Covalent Bonding

Section 8.1 The Covalent Bond

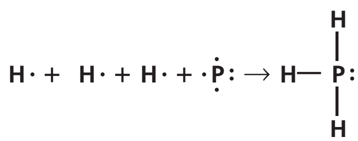
pages 240–247

Practice Problems

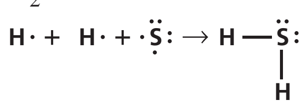
page 244

Draw the Lewis structure for each molecule.

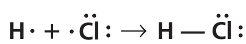
1. PH_3



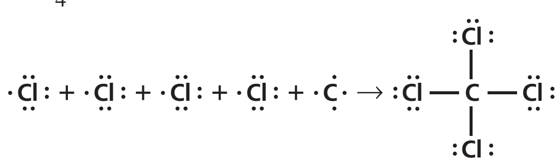
2. H_2S



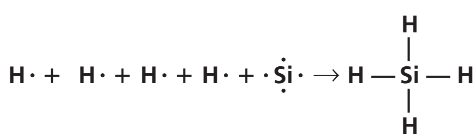
3. HCl



4. CCl_4



5. SiH_4



6. Challenge Draw a generic Lewis structure for a molecule formed between atoms of group 1 and group 16 elements.

Using 1 and 16 to represent atoms of groups 1 and 16, respectively, the generic structure is:

Section 8.1 Assessment

page 247

7. Identify the type of atom that generally forms covalent bonds.

The majority of covalent bonds form between nonmetallic elements.

8. Describe how the octet rule applies to covalent bonds.

Atoms share valence electrons; the shared electrons complete the octet of each atom.

9. Illustrate the formation of single, double, and triple covalent bonds using Lewis structures.

Student Lewis structures should show the sharing of a single pair of electrons, two pairs of electrons, and three pairs of electrons, respectively, for single, double, and triple covalent bonds.

10. Compare and contrast ionic bonds and covalent bonds.

Valence electrons are involved in both types of bonds. In covalent bonds, atoms share electrons, whereas is ionic bonds, electrons are transferred between atoms.

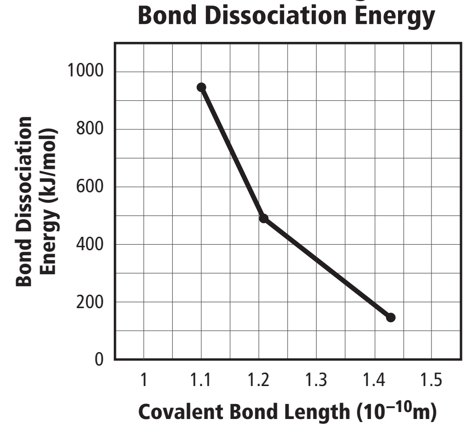
11. Contrast sigma bonds and pi bonds.

A sigma bond is a single covalent bond formed from the direct overlap of orbitals. A pi bond is the parallel overlap of p orbitals.

12. Apply Create a graph using the bond-dissociation energy data in Table 8.2 and the bond-length data in Table 8.1. Describe the relationship between bond length and bond-dissociation energy.

Student graphs should show that as bond length decreases the bond dissociation energy increases.

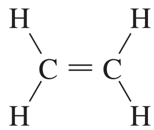
Covalent Bond Length vs. Bond Dissociation Energy



13. Predict the relative bond energies needed to break the bonds in the structures below.

a. H—C≡C—H

C—H: less energy than C≡C

b. 

C—H: less energy than C=C

Section 8.2 Naming Molecules

pages 248–252

Practice Problems

pages 249–251

Name each of the binary covalent compounds listed below.

14. CO_2

carbon dioxide

15. SO_2

sulfur dioxide

16. NF_3

nitrogen trifluoride

17. CCl_4

carbon tetrachloride

18. Challenge What is the formula for diarsenic trioxide?

Ar_2O_3

Name the following acids. Assume each compound is dissolved in water.

19. HI

hydroiodic acid

20. $HClO_3$

chloric acid

21. $HClO_2$

chlorous acid

22. H_2SO_4

sulfuric acid

23. H_2S

hydrosulfuric acid

24. Challenge What is the formula for periodic acid?

HIO_4

Give the formula for each compound.

25. silver chloride

AgCl

26. dihydrogen oxide

H_2O

27. chlorine trifluoride

ClF_3

28. diphosphorus trioxide

P_2O_3

29. disulfur decafluoride

S_2F_{10}

30. Challenge What is the formula for carbonic acid?

H_2CO_3

Section 8.2 Assessment

page 252

31. Summarize the rules for naming binary molecular compounds.

Name the first element in the formula first. Name the second element using its root plus the suffix *–ide.* Add prefixes to indicate the number of atoms of each element present.

32. Define a binary molecular compound.

a molecule composed of only two nonmetal elements

33. Describe the difference between a binary acid and an oxyacid.

A binary acid contains hydrogen and one other element. An oxyacid contains hydrogen, another element, and oxygen.

34. Apply Using the system of rules for naming binary molecular compounds, describe how you would name the molecule N_2O_4.

There are two atoms of nitrogen; use the prefix *di–* with the name nitrogen. There are four atoms of oxygen, so use the prefix *tetra–* + the root of oxygen + the ending *–ide*. The name is dinitrogen tetroxide.

35. Apply Write the molecular formula for each of these compounds:, iodic acid, disulfur trioxide, dinitrogen monoxide, hydrofluoric acid.

HIO_3, S_2O_3, N_2O, HF

36. State the molecular formula for each compound listed below.

a. dinitrogen trioxide

N_2O_3

b. nitrogen monoxide

NO

c. hydrochloric acid

HCl

d. chloric acid

$HClO_3$

e. sulfuric acid

H_2SO_4

f. sulfurous acid

H_2SO_3

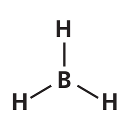
Section 8.3 Molecular Structures

pages 253–260

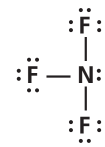
Practice Problems

pages 255–260

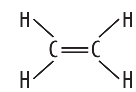
37. Draw the Lewis structure for BH_3.



38. Challenge A nitrogen trifluoride molecule contains numerous lone pairs. Draw its Lewis structure.

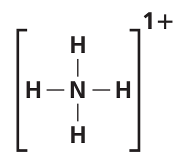


39. Draw the Lewis structure for ethylene, C_2H_4.

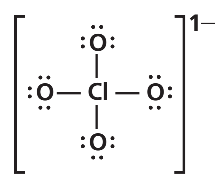


40. Challenge A molecule of carbon disulfide contains both lone pairs and multiple-covalent bonds. Draw its Lewis structure.

41. Draw the Lewis structure for ethylene, NH_4^+ ion.

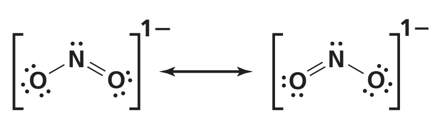


42. Challenge The ClO_4^- ion contains numerous lone pairs. Draw its Lewis structure.

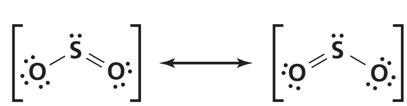


Draw the Lewis resonance structures for the following molecules.

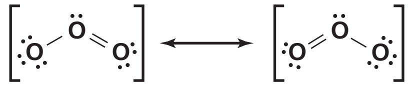
43. NO_2^-



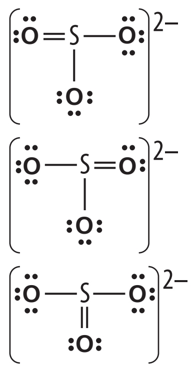
44. SO_2



45. O_3

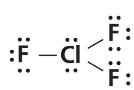


46. Challenge Draw the Lewis resonance structure for the ion SO_3^{2-}.

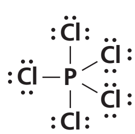


Draw the expanded octet Lewis structure of each molecule.

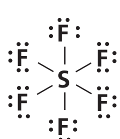
47. ClF_3



48. PCl_5



49. Challenge Draw the Lewis structure for the molecule formed when six fluorine atoms and one sulfur atom bond covalently.



Section 8.3 Assessment

page 260

50. Describe the information contained in a structural formula.

types of atoms, number of atoms, and a rough approximation of the molecular shape

51. State the steps used to draw Lewis structures.

1) determine central atom and terminal atoms, 2) determine number of bonding electrons, 3) determine bonding pairs, 4) connect terminal atoms to the central atom with single bonds, 5) determine remaining number of bonding pairs, 6) apply octet rule and form double or triple bonds if needed

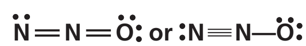
52. Summarize exceptions to the octet rule by correctly pairing these molecules and phrases: odd number of valence electrons, PCl_5, ClO_2, BH_3, expanded octet, less than an octet.

expanded octet, PCl_5; odd number of valence electrons, ClO_2; less than an octet, BH_3

53. Evaluate A classmate states that a binary compound having only sigma bonds displays resonance. Could the classmate's statement be true?

No, a molecule or polyatomic ion must have both a single bond and a double bond in order to display resonance. Only single bonds can be sigma bonds.

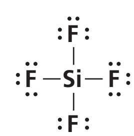
54. Draw the resonance structures for the dinitrogen oxide (N_2O) molecule.



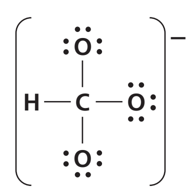
55. Draw the Lewis structure for CN^-, SiF_4, HCO_3^-, and AsF_6^-.

CN^- : $[:C\equiv N:]^-$

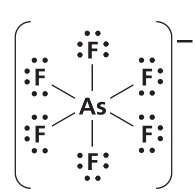
SiF_4 :



HCO_3^- :



AsF_6^- :



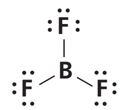
Section 8.4 Molecular Shapes

Practice Problems

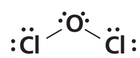
page 264

Determine the molecular shape, bond angle, and hybrid orbitals for each molecule.

56. BF_3 **trigonal planar, 120°, sp^2**

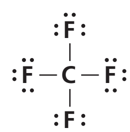


57. OCl_2 **bent, 104.5°, sp^3**



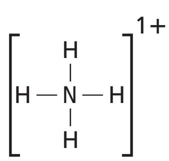
58. BeF_2 **linear, 180°, sp**

59. CF_4 **tetrahedral, 109°, sp^3**



60. Challenge For the NH_4^+ ion, identify its molecular shape, bond angle, and hybrid orbitals.

tetrahedral, 109°, sp^3



Section 8.4 Assessment

page 264

61. Summarize the VSEPR bonding theory.

VSEPR theory determines molecular geometry based on the repulsive nature of electron pairs around a central atom.

62. Define the term *bond angle*.

The bond angle is the angle formed by any two terminal atoms and the central atom.

63. Describe how the presence of a lone pair affects the spacing of shared bonding orbitals.

A lone pair occupies more space than a shared electron pair, thus, the presence of a lone pair pushes the bonding pairs closer together.

64. Compare the size of an orbital that has a shared electron pair with one that has a lone pair.

The orbital containing a lone electron pair occupies more space than a shared electron pair.

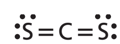
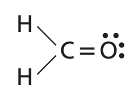
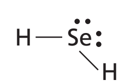
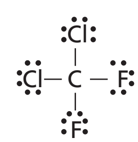
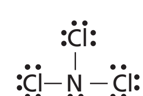
65. Identify the type of hybrid orbitals present and bond angles for a molecule with a tetrahedral shape.

sp^3 and 109°

66. Compare the molecular shapes and hybrid orbitals of PF_3 and PF_5 molecules. Explain why their shapes differ.

PF_3 is trigonal pyramidal with sp^3 hybrid orbitals. PF_5 is trigonal bipyramidal with sp^3d hybrid orbitals. Shape is determined by the type of hybrid orbital.

67. List in a table, the Lewis structure, molecular shape, bond angle, and hybrid orbitals for molecules of CS_2, CH_2O, H_2Se, CCl_2F_2, and NCl_3.

CS_2:		**linear, 180°, sp**
CH_2O:		**trigonal planar, 120°, sp^2**
H_2Se:		**bent, 104.5°, sp^3**
CCl_2F_2:		**tetrahedryl, 109°, sp^3**
NCL_3:		**trigonal pyramidal, 107°, sp^3**

Section 8.5 Electronegativity and Polarity

pages 265–270

Section 8.5 Assessment

page 270

68. Summarize how electronegativity difference is related to bond character.

The greater the electronegativity difference, the greater the ionic nature of the bond.

69. Describe a polar covalent bond.

A polar covalent bond has unequal sharing of electrons. The electrons are pulled toward one of the atoms, generating partial charges on the ends.

70. Describe a polar molecule.

A polar molecule is one that has a greater electron density on one side of the molecule.

71. List three properties of a covalent compound in the solid phase.

The solid state of a molecule is crystalline. A molecular solid is soft, a nonconductor, and has a low melting point.

72. Categorize bond types using electronegativity difference.

If the difference is zero, the bond is considered nonpolar covalent; if between zero and 0.4, mostly covalent; if between 0.4 and 1.7, polar covalent; if greater than 1.7, mostly ionic.

73. Generalize Describe the general characteristics of covalent network solids.

brittle, nonconductors of heat and electricity, extremely hard

74. Predict the type of bond that will form between the following pair of atoms:

a. H and S

electronegativity of S = 2.58
electronegativity of H = 2.20
EN difference = 0.38; mostly covalent

b. C and H

electronegativity of C = 2.55;
electronegativity of H = 2.20;
EN difference = 0.35; mostly covalent

c. Na and S.

electronegativity of S = 2.58;
electronegativity of Na = 0.93;
EN difference = 1.65; polar covalent

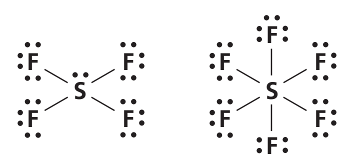
75. Identify each molecule as polar or nonpolar: SCl_2, CS_2, and CF_4.

SCl_2, polar; CS_2, nonpolar; CF_4, nonpolar

76. Determine whether a compound made of hydrogen and sulfur atoms is polar or nonpolar.

hydrogen and sulfur form H_2S, a molecule with a bent shape; the molecule is polar because it is asymmetric

77. Draw the Lewis structures for the molecules SF_4 and SF_6. Analyze each structure to determine whether the molecule is polar or nonpolar.



SF_4: polar **SF_6: nonpolar**

Chapter 8 Assessment

pages 274–277

Section 8.1

Mastering Concepts

78. What is the octet rule, and how is it used in covalent bonding?

Atoms lose, gain, or share electrons to end with a full outer energy level of eight electrons. Covalent bonding occurs when atoms share electrons to achieve an octet.

79. Describe the formation of a covalent bond.

The nucleus of one atom attracts the electrons of the other atom, and they share one or more pairs of electrons.

80. Describe the bonding in molecules.

Molecules bond covalently.

81. Describe the forces, both attractive and repulsive, that occur as two atoms come closer together.

Attractive forces occur between the nucleus of one atom and the electrons of the other atom. Repulsive forces occur between the nuclei of the two atoms and between the electrons of the two atoms. As the atoms approach, the net force of attraction increases. At a certain optimal distance between atoms, the net attractive force is maximized. If the atoms move closer than the optimal distance, repulsive force exceeds attractive force. See Figure 8.2 on page 241.

82. How could you predict the presence of a sigma or pi bond in a molecule?

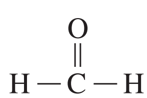
A single covalent bond is always a sigma bond; a double bond consists of a sigma bond and a pi bond; a triple bond consists of one sigma and two pi bonds.

Mastering Problems

83. Give the number of valence electrons in N, As, Br, and Se. Predict the number of covalent bonds needed for each of these elements to satisfy the octet rule.

N: 5, 3; As: 5, 3; Br: 7, 1; Se: 6, 2

84. Locate the sigma and pi bonds in each of the molecules shown below.

a. 

single bonds: sigma bonds; double bond: one sigma bond and one pi bond

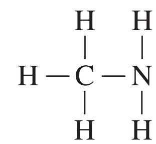
b. H—C≡C—H

single bonds: sigma bonds; triple bond: one sigma and two pi bonds

85. In the molecules CO, CO_2, and CH_2O, which C—O bond is the shortest? Which C—O bond is the strongest?

The triple bond in CO is the shortest and the strongest.

86. Consider the carbon-nitrogen bonds shown below:

$C\equiv N^-$ and 

Which bond is shorter? Which is stronger?

The triple bond in $C\equiv N^-$ is shorter and stronger.

87. Rank each of the molecules below in order of the shortest to the longest sulfur-oxygen bond length.

a. SO_2

b. SO_3^{2-}

c. SO_4^{2-}

a, c, b

Section 8.2

Mastering Concepts

88. Explain how molecular compounds are named.

Naming follows a specific set of rules depending on whether the compound forms an acidic aqueous solution. Answers should agree with Figure 8.12 on page 252.

89. When is a molecular compound named as an acid?

when it releases H^+ in water solution

90. Explain the difference between sulfur hexafluoride and disulfur tetrafluoride.

Sulfur hexafluoride is SF_6, which has one atom of sulfur bonded with six atoms of fluorine. Disulfur tetrafluoride is S_2F_4, which has two atoms of sulfur bonded with four atoms of fluorine.

91. Watches The quartz crystals used in watches are made of silicon dioxide. Explain how you use the name to determine the formula for silicon dioxide

The name silicon indicates one atom of Si. The prefix di- means two and oxide indicates oxygen. The correct formula is SiO_2.

Mastering Problems

92. Complete **Table 8.8**.

Acid Names	
Formula	**Name**
$HClO_2$	**chlorous acid**
H_3PO_4	**phosphoric acid**
H_2Se	**hydroselenic acid**
$HClO_3$	**chloric acid**

93. Name each molecule.

a. NF_3

nitrogen trifluoride

b. NO

nitrogen monoxide

c. SO_3

sulfur trioxide

d. SiF_4

silicon tetrafluoride

94. Name each molecule.

a. SeO_2

selenium dioxide

b. SeO_3

selenium trioxide

c. N_2F_4

dinitrogen tetrafluoride

d. S_4N_4

tetrasulfur tetranitride

95. Write the formula for each molecule.

a. sulfur difluoride

SF_2

b. silicon tetrachloride

$SiCl_4$

c. carbon tetrafluoride

CF_4

d. sulfurous acid

H_2SO_3

96. Write the formula for each molecule.

a. silicon dioxide

SiO_2

b. bromous acid

$HBrO_2$

c. chlorine trifluoride

ClF_3

d. hydrobromic acid

HBr

Section 8.3

Mastering Concepts

97. What must you know in order to draw the Lewis structure for a molecule?

the number of valence electrons for each atom

98. Doping Agent Material scientists are studying the properties of polymer plastics doped with AsF_5. Explain why the compound AsF_5 is an exception to the octet rule.

Arsenic has five bonding positions with a total of 10 shared electrons. This is greater that the eight electrons that occupy an octet.

99. Reducing Agent Boron trihydride (BH_3) is used as reducing agent in organic chemistry. Explain why BH_3 often forms coordinate covalent bonds with other molecules.

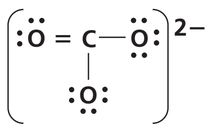
BH_3 only has 6 electrons and does not have an electron arrangement with a low amount of potential energy. It will share a lone pair with another molecule to form this electron arrangement.

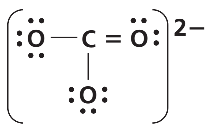
100. Antimony and chlorine can form antimony trichloride and antimony pentachloride. Explain how these elements are able to form two different compounds.

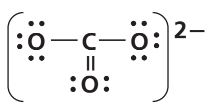
Antimony has five valence electrons, one lone pair, and three positions where it can share one electron with a chlorine atom. This will form $SbCl_3$. Antimony can also expand its octet and bond with all five valence electrons, forming $SbCl_5$.

Mastering Problems

101. Draw three resonance structures for the polyatomic ion CO_3^{2-}.

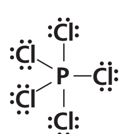




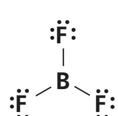


102. Draw the Lewis structure for these molecules, each of which has a central atom that does not obey the octet rule.

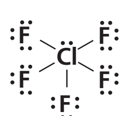
a. PCl_5



b. BF_3



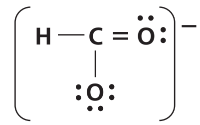
c. ClF_5

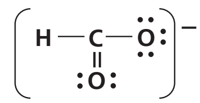


d. BeH_2

H—Be—H

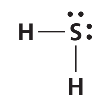
103. Draw two resonance structures for the polyatomic ion HCO_2^-.



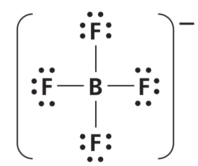


104. Draw the Lewis structure for a molecule of each of these compounds and ions.

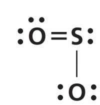
a. H_2S



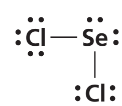
b. BF_4^-



c. SO_2



d. $SeCl_2$



105. Which elements in the list below are capable of forming molecules in which one of its atoms has an expanded octet? Explain your answer.

a. B
b. C
c. P
d. O
e. Se

P and Se because they are from period 3 or higher and have a d sublevel available

Section 8.4

Mastering Concepts

106. What is the basis of the VSEPR model?

the repulsive nature of electron pairs around a central atom

107. What is the maximum number of hybrid orbitals a carbon atom can form?

four

108. What is the molecular shape of each molecule? Estimate the bond angle for each molecule, assuming that there is not a lone pair.

a. A—B

linear, 180°

b. A—B—A

linear, 180°

c.
$$\begin{array}{c} \text{A—B—A} \\ | \\ \text{A} \end{array}$$

trigonal planar, 120°

d.
$$\begin{array}{c} \text{A} \\ | \\ \text{A—B—A} \\ | \\ \text{A} \end{array}$$

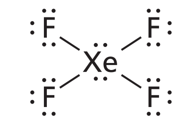
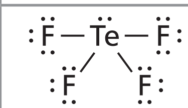
tetrahedral, 109°

109. Parent Compound PCl_5 is used as a parent compound to form many other compounds. Explain the theory of hybridization and determine the number of hybrid orbitals present in a molecule of PCl_5.

The theory of hybridization explains the shapes of molecules by the formation of identical hybrid orbitals from the atomic orbitals of the atoms in the molecule; five identical sp^3d orbitals

Mastering Problems

110. Complete **Table 8.9** by identifying the expected hybrid on the central atom. You might find drawing the molecule's Lewis structure helpful.

Structures		
Formula	**Hybrid Orbital**	**Lewis Structure**
XeF_4	**sp^3d^2**	
TeF_4	**sp^3d**	
KrF_2	**sp^3d**	
OF_2	**sp^3**	

111. Predict the molecular shape of each molecule.

a. COS

linear

b. CF_2Cl_2

tetrahedral

112. For each molecule listed below, predict its molecular shape and bond angle, and identify the hybrid orbitals. Drawing the Lewis structure might help you.

a. SCl_2

bent, 104.5°, sp^3

b. NH_2Cl

trigonal pyramidal, 107°, sp^3

c. HOF

bent, 104.5°, sp^3

d. BF_3

trigonal planar, 120°, sp^2

Section 8.5

Mastering Concepts

113. Describe electronegativity trends in the periodic table.

It increases left to right in a period and decreases top to bottom in a group.

114. Explain the difference between nonpolar molecules and polar molecules.

A nonpolar molecule has a symmetric distribution of charge, while a polar molecule has a concentration of electrons on one side of the molecule.

115. Compare the location of bonding electrons in a polar covalent bond with those in a nonpolar covalent bond. Explain your answer.

Electrons in a polar bond are closer to the more electronegative atom because of unequal sharing. Those in a nonpolar bond are shared equally.

116. What is the difference between a covalent molecular solid and a covalent network solid? Do their physical properties differ? Explain your answer.

A covalent molecular solid is soft and has a low melting point because of weak intermolecular forces. A covalent network solid has a high melting point and is very hard because of the strength of the network of covalent bonds.

Mastering Problems

117. For each pair, indicate the more polar bond by circling the negative end of its dipole.

a. C—S, C—O

O is circled because it has the greatest electronegativity. C—O is the more polar bond. There is a greater electronegativity difference between C and O. Electronegativity S = 2.58; electronegativity C = 2.55; EN difference = 0.03; electronegativity O = 3.44; electronegativity C = 2.55; EN difference = 0.89

b. C—F, C—N

F is circled because it has the greatest electronegativity. C—F is the more polar bond. There is a greater electronegativity difference between C and F. Electronegativity F = 3.98; electronegativity C = 2.55; EN difference = 1.43. Electronegativity N = 3.04; electronegativity C = 2.55; EN difference = 0.49

c. P—H, P—Cl

Cl is circled because it has the greatest electronegativity. P—Cl is the more polar bond. There is a greater electronegativity difference between Cl and P. Electronegativity H = 2.20; electronegativity P = 2.19; EN difference = 0.01. Electronegativity Cl = 3.16; electronegativity P = 2.19; EN difference = 0.97

118. For each of the bonds listed, tell which atom is more negatively charged.

The most negatively charged atom has the greatest electronegativity. Use Figure 8.20.

a. C—H

C

b. C—N

N

c. C—S

S

d. C—O

O

119. Predict which bond is the most polar.

a. C—O

electronegativity O = 3.44, electronegativity C = 2.55, EN difference = 0.89

b. Si—O

electronegativity O = 3.44, electronegativity Si = 1.90, EN difference = 1.54

c. C—Cl

electronegativity Cl = 3.16, electronegativity C = 2.55, EN difference = 0.61

d. C—Br

electronegativity Br = 2.96, electronegativity C = 2.55, EN difference = 0.41

Si—O is the most polar because it has the greatest electronegativity difference.

120. Rank the bonds according to increasing polarity.

a. C—H

electronegativity C = 2.55, electronegativity H = 2.20, EN difference = 0.35

b. N—H

electronegativity N = 3.04, electronegativity H = 2.20, EN difference = 0.84

c. Si—H

electronegativity H = 2.20, electronegativity Si = 1.90, EN difference = 0.30

d. O—H

electronegativity O = 3.44, electronegativity H = 2.20, EN difference = 1.24

e. Cl—H

electronegativity Cl = 3.16, electronegativity H = 2.20, EN difference = 0.96

in order of increasing polarity: c, a, b, e, d

121. Refrigerant The refrigerant known as freon-14 is an ozone-damaging compound with the formula CF_4. Why is the CF_4 molecule nonpolar even though it contains polar bonds?

equal distribution of charge in a symmetrical molecule

122. Determine if these molecules and ion are polar. Explain your answers.

a. H_3O^+

polar, asymmetrical

b. PCl_5

nonpolar, symmetrical

c. H_2S

polar, asymmetrical

d. CF_4

nonpolar, symmetrical

123. Use Lewis structures to predict the molecular polarities for sulfur difluoride, sulfur tetrafluoride, and sulfur hexafluoride.

SF_2 and SF_4 are polar. SF_6 is nonpolar.

Mixed Review

124. Write the formula for each molecule.

a. chlorine monoxide

ClO

b. arsenic acid

H_3AsO_4

c. phosphorus pentachloride

PCl_5

d. hydrosulfuric acid

H_2S

125. Name each molecule.

a. PCl_3

phosphorus trichloride

b. Cl_2O_7

dichlorine heptoxide

c. P_4O_6

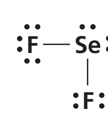
tetraphosphorus hexoxide

d. NO

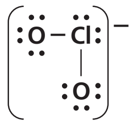
nitrogen monoxide

126. Draw the Lewis structure for each molecule or ion.

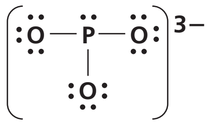
a. SeF_2



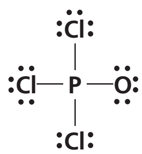
b. ClO_2^-



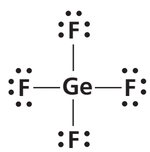
c. PO_3^{3-}



d. $POCl_3$



e. GeF_4



127. Determine which of the molecules are polar. Explain your answer.

a. CH_3Cl
b. ClF
c. NCl_3
d. BF_3
e. CS_2

The polar molecules are CH_3Cl, ClF, and NCl_3 because each molecule is asymmetric and the charge is not distributed evenly.

128. Arrange the bonds in order of least to greatest polar character.

a. C—O

electronegativity O = 3.44, electronegativity C = 2.55, difference = 0.89

b. Si—O

electronegativity O = 3.44, electronegativity Si = 1.90, difference = 1.54

c. Ge—O

electronegativity O = 3.44, electronegativity Ge = 2.01, difference = 1.43

d. C—Cl

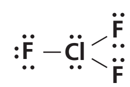
electronegativity Cl = 3.16, electronegativity CP = 2.55, difference = 0.61

e. C—Br

electronegativity Br = 2.96, electronegativity C = 2.55, difference = 0.41

in order of least to greatest polar bond character is the same as the order of increasing electronegativity differences (refer to Figure 8.21): e, d, a, c, b

129. Rocket Fuel In the 1950s, the reaction of hydrazine with chlorine trifluoride (ClF_3) was used as a rocket fuel. Draw the Lewis structure for ClF_3 and identify the hybrid orbitals.



sp^3d

130. Complete **Table 8.10,** which shows the number of electrons shared in a single covalent bond, a double covalent bond, and a triple covalent bond. Identify the group of atoms that will form each of these bonds.

Shared Pairs		
Bond Type	**Number of Shared Electrons**	**Atoms that Form the Bond**
Single covalent	**2 shared electrons**	**any halogen, group 17**
Double covalent	**4 shared electrons**	**A group 16 element**
Triple covalent	**6 shared electrons**	**A group 15 element**

Think Critically

131. Organize Design a concept map that explains how VSEPR model theory, hybridization theory, and molecular shape are related.

Concept maps will vary.

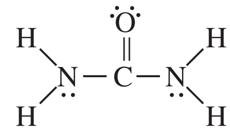
132. Compare and contrast the two compounds identified by the names arsenic(III) oxide and diarsenic trioxide.

The name arsenic(III) oxide states that arsenic has an oxidation number of 3+ and oxide is 2−. The correct formula is As_2O_3. The name diarsenic trioxide states that there are two atoms of arsenic and three atoms of oxygen. The correct formula is As_2O_3. Even though they are named differently, they both represent the same formula.

133. Make and Use Tables Complete **Table 8.11** using Chapters 7 and 8.

Solid	Bond Description	Characteristic of Solid	Example
Ionic	the electrostatic attraction of a positive ion for a negative ion	hard, rigid, brittle, crystalline, high melting point, nonconductor in the solid state	NaCl
Covalent molecular	the sharing of electrons between two atoms	soft, low melting point, non-conductor in the solid state	CO_2
Metallic	the attraction of a positive ion for delocalized electrons	a crystal that conducts heat and electricity, malleable, ductile, high melting point	Ag
Covalent network	atoms covalently bonded to many other atoms	crystal is hard, rigid, brittle, nonconductor	diamond

134. Apply Urea, whose structure is shown in the next column, is a compound used in manufacturing plastics and fertilizers. Identify the sigma bond, pi bonds, and lone pairs present in a molecule of urea.



Sigma bonds are the N—H bonds and the C—N bonds as well as one of the C—O bonds. The other C—O bond is a pi bond. The lone pairs are located on both N atoms and on the O atom.

135. Analyze For each of the characteristics listed below, identify the polarity of a molecule with that characteristic.

a. a solid at room temperature

polar

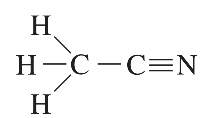
b. a gas at room temperature

nonpolar

c. attracted to an electric current

polar

136. Apply The structural formula for acetonitrile, CH_3CN, is shown below.



Examine the structure of acetonitrile molecule. Determine the number of carbon atoms in the molecule, identify the hybrid present in each carbon atom, and explain your reasoning.

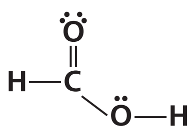
The first carbon (bonded to three H atoms and one C atom) atom is a sp^3 hybrid because it has 4 bonding positions. The 2nd carbon atom (bonded to one C atom and one N atom) is an sp hybrid because it has two bonding positions.

Challenge Problem

137. Examine the bond-dissociation energies for the various bonds listed in **Table 8.12.**

Bond-Dissociation Energies			
Bond	Bond-Dissociation Energy (kJ/mol)	Bond	Bond-Dissociation Energy (kJ/mol)
C–C	348	O–H	467
C=C	614	C–N	305
C≡C	839	O=O	498
N–N	163	C–H	416
N=N	418	C–O	358
N≡N	945	C=O	745

a. Draw the correct Lewis structures for C_2H_2 and HCOOH.



b. Determine the amount of energy needed to break apart each of these molecules.

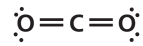
C_2H_2: (416 × 2) + 839 kJ/mol = 1671 kJ/mol

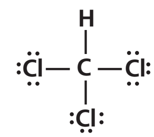
HCOOH: 416 + 745 + 358 + 467 kJ/mol = 1986 kJ/mol

Cumulative Review

138. Table 8.13 lists a liquid's mass and volume data. Create a line graph of this data with the volume on the *x*-axis and the mass on the *y*-axis. Calculate the slope of the line. What information does the slope give you? *(Chapter 2)*

Mass v. Volume	
Volume	Mass
4.1 mL	9.36 g
6.0 mL	14.04 g
8.0 mL	18.72 g
10.0 mL	23.40 g





$$\text{slope} = \frac{\Delta y}{\Delta x} = \frac{23.40\text{ g} - 14.04\text{ g}}{10.0\text{ mL} - 6.0\text{ mL}} = 2.34\text{ g/mL}$$

The unit of the slope is g/mL, which is the unit of density. The slope gives you the density of the liquid.

139. Write the correct chemical formula for each compound. *(Chapter 7)*

a. calcium carbonate

$CaCO_3$

b. potassium chlorate

$KClO_3$

c. silver acetate

$AgC_2H_3O_2$

d. copper(II) sulfate

$CuSO_4$

e. ammonium phosphate

$(NH_4)_3PO_4$

140. Write the correct chemical name for each compound.

a. NaI

sodium iodide

b. $Fe(NO_3)_3$

iron(III) nitrate

c. $Sr(OH)_2$

strontium hydroxide

d. $CoCl_2$

cobalt(II) chloride

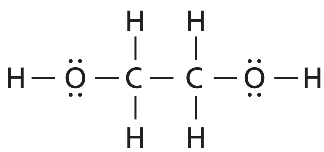
e. $Mg(BrO_3)_2$

magnesium bromate

Additional Assessment

Writing in Chemistry

141. Antifreeze Research ethylene glycol, an antifreeze-coolant, to learn its chemical formula. Explain how its structure makes it a useful antifreeze and coolant.



Answers will vary. Students might note that the presence of –OH groups make ethylene glycol miscible in water and contribute to its relatively high boiling point and relatively low freezing point.

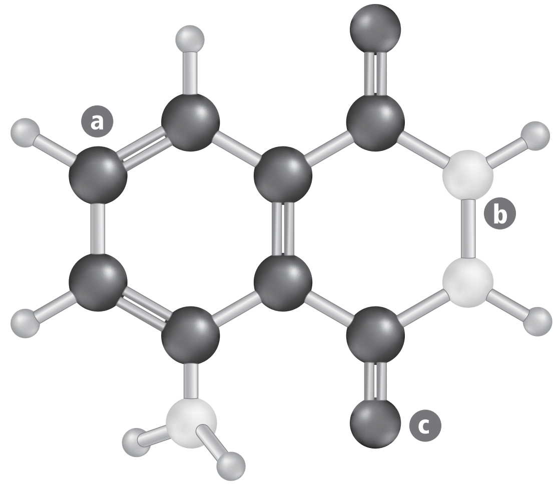
142. Detergents Choose a laundry detergent to research and write an essay about its chemical composition. Explain how it removes oil and grease from of fabrics.

Answers should include a discussion of the nonpolar end of a detergent molecule and the polar end of the same molecule allowing it to attract both water and oil.

Document-Based Questions

Luminol Crime-scene investigators often use the covalent compound luminol to find blood evidence. The reaction between luminol, certain chemicals, and hemoglobin, a protein in blood, produces light. **Figure 8.26** shows a ball-and-stick model of luminol.

Data obtained from: Fleming, Declan., 2002. The Chemiluminescence of Luminol, Exemplarchem, Royal Society of Chemistry.



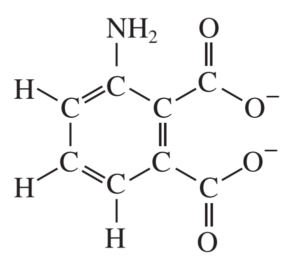
143. Determine the molecular formula for luminol and draw its Lewis structure.

$C_8H_7O_2N_3$

144. Indicate the hybrid present on the atoms labeled *A, B,* and *C* in **Figure 8.26.**

a, sp^2; b, sp^3; c, sp^2

145. When luminol comes in contact with the iron ion in hemoglobin, it reacts to produce Na_2APA, water, nitrogen, and light energy. Given the structural formula of the APA ion in **Figure 8.27,** write the chemical formula for the polyatomic APA ion.



APA ion

$C_8H_5NO_4{}^{2-}$

Standardized Test Practice

pages 278–279

1. The common name of SiI_4 is tetraiodosilane. What is its molecular compound name?

a. silane tetraiodide
b. silane tetraiodine
c. silicon iodide
d. silicon tetraiodide

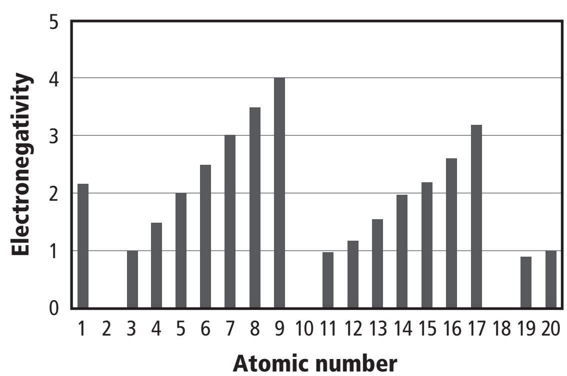
d

2. Which compound contains at least one pi bond?

a. CO_2
b. $CHCl_3$
c. AsI_3
d. BeF_2

a

Use the graph below to answer Questions 3 and 4.



3. What is the electronegativity of the element with atomic number 14?

a. 1.5
b. 1.9
c. 2.0
d. 2.2

c

4. An ionic bond would form between which pairs of elements?

a. atomic number 3 and atomic number 4
b. atomic number 7 and atomic number 8
c. atomic number 4 and atomic number 18
d. atomic number 8 and atomic number 12

d

5. Which is the Lewis structure for silicon disulfide?

a. 

b. 

c. 

d. 

b

6. The central selenium atom in selenium hexafluoride forms an expanded octet. How many electron pairs surround the central Se atom?

a. 4
b. 5
c. 6
d. 7

c

Use the table below to answer Questions 7 and 8.

Bond Dissociation Energies at 298 K

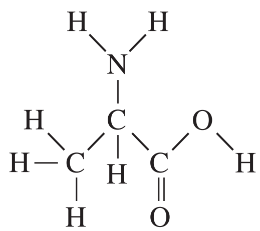
Bond	kJ/mol	Bond	kJ/mol
Cl—Cl	242	N≡N	945
C—C	345	O—H	467
C—H	416	C—O	358
C—N	305	C=O	745
H—I	299	O=O	498
H—N	391		

7. Which diatomic gas has the shortest bond between its two atoms?

a. HI
b. O_2
c. Cl_2
d. N_2

d

8. Approximately how much energy will it take to break all the bonds present in the molecule below?



a. 3024 kJ/mol
b. 4318 kJ/mol
c. 4621 kJ/mol
d. 5011 kJ/mol

d

E_{total} = (2 × E_{HN}) + E_{CN} + (4 × E_{CH}) + (2 × E_{CC}) + EC = O + E_{CO} + E_{OH} E_{total} = (2 × 391) + 305 + (4 × 416) + (2 × 345) + 745 + 358 + 467 kJ/mol
E_{total} = 5011 kJ/mol

9. Which compound does NOT have a bent molecular shape?

a. BeH_2
b. H_2S
c. H_2O
d. SeH_2

a

10. Which compound is nonpolar?

a. H_2S
b. CCl_4
c. SiH_3Cl
d. AsH_3

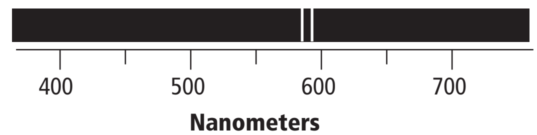
b

11. Oxyacids contain hydrogen and an oxyanion. There are two different oxyacids that contain hydrogen, nitrogen, and oxygen. Identify these two oxyacids. How can they be distinguished on the basis of their names and formulas?

Nitric acid (HNO_3) and nitrous acid (NHO_2)

The –ic suffix indicates the larger number of oxygen atoms; the –ous suffix indicates the lower number of oxygen atoms.

Use the atomic emissions spectrum below to answer Questions 12 and 13.



12. Estimate the wavelength of the photons being emitted by this element.

580 nm

13. Find the frequency of the photons being emitted by this element.

5.2×10^{14} Hz

Solution: $c = \lambda\nu$
$\nu = c/\lambda$

$$\nu = \frac{3 \times 10^8\ \cancel{m}/s}{580 \times 10^{-9}\ \cancel{m}}$$

$= 5.2 \times 10^{14}\ s^{-1}$ or Hz

Use the table below to answer Question 14.

Percent Abundance of Silicon Isotopes

Isotope	Mass	Percent Abundance
^{28}Si	27.98 amu	92.21%
^{29}Si	28.98 amu	4.70%
^{30}Si	29.97 amu	3.09%

14. Your lab partner calculates the average atomic mass of these three silicon isotopes. His average atomic mass value is 28.98 amu. Explain why your lab partner is incorrect, and show how to calculate the correct average atomic mass.

The correct calculation uses a weighted average to account for the proportions of each isotope in the sample:

$M_{avg} = 0.9221 \times M_{^{28}Si} + 0.0470 \times M_{^{29}Si} + 0.0309 \times M_{^{30}Si}$ amu

$M_{avg} = 0.9221 \times 27.98 + 0.0470 \times 28.98 + 0.0309 \times 29.97$ amu

M_{avg} = 28.09 amu

SAT Subject Test: Chemistry

Use the list of separation techniques below to answer Questions 15 through 17.

a. filtration
b. distillation
c. crystallization
d. chromatography
e. sublimation

15. Which technique separates components of a mixture with different boiling points?

b

16. Which technique separates components of a mixture based on the size of its particles?

a

17. Which technique is based on the stronger attraction some components have for the stationary phase compared to the mobile phase?

d

Use the table below to answer Questions 18 and 19.

Electron-Dot Structures

Group	1	2	13	14	15	16	17	18
Diagram	Li·	·Be·	·B·	·C·	·N·	·O:	:F:	:Ne:

18. Based on the Lewis structures shown, which elements will combine in a 2:3 ratio?

a. lithium and carbon
b. beryllium and fluorine
c. beryllium and nitrogen
d. boron and oxygen
e. boron and carbon

d

19. How many electrons will beryllium have in its outer energy level after it forms an ion to become chemically stable?

a. 0
b. 2
c. 4
d. 6
e. 8

a