Continued—

Q14. Use bond energies from Table 10.3 to determine ΔH_{rxn} for the reaction between ethanol and hydrogen chloride.
MISSED THIS? *Read Section 10.10; Watch IWE 10.11*

$$CH_3CH_2OH(g) + HCl(g) \longrightarrow CH_3CH_2Cl(g) + H_2O(g)$$

a) −1549 kJ b) 1549 kJ c) −12 kJ d) 12 kJ

Q15. Consider the halogenation of ethene, where X is a generic halogen:

$$H_2C{=}CH_2(g) + X_2(g) \longrightarrow H_2XC{-}CH_2X(g)$$

Use bond energies to determine which halogen produces the most exothermic halogenation reaction with ethene. The C—F, C—Br, and C—I bond energies are 552 kJ/mol, 280 kJ/mol, and 209 kJ/mol, respectively. Look up all other necessary bond energies in Table 10.3.
MISSED THIS? *Read Section 10.10; Watch IWE 10.11*

a) fluorine b) chlorine
c) bromine d) iodine

Answers: **1.** (d) **2.** (b) **3.** (c) **4.** (d) **5.** (a) **6.** (a) **7.** (c) **8.** (a) **9.** (c) **10.** (b) **11.** (a) **12.** (d) **13.** (b) **14.** (c) **15.** (a)

CHAPTER 10 IN REVIEW

TERMS

Section 10.1
Lewis model (394)
Lewis electron-dot structures (Lewis structures) (394)

Section 10.2
ionic bond (395)
covalent bond (395)
metallic bonding (396)

Section 10.3
Lewis symbol (396)
octet (396)
duet (396)
chemical bond (396)
octet rule (396)

Section 10.4
lattice energy (398)
Born–Haber cycle (398)

Section 10.5
bonding pair (404)
lone pair (404)
nonbonding electrons (404)
double bond (404)
triple bond (405)

Section 10.6
polar covalent bond (407)
electronegativity (407)
dipole moment (μ) (408)
percent ionic character (409)

Section 10.8
resonance structures (413)
resonance hybrid (413)
formal charge (414)

Section 10.9
free radical (418)

Section 10.10
bond energy (422)
bond length (424)

CONCEPTS

Bonding Models and AIDS Drugs (10.1)

- Theories that predict how and why atoms bond together are central to chemistry because they explain compound stability and molecule shape.
- Bonding theories have been useful in combating HIV because they help in the design of molecules that bind to the active site of a protein crucial for the development of AIDS.

Types of Chemical Bonds (10.2)

- We divide chemical bonds into three general types: ionic bonds, which occur between a metal and a nonmetal; covalent bonds, which occur between two nonmetals; and metallic bonds, which occur within metals.
- In an ionic bond, an electron transfers from the metal to the nonmetal, and the resultant ions attract each other by coulombic forces.
- In a covalent bond, nonmetals share electrons that interact with the nuclei of both atoms via coulombic forces, holding the atoms together.
- In a metallic bond, the atoms form a lattice in which each metal loses electrons to an "electron sea." The attraction of the positively charged metal ions to the electron sea holds the metal together.

The Lewis Model and Electron Dots (10.3)

- In the Lewis model, chemical bonds form when atoms transfer (ionic bonding) or share (covalent bonding) valence electrons to attain noble gas electron configurations.
- The Lewis model represents valence electrons as dots surrounding the symbol for an element. When two or more elements bond together, the dots are transferred or shared so that every atom has eight dots, an octet (or two dots, a duet, in the case of hydrogen).

Ionic Lewis Structures and Lattice Energy (10.4)

- In an ionic Lewis structure involving main-group metals, the metal transfers its valence electrons (dots) to the nonmetal.
- The formation of most ionic compounds is exothermic because of lattice energy, the energy released when metal cations and nonmetal anions coalesce to form the solid. The smaller the radius of the ions and the greater their charge, the more exothermic the lattice energy.

Covalent Lewis Structures, Electronegativity, and Polarity (10.5, 10.6, 10.7)

- In a covalent Lewis structure, neighboring atoms share valence electrons to attain octets (or duets).

- A single shared electron pair constitutes a single bond, while two or three shared pairs constitute double or triple bonds, respectively.
- The shared electrons in a covalent bond are not always *equally* shared; when two dissimilar nonmetals form a covalent bond, the electron density is greater on the more electronegative element. The result is a polar bond, in which one element has a partial positive charge and the other a partial negative charge.
- Electronegativity—the ability of an atom to attract electrons to itself in chemical bonding—increases as we move to the right across a period in the periodic table and decreases as we move down a column.
- Elements with very dissimilar electronegativities form ionic bonds, those with very similar electronegativities form nonpolar covalent bonds, and those with intermediate electronegativity differences form polar covalent bonds.

Resonance and Formal Charge (10.8)

- We can best represent some molecules not by a single Lewis structure, but by two or more resonance structures. The actual structure of these molecules is a resonance hybrid: a combination or average of the contributing structures.
- The formal charge of an atom in a Lewis structure is the charge the atom would have if all bonding electrons were shared equally between bonding atoms.
- In general, the best Lewis structures have the fewest atoms with formal charge, and any negative formal charges are on the most electronegative atom.

Exceptions to the Octet Rule (10.9)

- Although the octet rule normally applies when we draw Lewis structures, some exceptions occur.
- These exceptions include odd-electron species, which necessarily have Lewis structures with only seven electrons around an atom. Such molecules, called free radicals, tend to be unstable and chemically reactive.
- Other exceptions to the octet rule include molecules with incomplete octets—usually totaling six electrons (especially important in compounds containing boron)—and molecules with expanded octets—usually 10 or 12 electrons (which can occur in compounds containing elements from the third row of the periodic table and below). Expanded octets never occur in second-period elements.

Bond Energies and Bond Lengths (10.10)

- The bond energy of a chemical bond is the energy required to break one mole of the bonds in the gas phase.
- Average bond energies for a number of different bonds are tabulated, and we can use them to calculate enthalpies of reaction.
- Average bond lengths are also tabulated.
- In general, triple bonds are shorter and stronger than double bonds, which are in turn shorter and stronger than single bonds.

Bonding in Metals (10.11)

- When metal atoms bond together to form a solid, each metal atom donates one or more electrons to an *electron sea*. The metal cations are held together by their attraction to the sea of electrons.
- The *electron sea* model accounts for the electrical conductivity, thermal conductivity, malleability, and ductility of metals.

EQUATIONS AND RELATIONSHIPS

Dipole Moment (μ): Separation of Two Particles of Equal but Opposite Charge of Magnitude q by a Distance r (10.6)

$$\mu = qr$$

Percent Ionic Character (10.6)

$$\text{Percent ionic character} = \frac{\text{measured dipole moment of bond}}{\text{dipole moment if electron were completely transferred}} \times 100\%$$

Formal Charge (10.8)

$$\text{Formal charge} = \text{number of valence electrons} - (\text{number of nonbonding electrons} + \tfrac{1}{2}\,\text{number of shared electrons})$$

Enthalpy Change of a Reaction (ΔH_{rxn}): Relationship of Bond Energies (10.10)

$$\Delta H_{rxn} = \Sigma(\Delta H\text{'s bonds broken}) + \Sigma(\Delta H\text{'s bonds formed})$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Write Lewis symbols for main group elements (10.3)	**Exercises** 35–38
Write and use Lewis symbols to describe and analyze ionic compounds (10.4)	Example 10.1 **For Practice** 10.1 **Exercise** 39–42
Compare the relative lattice energies of different compounds (10.4)	Example 10.2 **For Practice** 10.2 **For More Practice** 10.2 **Exercises** 43–46
Calculate lattice energies of ionic solids (10.4)	**Exercises** 47–48
Classify bonds as pure covalent, polar covalent, or ionic (10.5–10.6)	Example 10.3 **For Practice** 10.3 **Exercises** 55–58
Write and use Lewis structures to describe and analyze molecular compounds and polyatomic ions (10.7)	Examples 10.4–10.7 **For Practice** 10.4–10.7 **Exercises** 49–54, 59–62

—Continued on the next page

Continued—

Write the best possible Lewis structure using formal charge (10.8)	Examples 10.8, 10.9 **For Practice** 10.8, 10.9 **For More Practice** 10.8 **Exercises** 63–72
Write Lewis structures for molecules and ions that are exceptions to the octet rule (10.9)	Example 10.10 **For Practice** 10.10 **For More Practice** 10.10 **Exercises** 73–78
Analyze bonds in terms of bond energy and bond length (10.10)	Example 10.11 **For Practice** 10.11 **For More Practice** 10.11 **Exercises** 79–84

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. Why are bonding theories important? Provide some examples of what bonding theories can predict.

2. Why do chemical bonds form? What basic forces are involved in bonding?

3. What are the three basic types of chemical bonds? What happens to electrons in the bonding atoms in each type?

4. How do you determine how many dots to put around the Lewis symbol of an element?

5. Describe the octet rule in the Lewis model.

6. According to the Lewis model, what is a chemical bond?

7. How do you draw an ionic Lewis structure?

8. How can Lewis structures be used to determine the formula of ionic compounds? Give an example.

9. What is lattice energy?

10. Why is the formation of solid sodium chloride from solid sodium and gaseous chlorine exothermic, even though it takes more energy to form the Na^+ ion than the amount of energy released upon formation of Cl^-?

11. What is the Born–Haber cycle? List each step in the cycle and show how the cycle is used to calculate lattice energy.

12. How does lattice energy relate to ionic radii? To ion charge?

13. How does the ionic bonding model explain the relatively high melting points of ionic compounds?

14. How does the ionic bonding model explain the nonconductivity of ionic solids, and at the same time the conductivity of ionic solutions?

15. In a covalent Lewis structure, what is the difference between lone pair and bonding pair electrons?

16. In what ways are double and triple covalent bonds different from single covalent bonds?

17. How does the Lewis model for covalent bonding account for why certain combinations of atoms are stable while others are not?

18. How does the Lewis model for covalent bonding account for the relatively low melting and boiling points of molecular compounds (compared to ionic compounds)?

19. What is electronegativity? What are the periodic trends in electronegativity?

20. How do a pure covalent bond, a polar covalent bond, and an ionic bond differ?

21. Explain percent ionic character of a bond. Do any bonds have 100% ionic character?

22. What is a dipole moment?

23. What is the magnitude of the dipole moment formed by separating a proton and an electron by 100 pm? 200 pm?

24. What is the basic procedure for writing a covalent Lewis structure?

25. How do you determine the number of electrons in the Lewis structure of a molecule? A polyatomic ion?

26. What are resonance structures? What is a resonance hybrid?

27. Do resonance structures always contribute equally to the overall structure of a molecule? Explain.

28. What is formal charge? How is formal charge calculated? How is it helpful?

29. Why does the octet rule have exceptions? List the three major categories of exceptions and an example of each.

30. Which elements can have expanded octets? Which elements should never have expanded octets?

31. What is bond energy? How can you use average bond energies to calculate enthalpies of reaction?

32. Explain the difference between endothermic reactions and exothermic reactions with respect to the bond energies of the bonds broken and formed.

33. What is the electron sea model for bonding in metals?

34. How does the electron sea model explain the conductivity of metals? The malleability and ductility of metals?

PROBLEMS BY TOPIC

Valence Electrons and Dot Structures

35. Write the electron configuration for N. Then write the Lewis symbol for N and show which electrons from the electron configuration are included in the Lewis symbol.
MISSED THIS? *Read Section 10.3*

36. Write the electron configuration for Ne. Then write the Lewis symbol for Ne and show which electrons from the electron configuration are included in the Lewis symbol.

37. Write the Lewis symbol for each atom or ion.
MISSED THIS? *Read Section 10.3*
a. Al **b.** Na^+ **c.** Cl **d.** Cl^-

38. Write the Lewis symbol for each atom or ion.
a. S^{2-} **b.** Mg **c.** Mg^{2+} **d.** P

Ionic Lewis Symbols and Lattice Energy

39. Write the Lewis symbols for the ions in each ionic compound.
MISSED THIS? *Read Section 10.4*
a. NaF **b.** CaO **c.** $SrBr_2$ **d.** K_2O

40. Write the Lewis symbols for the ions in each ionic compound.
a. SrO **b.** Li_2S **c.** CaI_2 **d.** RbF

41. Use Lewis symbols to determine the formula for the compound that forms between each pair of elements.
MISSED THIS? *Read Section 10.4*
a. Sr and Se **b.** Ba and Cl **c.** Na and S **d.** Al and O

42. Use Lewis symbols to determine the formula for the compound that forms between each pair of elements:
a. Ca and N **b.** Mg and I **c.** Ca and S **d.** Cs and F

43. Explain the trend in the lattice energies of the alkaline earth metal oxides. **MISSED THIS?** *Read Section 10.4*

Metal Oxide	Lattice Energy (kJ/mol)
MgO	−3795
CaO	−3414
SrO	−3217
BaO	−3029

44. Rubidium iodide has a lattice energy of −617 kJ/mol, while potassium bromide has a lattice energy of −671 kJ/mol. Why is the lattice energy of potassium bromide more exothermic than the lattice energy of rubidium iodide?

45. The lattice energy of CsF is −744 kJ/mol, whereas that of BaO is −3029 kJ/mol. Explain this large difference in lattice energy.
MISSED THIS? *Read Section 10.4*

46. Arrange these compounds in order of increasing magnitude of lattice energy: KCl, SrO, RbBr, CaO.

47. Use the Born–Haber cycle and data from Appendix IIB, Chapter 9, and this chapter to calculate the lattice energy of KCl. (ΔH_{sub} for potassium is 89.0 kJ/mol.) **MISSED THIS?** *Read Section 10.4*

48. Use the Born–Haber cycle and data from Appendix IIB and Table 10.3 to calculate the lattice energy of CaO. (ΔH_{sub} for calcium is 178 kJ/mol; IE_1 and IE_2 for calcium are 590 kJ/mol and 1145 kJ/mol, respectively; EA_1 and EA_2 for O are −141 kJ/mol and 744 kJ/mol, respectively.)

Simple Covalent Lewis Structures, Electronegativity, and Bond Polarity

49. Use covalent Lewis structures to explain why each element (or family of elements) occurs as diatomic molecules.
MISSED THIS? *Read Section 10.5; Watch KCV 10.5*
a. hydrogen
b. the halogens
c. oxygen
d. nitrogen

50. Use covalent Lewis structures to explain why the compound that forms between nitrogen and hydrogen has the formula NH_3. Show why NH_2 and NH_4 are not stable.

51. Write the Lewis structure for each molecule.
MISSED THIS? *Read Section 10.7; Watch KCV 10.7, IWE 10.4*
a. PH_3 **b.** SCl_2 **c.** HI **d.** CH_4

52. Write the Lewis structure for each molecule.
a. NF_3 **b.** HBr **c.** SBr_2 **d.** CCl_4

53. Write the Lewis structure for each molecule.
MISSED THIS? *Read Section 10.7; Watch KCV 10.7, IWE 10.4*
a. SF_2 **b.** SiH_4
c. HCOOH (both O bonded to C) **d.** CH_3SH (C and S central)

54. Write the Lewis structure for each molecule.
a. CH_2O **b.** C_2Cl_4
c. CH_3NH_2 **d.** $CFCl_3$ (C central)

55. Determine if a bond between each pair of atoms would be pure covalent, polar covalent, or ionic.
MISSED THIS? *Read Section 10.6; Watch KCV 10.6*
a. Br and Br **b.** C and Cl **c.** C and S **d.** Sr and O

56. Determine if a bond between each pair of atoms would be pure covalent, polar covalent, or ionic.
a. C and N **b.** N and S **c.** K and F **d.** N and N

57. Draw the Lewis structure for CO with an arrow representing the dipole moment. Refer to Figure 10.10 to estimate the percent ionic character of the CO bond.
MISSED THIS? *Read Section 10.6; Watch KCV 10.6*

58. Draw the Lewis structure for BrF with an arrow representing the dipole moment. Refer to Figure 10.10 to estimate the percent ionic character of the BrF bond.

Covalent Lewis Structures, Resonance, and Formal Charge

59. Write the Lewis structure for each molecule or ion.
MISSED THIS? *Read Section 10.7; Watch KCV 10.7, IWE 10.4*
a. CI_4 **b.** N_2O **c.** SiH_4 **d.** Cl_2CO

60. Write the Lewis structure for each molecule or ion.
a. H_3COH **b.** OH^- **c.** BrO^- **d.** O_2^{2-}

61. Write the Lewis structure for each molecule or ion.
MISSED THIS? *Read Section 10.7; Watch KCV 10.7, IWE 10.4*
a. N_2H_2 **b.** N_2H_4 **c.** C_2H_2 **d.** C_2H_4

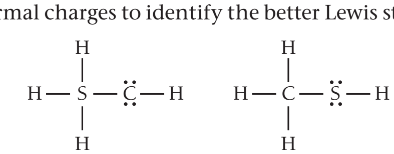
62. Write the Lewis structure for each molecule or ion.
a. H_3COCH_3 **b.** CN^- **c.** NO_2^- **d.** ClO^-

63. Write a Lewis structure that obeys the octet rule for each molecule or ion. Include resonance structures if necessary and assign formal charges to each atom.
MISSED THIS? *Read Section 10.8; Watch KCV 10.8, IWE 10.7*
a. SeO_2 **b.** CO_3^{2-} **c.** ClO^- **d.** NO_2^-

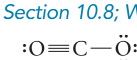
64. Write a Lewis structure that obeys the octet rule for each ion. Include resonance structures if necessary and assign formal charges to each atom.
a. ClO_3^- **b.** ClO_4^- **c.** NO_3^- **d.** NH_4^+

65. Use formal charge to identify the better Lewis structure.
MISSED THIS? *Read Section 10.8; Watch KCV 10.8, IWE 10.8*

66. Use formal charges to identify the better Lewis structure.



67. How important is the resonance structure shown here to the overall structure of carbon dioxide? Explain.
MISSED THIS? *Read Section 10.8; Watch KCV 10.8, IWE 10.8*

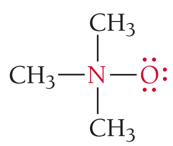


68. In N_2O, nitrogen is the central atom, and the oxygen atom is terminal. In OF_2, however, oxygen is the central atom. Use formal charges to explain why.

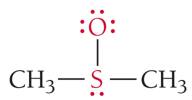
69. Draw the Lewis structure (including resonance structures) for the acetate ion (CH_3COO^-). For each resonance structure, assign formal charges to all atoms that have formal charge. **MISSED THIS?** *Read Section 10.8; Watch KCV 10.8, IWE 10.9*

70. Draw the Lewis structure (including resonance structures) for methyl azide (CH_3N_3). For each resonance structure, assign formal charges to all atoms that have formal charge.

71. What are the formal charges of the atoms shown in red? **MISSED THIS?** *Read Section 10.8; Watch KCV 10.8, IWE 10.9*



72. What are the formal charges of the atoms shown in red?



Odd-Electron Species, Incomplete Octets, and Expanded Octets

73. Write the Lewis structure for each molecule (octet rule not followed). **MISSED THIS?** *Read Section 10.9; Watch KCV 10.9*

a. BCl_3 **b.** NO_2 **c.** BH_3

74. Write the Lewis structure for each molecule (octet rule not followed).

a. BBr_3
b. NO
c. ClO_2

75. Write the Lewis structure for each ion. Include resonance structures if necessary and assign formal charges to all atoms. If necessary, expand the octet on the central atom to lower formal charge. **MISSED THIS?** *Read Section 10.9; Watch KCV 10.9, IWE 10.10*

a. PO_4^{3-} **b.** CN^- **c.** SO_3^{2-} **d.** ClO_2^-

76. Write Lewis structures for each molecule or ion. Include resonance structures if necessary and assign formal charges to all atoms. If necessary, expand the octet on the central atom to lower formal charge.

a. SO_4^{2-} **b.** HSO_4^- **c.** SO_3 **d.** BrO_2^-

77. Write Lewis structures for each molecule or ion. Use expanded octets as necessary.
MISSED THIS? *Read Section 10.9; Watch KCV 10.9, IWE 10.10*

a. PF_5 **b.** I_3^-
c. SF_4 **d.** GeF_4

78. Write Lewis structures for each molecule or ion. Use expanded octets as necessary.

a. ClF_5 **b.** AsF_6^-
c. Cl_3PO **d.** IF_5

Bond Energies and Bond Lengths

79. List these compounds in order of increasing carbon–carbon bond strength and in order of decreasing carbon–carbon bond length: HCCH, H_2CCH_2, H_3CCH_3. **MISSED THIS?** *Read Section 10.10*

80. Which compound shown here has the stronger nitrogen–nitrogen bond? The shorter nitrogen–nitrogen bond?

$$H_2NNH_2, HNNH$$

81. Hydrogenation reactions are used to add hydrogen across double bonds in hydrocarbons and other organic compounds. Use average bond energies to calculate ΔH_{rxn} for the hydrogenation reaction. **MISSED THIS?** *Read Section 10.10; Watch IWE 10.11*

$$H_2C=CH_2(g) + H_2(g) \longrightarrow H_3C-CH_3(g)$$

82. Ethanol is a possible fuel. Use average bond energies to calculate ΔH_{rxn} for the combustion of ethanol.

$$CH_3CH_2OH(g) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$$

83. Ethane burns in air to form carbon dixode and water vapor. **MISSED THIS?** *Read Section 10.10; Watch IWE 10.11*

$$2\,H_3C-CH_3(g) + 7\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$$

Use average bond energies to calculate ΔH_{rxn} for the reaction.

84. In the *Chemistry and the Environment* box on free radicals in this chapter, we discussed the importance of the hydroxyl radical in reacting with and eliminating many atmospheric pollutants. However, the hydroxyl radical does not clean up everything. For example, chlorofluorocarbons—which destroy stratospheric ozone—are not attacked by the hydroxyl radical. Consider the hypothetical reaction by which the hydroxyl radical might react with a chlorofluorocarbon:

$$OH(g) + CF_2Cl_2(g) \longrightarrow HOF(g) + CFCl_2(g)$$

Use bond energies to explain why this reaction is improbable. (The C—F bond energy is 552 kJ/mol.)

CUMULATIVE PROBLEMS

85. Write an appropriate Lewis structure for each compound. Make certain to distinguish between ionic and molecular compounds.

a. BI_3 **b.** K_2S **c.** HCFO **d.** PBr_3

86. Write an appropriate Lewis structure for each compound. Make certain to distinguish between ionic and molecular compounds.

a. Al_2O_3 **b.** ClF_5 **c.** MgI_2 **d.** XeO_4

87. Each compound contains both ionic and covalent bonds. Write ionic Lewis structures for each of them, including the covalent structure for the ion in brackets. Write resonance structures if necessary.

a. $BaCO_3$ **b.** $Ca(OH)_2$
c. KNO_3 **d.** LiIO

88. Each compound contains both ionic and covalent bonds. Write ionic Lewis structures for each of them, including the covalent structure for the ion in brackets. Write resonance structures if necessary.

a. $RbIO_2$ **b.** NH_4Cl **c.** KOH **d.** $Sr(CN)_2$

89. Carbon ring structures are common in organic chemistry. Draw a Lewis structure for each carbon ring structure, including any necessary resonance structures.

a. C_4H_8 **b.** C_4H_4 **c.** C_6H_{12} **d.** C_6H_6

90. Amino acids are the building blocks of proteins. The simplest amino acid is glycine (H_2NCH_2COOH). Draw a Lewis structure for glycine. (*Hint:* The central atoms in the skeletal structure are nitrogen and the two carbon atoms. Each oxygen atom is bonded directly to the right-most carbon atom.)

91. Formic acid is responsible for the sting of ant bites. By mass, formic acid is 26.10% C, 4.38% H, and 69.52% O. The molar mass of formic acid is 46.02 g/mol. Determine the molecular formula of formic acid and draw its Lewis structure.

92. Diazomethane is a highly poisonous, explosive compound because it readily evolves N_2. Diazomethane has the following composition by mass: 28.57% C; 4.80% H; and 66.64% N. The molar mass of diazomethane is 42.04 g/mol. Find the molecular formula of diazomethane, draw its Lewis structure, and assign formal charges to each atom. Why is diazomethane not very stable? Explain.

93. The reaction of $Fe_2O_3(s)$ with $Al(s)$ to form $Al_2O_3(s)$ and $Fe(s)$ is called the thermite reaction and is highly exothermic. What role does lattice energy play in the exothermicity of the reaction?

94. NaCl has a lattice energy of −787 kJ/mol. Consider a hypothetical salt XY. X^{3+} has the same radius as Na^+ and Y^{3-} has the same radius as Cl^-. Estimate the lattice energy of XY.

95. Draw the Lewis structure for nitric acid (the hydrogen atom is attached to one of the oxygen atoms). Include all three resonance structures by alternating the double bond among the three oxygen atoms. Use formal charge to determine which of the resonance structures is most important to the structure of nitric acid.

96. Phosgene (Cl_2CO) is a poisonous gas used as a chemical weapon during World War I. It is a potential agent for chemical terrorism today. Draw the Lewis structure of phosgene. Include all three resonance forms by alternating the double bond among the three terminal atoms. Which resonance structure is the best?

97. The cyanate ion (OCN^-) and the fulminate ion (CNO^-) share the same three atoms but have vastly different properties. The cyanate ion is stable, while the fulminate ion is unstable and forms explosive compounds. The resonance structures of the cyanate ion are explored in Example 10.8. Draw Lewis structures for the fulminate ion—including possible resonance forms—and use formal charge to explain why the fulminate ion is less stable (and therefore more reactive) than the cyanate ion.

98. Draw the Lewis structure for each organic compound from its condensed structural formula.
a. C_3H_8 **b.** CH_3OCH_3 **c.** CH_3COCH_3
d. CH_3COOH **e.** CH_3CHO

99. Draw the Lewis structure for each organic compound from its condensed structural formula.
a. C_2H_4 **b.** CH_3NH_2 **c.** HCHO
d. CH_3CH_2OH **e.** HCOOH

100. Use Lewis structures to explain why Br_3^- and I_3^- are stable, while F_3^- is not.

101. Draw the Lewis structure for $HCSNH_2$. (The carbon and nitrogen atoms are bonded together, and the sulfur atom is bonded to the carbon atom.) Label each bond in the molecule as polar or nonpolar.

102. Draw the Lewis structure for urea, H_2NCONH_2, one of the compounds responsible for the smell of urine. (The central carbon atom is bonded to both nitrogen atoms and to the oxygen atom.) Does urea contain polar bonds? Which bond in urea is most polar?

103. Some theories of aging suggest that free radicals cause certain diseases and perhaps aging in general. As you know from the Lewis model, such molecules are not chemically stable and will quickly react with other molecules. According to certain theories, free radicals may attack molecules within the cell, such as DNA, changing them and causing cancer or other diseases. Free radicals may also attack molecules on the surfaces of cells, making them appear foreign to the body's immune system. The immune system then attacks the cells and destroys them, weakening the body. Draw Lewis structures for each free radical implicated in this theory of aging.
a. O_2^- **b.** O^- **c.** OH
d. CH_3OO (unpaired electron on terminal oxygen)

104. Free radicals are important in many environmentally significant reactions (see the *Chemistry in the Environment* box on free radicals in this chapter). For example, photochemical smog—smog that results from the action of sunlight on air pollutants—forms in part by these two steps:

$$NO_2 \xrightarrow{\text{UV light}} NO + O$$
$$O + O_2 \longrightarrow O_3$$

The product of this reaction, ozone, is a pollutant in the lower atmosphere. (Upper atmospheric ozone is a natural part of the atmosphere that protects life on Earth from ultraviolet light.) Ozone is an eye and lung irritant and also accelerates the weathering of rubber products. Rewrite the given reactions using the Lewis structure of each reactant and product. Identify the free radicals.

105. If hydrogen were used as a fuel, it could be burned according to this reaction:

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)$$

Use average bond energies to calculate ΔH_{rxn} for this reaction and also for the combustion of methane (CH_4). Which fuel yields more energy per mole? Per gram?

106. Calculate ΔH_{rxn} for the combustion of octane (C_8H_{18}), a component of gasoline, by using average bond energies and then calculate it using enthalpies of formation from Appendix IIB. What is the percent difference between your results? Which result would you expect to be more accurate?

107. Draw the Lewis structure for each compound.
a. Cl_2O_7 (no Cl—Cl bond)
b. H_3PO_3 (two OH bonds)
c. H_3AsO_4

108. The azide ion, N_3^-, is a symmetrical ion, all of whose contributing resonance structures have formal charges. Draw three important contributing structures for this ion.

109. List the following gas-phase ion pairs in order of the quantity of energy released when they form from separated gas-phase ions. List the pair that releases the least energy first. Na^+F^-, $Mg^{2+}F^-$, Na^+O^{2-}, $Mg^{2+}O^{2-}$, $Al^{3+}O^{2-}$.

110. Calculate ΔH° for the reaction $H_2(g) + Br_2(g) \longrightarrow 2\,HBr(g)$ using the bond energy values. The ΔH_f° of $HBr(g)$ is not equal to one-half of the value calculated. Account for the difference.

111. The heat of atomization is the heat required to convert a molecule in the gas phase into its constituent atoms in the gas phase. The heat of atomization is used to calculate average bond energies. Without using any tabulated bond energies, calculate the average C—Cl bond energy from the following data: the heat of atomization of CH_4 is 1660 kJ/mol, and the heat of atomization of CH_2Cl_2 is 1495 kJ/mol.

112. Calculate the heat of atomization (see previous problem) of C_2H_3Cl, using the average bond energies in Table 10.3.

113. A compound composed of only carbon and hydrogen is 7.743% hydrogen by mass. Propose a Lewis structure for the compound.

114. A compound composed of only carbon and chlorine is 85.5% chlorine by mass. Propose a Lewis structure for the compound.

CHALLENGE PROBLEMS

115. The main component of acid rain (H_2SO_4) forms from the SO_2 pollutant in the atmosphere via these steps:

$$SO_2 + OH\cdot \longrightarrow HSO_3\cdot$$
$$HSO_3\cdot + O_2 \longrightarrow SO_3 + HOO\cdot$$
$$SO_3 + H_2O \longrightarrow H_2SO_4$$

Draw the Lewis structure for each of the species in these steps and use bond energies and Hess's law to estimate ΔH_{rxn} for the overall process. (Use 265 kJ/mol for the S—O single bond energy.)

116. A 0.167-g sample of an unknown acid requires 27.8 mL of 0.100 M NaOH to titrate to the equivalence point. Elemental analysis of the acid gives the following percentages by mass: 40.00% C, 6.71% H, 53.29% O. Determine the molecular formula, molar mass, and Lewis structure of the unknown acid.

117. Use the dipole moments of HF and HCl (given at the end of the problem) together with the percent ionic character of each bond (Figure 10.10) to estimate the bond length in each molecule. How well does your estimated bond length agree with the bond length in Table 10.4?

$$HCl\ \mu = 1.08\ D$$
$$HF\ \mu = 1.82\ D$$

118. Use average bond energies together with the standard enthalpy of formation of C(*g*) (718.4 kJ/mol) to estimate the standard enthalpy of formation of gaseous benzene, $C_6H_6(g)$. (Remember that average bond energies apply to the gas phase only.) Compare the value you obtain using average bond energies to the actual standard enthalpy of formation of gaseous benzene, 82.9 kJ/mol. What does the difference between these two values tell you about the stability of benzene?

119. The standard state of phosphorus at 25 °C is P_4. This molecule has four equivalent P atoms, no double or triple bonds, and no expanded octets. Draw its Lewis structure.

120. The standard heat of formation of $CaBr_2$ is −675 kJ/mol. The first ionization energy of Ca is 590 kJ/mol, and its second ionization energy is 1145 kJ/mol. The heat of sublimation of Ca[$Ca(s) \longrightarrow Ca(g)$] is 178 kJ/mol. The bond energy of Br_2 is 193 kJ/mol, the heat of vaporization of $Br_2(l)$ is 31 kJ/mol, and the electron affinity of Br is −325 kJ/mol. Calculate the lattice energy of $CaBr_2$.

121. The standard heat of formation of $PI_3(s)$ is −24.7 kJ/mol, and the PI bond energy in this molecule is 184 kJ/mol. The standard heat of formation of P(*g*) is 334 kJ/mol, and that of $I_2(g)$ is 62 kJ/mol. The I_2 bond energy is 151 kJ/mol. Calculate the heat of sublimation of PI_3[$PI_3(s) \longrightarrow PI_3(g)$].

122. A compound has the formula C_8H_8 and does not contain any double or triple bonds. All the carbon atoms are chemically identical, and all the hydrogen atoms are chemically identical. Draw the Lewis structure for this molecule.

123. Find the oxidation number of each sulfur in the molecule H_2S_4, which has a linear arrangement of its atoms.

124. Ionic solids of the O^- and O^{3-} anions do not exist, while ionic solids of the O^{2-} anion are common. Explain.

125. The standard state of sulfur is solid rhombic sulfur. Use the appropriate standard heats of formation given in Appendix II to find the average bond energy of the S=O in SO_2.

CONCEPTUAL PROBLEMS

126. Which statement is true of an endothermic reaction?

a. Strong bonds break and weak bonds form.
b. Weak bonds break and strong bonds form.
c. The bonds that break and those that form are of approximately the same strength.

127. When a firecracker explodes, energy is obviously released. The compounds in the firecracker can be viewed as being "energy rich." What does this mean? Explain the source of the energy in terms of chemical bonds.

128. A fundamental difference between compounds containing ionic bonds and those containing covalent bonds is the existence of molecules. Explain why molecules exist in solid covalent compounds but not in solid ionic compounds.

129. In the very first chapter of this book, we described the scientific approach and put a special emphasis on scientific models or theories. In this chapter, we looked carefully at a model for chemical bonding (the Lewis model). Why is this theory successful? What are some of the limitations of the theory?

QUESTIONS FOR GROUP WORK

Active Classroom Learning

Discuss these questions with the group and record your consensus answer.

130. Have each member of your group represent an atom of a metal or an atom of a nonmetal. Each group member holds a coin to represent an electron. Which group members are most reluctant to give up their electrons? Which group members are most willing to give up their electrons? Determine which kind of bond could form between each pair of group members. Tabulate your results.

131. Spend a few minutes reviewing the Lewis dot symbols for the atoms H through Ne. Form a circle and have each group member ask the group member on his or her right to draw the Lewis symbol for a specific atom. Keep going around until each group member can write all the Lewis dot symbols for the atoms H through Ne. Determine the formal charge for each symbol. In a complete sentence or two, describe why they are all the same.

132. Draw the Lewis dot symbols for the atoms Al and O. Use the Lewis model to determine the formula for the compound formed from these two atoms.

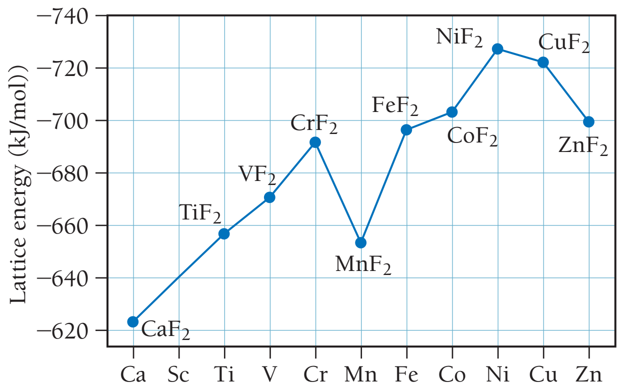
133. Draft a list of step-by-step instructions for writing the correct Lewis dot structure for any molecule or polyatomic ion.

134. Pass a piece of paper around the group and ask each group member in turn to perform the next step in the process of determining a correct Lewis structure (including formal charges on all atoms and resonance structures, if appropriate) for the following molecules and ions: N_2H_4, CCl_4, CO_3^{2-}, and NH_4^+.

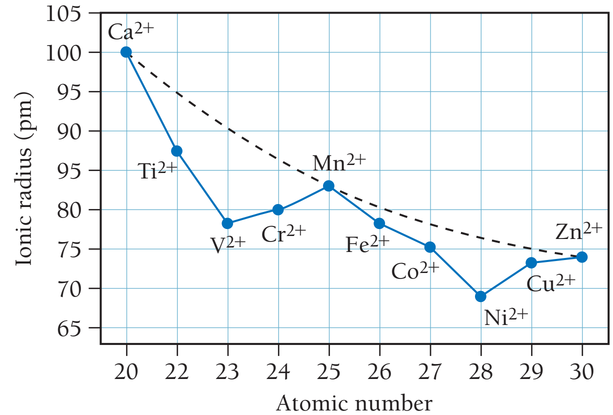
DATA INTERPRETATION AND ANALYSIS

Lattice Energy

135. Evidence for the additional stabilization of certain electron configurations comes from the experimental lattice energies of the metal fluorides, MF_2. The first figure below plots lattice energy for the 2+ metal cations for period 4 elements. The figure that follows plots the ionic radii of the 2+ metal cations of period 4 elements versus atomic number.



▲ **Lattice Energy of Period 4 Metal Fluorides**
Source: After C.S.G. Phillips, R. J. P Williams, Inorganic Chemistry, Volume 1, p. 179.



▲ **Ionic Radius Versus Atomic Number of 2+ Metal Cations for Period 4 Elements**

Use the information provided in the figures to answer the following questions:

a. Explain the general trend in lattice energy.

b. Is there a correlation between ionic radius and lattice energy? Explain.

c. What could account for the decrease in lattice energy between CrF_2 and MnF_2?

d. Which has the higher lattice energy: VF_2 or VCl_2? Explain.

Cc ANSWERS TO CONCEPTUAL CONNECTIONS

Bond Types

10.1 **(c)** MgF_2 contains a metal (Mg) bonded to a nonmetal (F), so the bonding is ionic.

Lewis Symbols

10.2 **(d)** This Lewis symbol has four dots, corresponding to silicon's four valence electrons.

Melting Points of Ionic Solids

10.3 **(a)** You would expect MgO to have the higher melting point because, in our bonding model, the magnesium and oxygen ions are held together in a crystalline lattice by charges of 2+ for magnesium and 2− for oxygen. In contrast, the NaCl lattice is held together by charges of 1+ for sodium and 1− for chlorine. The experimentally measured melting points of these compounds are 801 °C for NaCl and 2852 °C for MgO, in accordance with our model.

Energy and the Octet Rule

10.4 **(b)** The reasons that atoms form bonds are complex. One contributing factor is the lowering of their potential energy. The octet rule is just a handy way to predict the combinations of atoms that have a lower potential energy when they bond.

Periodic Trends in Electronegativity

10.5 **(a)** $N > P > Al > Na$

Percent Ionic Character

10.6 **(b)** You are given that the dipole moment of the HCl bond is about 1 D and that the bond length is 127 pm. Previously, we calculated the dipole moment for a 130-pm bond that is 100% ionic to be about 6.2 D. You can therefore estimate the bond's ionic character as $1/6 \times 100$, which is closest to 15%.

Resonance Structures

10.7 **(b)** This structure is not a resonance structure of the ion because it has a different skeletal structure.

Odd-Electron Species

10.8 **(d)** ClO because the sum of the valence electrons of its atoms is an odd number

Expanded Octets

10.9 **(b)** The only molecule in this group that could have an expanded octet is H_3PO_4 because phosphorus is a third-period element. Expanded octets *never* occur in second-period elements such as carbon and nitrogen.

Bond Energies and ΔH_{rxn}

10.10 **(b)** In a highly exothermic reaction, the energy needed to break bonds is less than the energy released when the new bonds form, resulting in a net release of energy.