO_2^-

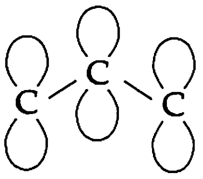


NO_2^- showing only the σ bonds (N is sp^2 hybridized)

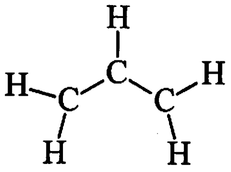


delocalized MO model for π bonding in NO_2^-

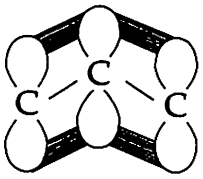
8.109



p orbitals in allyl cation



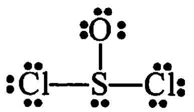
allyl cation showing only the σ bonds (each C is sp^2 hybridized)



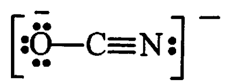
delocalized MO model for π bonding in the allyl cation

Multiconcept Problems

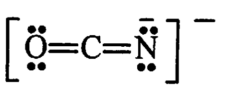
8.110 (a) $H{-}C{\equiv}C{-}H$ (b) $H{-}\ddot{N}{=}\ddot{N}{-}H$ (c)



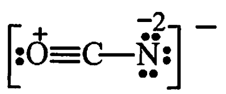
8.111 (a) (1)



(2)



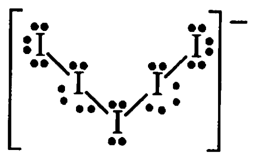
(3)



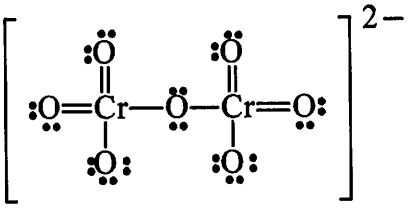
(b) Structure (1) makes the greatest contribution to the resonance hybrid because of the –1 formal charge on the oxygen. Structure (3) makes the least contribution to the resonance hybrid because of the +1 formal charge on the oxygen.

(c) and (d) OCN^- is linear because the C has 2 charge clouds. It is sp hybridized in all three resonance structures. It forms two π bonds.

8.112

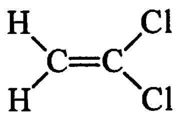


8.113 (a)

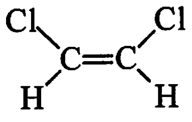


(b) Each Cr atom has 6 pairs of electrons around it. The likely geometry about each Cr atom is tetrahedral because each Cr has 4 charge clouds.

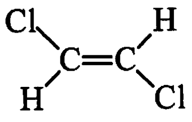
8.114 (a)



polar



polar



nonpolar

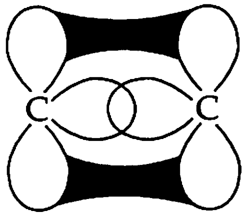
(b) All three molecules are planar. The first two structures are polar because they both have an unsymmetrical distribution of atoms about the center of the molecule (the middle of the double bond), and bond polarities do not cancel. Structure 3 is nonpolar because the H's and Cl's, respectively, are symmetrically distributed about the center of the molecule, both being opposite each other. In this arrangement, bond polarities cancel.

(c) 200 nm = 200 x 10^{-9} m

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ \text{J}\cdot\text{s})(3.00 \times 10^{8}\ \text{m/s})}{200 \times 10^{-9}\ \text{m}}(6.022 \times 10^{23}/\text{mol})$$

E = 5.99 x 10^5 J/mol = 599 kJ/mol

(d)



The π bond must be broken before rotation can occur.

8.115 (a) Each carbon is sp^2 hybridized.

(b) and (c)

antibonding	___
antibonding	___ ___
nonbonding	↑↓ ↑↓
bonding	↑↓ ↑↓
bonding	↑↓

(d) The cyclooctatetraene dianion has only paired electrons and is diamagnetic.