8. The electrons in a covalent bond occupy overlapping orbitals; each electron is free to occupy either of the orbitals, but both are more likely to be in the space between the nuclei of the bonded atoms.

9. a pair of electrons that is not involved in bonding but instead belongs exclusively to one atom

10. A noble-gas configuration corresponds to a set of outer s and p orbitals that is completely filled with a total of eight electrons. These eight outer electrons are referred to as an octet. Atoms that possess a noble-gas configuration are very stable because the potential energy of their electrons is relatively low. The octet rule states that elements will gain or lose electrons to form a noble-gas configuration.

11. a. 1 d. 6 g. 4
   b. 7 e. 3
   c. 2 f. 5

12. The least-electronegative atom is usually the central atom (except for hydrogen, which is never central). If carbon is present, however, it is usually the central atom, regardless of what other atoms are in the molecule.

13. A single bond involves one pair of electrons, as in the bond between hydrogen and oxygen in water, H₂O. A double bond involves two electron pairs, as in the bond between the carbon and oxygen atoms in a molecule of carbon dioxide, CO₂. A triple bond involves three electron pairs, as in the bond between carbon and nitrogen in hydrogen cyanide, HCN.

14. A multiple bond is needed when there are not enough valence electrons to complete octets by adding unshared pairs.

15. a. An ionic compound is composed of cations and anions such that the total positive and negative charges are equal.
   b. Most ionic compounds occur naturally as crystalline solids.

16. a. the simplest collection of atoms from which an ionic compound’s formula can be established
   b. one calcium ion, Ca²⁺, and two fluorine ions, F⁻

17. a. the energy released when one mole of an ionic compound is formed from gaseous ions
   b. The greater the lattice energy, the stronger the ionic bonding.

18. a. Ionic compounds have higher melting and boiling points than molecular compounds do, and they do not vaporize at room temperature.
   b. The differences in the properties of ionic and molecular compounds are generally a result of differences in how strongly the compound’s basic units are held together.
   c. hardness, brittleness, electrical conductivity in the molten state

19. a. a charged group of covalently bonded atoms
   b. Some common examples of polyatomic ions include the nitrate ion, NO₃⁻, the ammonium ion, NH₄⁺, the sulfate ion, SO₄²⁻, and the phosphate ion, PO₄³⁻.
   c. Polyatomic ions combine with ions of opposite charge to form ionic compounds.

20. a. Metals are better conductors of heat than ionic or molecular compounds. In the solid state, metals are more easily deformed and are better electrical conductors than solid ionic or molecular compounds. Unlike ionic and molecular compounds, metals are also shiny in appearance.
b. Metals are good electrical conductors because of the presence of highly mobile electrons within the bonding networks of their atoms.

21. Most contain sparsely populated outermost orbitals, they have low ionization energies, and they have low electronegativities.

22. a. Metallic bonding results from the attraction between metal atoms and a sea of surrounding electrons.
   b. A metal’s heat of vaporization is a measure of the strength of the metal’s bonding.

23. a. According to VSEPR theory, the shapes of molecules are classified based on the number of bonding electron pairs and lone pairs that surround a molecule’s central atom.
   b. Both molecules are linear.

24. a. linear
   b. trigonal-planar
   c. tetrahedral
   d. trigonal-bipyramidal
   e. octahedral

25. a. Unshared electron pairs occupy space as bonded electrons do, but they are not part of the visualized molecular geometry.
   b. Double (and triple) bonds are treated the same as single bonds.

26. a. Hybrid orbitals are identically shaped orbitals of equal energy that are produced by the mixing of two or more atomic orbitals of similar, but not identical, energies on the same atom.
   b. The number of hybrid orbitals produced is always equal to the number of orbitals that have combined.

27. a. Intermolecular forces are the forces of attraction between molecules.
   b. Intermolecular forces are weaker than the forces involved

in ionic and metallic bonding.

c. The strongest intermolecular forces occur between polar molecules.

28. The more-electronegative atom in a covalent bond draws electrons toward it, creating a polar bond.

29. a. Dipole-dipole forces are the forces of attraction between polar molecules.
   b. The overall polarity of a molecule is determined by the polarity of the molecule’s individual bonds as well as the orientation of the bonds with respect to one another.

30. a. An induced dipole is an instantaneous dipole that is produced in a nonpolar molecule when the molecule’s electrons are momentarily attracted by a polar molecule.
   b. Induced dipoles account for the solubility of nonpolar compounds, such as oxygen, in polar compounds, such as water.

31. a. Hydrogen bonding is a particularly strong dipole-dipole force that occurs among molecules containing hydrogen atoms and highly electronegative atoms, such as those of N, O, Cl, and F.
   b. Because of the great electronegativity difference between H and F, N, O, or Cl, an atom of hydrogen has a positive charge approaching that of a proton. This, coupled with the small size of the hydrogen atom, results in a very strong dipole-dipole attraction.

32. London dispersion forces are intermolecular forces resulting from the creation of instantaneous dipoles.

33. See page 199A.

34. K and Br, H and F, Si and Cl, S and O, H and I/Cand H, Se and S

35. See page 199A.

36. Bonding is stronger between the ions in sodium chloride because its lattice energy is greater (more negative). Greater lattice energy indicates stronger ionic bonding.

37. a. Li
   b. Ca
   c. Cl
   d. O
   e. C
   f. P
   g. Al
   h. S

38. See page 199A.

39. See page 199A.

40. sp^3 hybrid orbitals

41. See page 199A.

42. See page 199A.

43. a. trigonal-pyramidal
   b. bent or angular
   c. bent or angular

44. The carbon atom contains four valence electrons, two in the 2s and two in the 2p orbitals. Hybridization of the 2s orbital and the three 2p orbitals creates four sp^3 hybrid orbitals, each of which can bond with a hydrogen atom to form four covalent bonds.

45. The direction of the dipole is toward
   a. F
   b. Cl
   c. Br
   d. I

46. a. nonpolar
   b. polar
   c. polar
   d. nonpolar
   e. polar
   f. polar

47. a. polar
   b. nonpolar
   c. nonpolar

48. See page 199A.

49. See page 199B.

50. d, c, a, b

51. a. tetrahedral
   b. linear
   c. trigonal-pyramidal

52–75. See page 199B.
Additional Sample Problem from page 163

6-1 Ask students to complete the following chart:

<table>
<thead>
<tr>
<th>Elements bonded</th>
<th>Electro-negativity difference</th>
<th>Bond type</th>
<th>More-negative atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. C and H</td>
<td>0.4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>b. C and S</td>
<td>0.0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>c. O and H</td>
<td>1.4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>d. Na and Cl</td>
<td>2.1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>e. Cs and S</td>
<td>1.8</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Ans.

a. polar-covalent; C
b. nonpolar-covalent; same electronegativity

c. polar-covalent; O
d. ionic; Cl
e. ionic; S

Additional Sample Problem from page 171

6-3 Draw the Lewis structure for each of the following compounds:

a. H₂O
b. CH₄
c. CH₃O

d. HCl

Ans.

a. H·O·H
b. H
H·C·H
H

c. H·C·O·H

Additional Sample Problem from page 174

6-4 Determine the Lewis structure for each of the following molecules:

a. O₂
b. C₂H₄
c. C₂H₂

Ans.

a. Ō·Ō
b. H·H
H·C·C
H·H

c. H·C·C·H

Answers from page 198

33.

<table>
<thead>
<tr>
<th>Elements bonded</th>
<th>Electro-negativity difference</th>
<th>Bond type</th>
<th>More-negative atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. H and I</td>
<td>0.4</td>
<td>polar-covalent</td>
<td>I</td>
</tr>
<tr>
<td>b. S and O</td>
<td>1.0</td>
<td>polar-covalent</td>
<td>O</td>
</tr>
<tr>
<td>c. K and Br</td>
<td>2.0</td>
<td>ionic</td>
<td>Br</td>
</tr>
<tr>
<td>d. Si and Cl</td>
<td>1.2</td>
<td>polar-covalent</td>
<td>Cl</td>
</tr>
<tr>
<td>e. H and F</td>
<td>1.9</td>
<td>ionic</td>
<td>F</td>
</tr>
<tr>
<td>f. Se and S</td>
<td>0.1</td>
<td>nonpolar-covalent</td>
<td>S</td>
</tr>
<tr>
<td>g. C and H</td>
<td>0.4</td>
<td>polar-covalent</td>
<td>C</td>
</tr>
</tbody>
</table>

35.

a. Cl

b. O

c. F
38. a. $Na^+ + Na^+ + S^{2-} \rightarrow Na^+ + Na^+ + S^{2-}$ ($Na_2S$)

b. $Ca^+ + O^2- \rightarrow Ca^{2+} + O^2-$ (CaO)

c. $Al^{3+} + Al^3+ + S^{2-} + S^{2-} \rightarrow Al^{3+} + Al^{3+} + S^{2-} + S^{2-}$ ($Al_2S_3$)

39. a. $F:F:

b. H:Se:H
e. $H:Cl:H$

c. $I:Br:1$

d. Br:Si:Br:

40. a. $:\cdot O:O:\cdot$
c. $:\cdot C:O:\cdot$
b. $:\cdot N:O:\cdot$
d. $:\cdot O:S:O: \leftrightarrow O:S:O:$

42. a. $[O:H]^-$
c. $[O:Br:O:]-$

43. a. $:\cdot Cl:Cl:S:Cl:\cdot$
b. $:\cdot I:Br:I:1$
c. $:\cdot Cl:O:Cl:\cdot$
d. $H:N:H$
e. $\cdot Br:$

44. a. $\cdot O:O:\cdot$
b. $\cdot O:\cdot$
c. $\cdot O:\cdot$

45. a. In ionic bonding, valence electrons of the atoms of the less-electronegative element are donated entirely to the atoms of the more-electronegative element. In covalent bonding, valence electrons are shared between the bonded atoms.
b. A molecular compound consists of individual units capable of existing on their own. An ionic compound consists of an arrangement of a large number of ions. There is no discrete, independent particle in an ionic compound.
c. An ionic compound is held together by electrical attraction between ions. A metal is held together by the sharing by atoms of a sea of mobile valence electrons.

52. a. metals and nonmetals
b. nonmetals only
c. metals only

53. The energies of the atoms increase and the atoms become less stable.

54. $\cdot O:O: \leftrightarrow O:S:O: \leftrightarrow S\cdots S\cdots S\cdots S\cdots S\cdots O:O:$

55. a. In ionic bonding, valence electrons of the atoms of the less-electronegative element are donated entirely to the atoms of the more-electronegative element. In covalent bonding, valence electrons are shared between the bonded atoms.
b. A molecular compound consists of individual units capable of existing on their own. An ionic compound consists of an arrangement of a large number of ions. There is no discrete, independent particle in an ionic compound.
c. An ionic compound is held together by electrical attraction between ions. A metal is held together by the sharing by atoms of a sea of mobile valence electrons.

56. a. $\cdot He:\cdot$
c. $\cdot O:\cdot$
e. $\cdot B:\cdot$

57. $\cdot H:\cdot$

58. two potassium cations, $K^+$, and one sulfide anion, $S^{2-}$

59. Vacant orbitals in the outer energy levels of the metal's atoms overlap and are occupied by electrons from adjacent atoms.
60. Although a particular bond in a molecule may be polar, it is the arrangement of all the polar bonds in space—or molecular geometry—that determines whether a molecule is polar.

61. It lies between the energy levels of the orbitals from which it was made.

62. aluminum

63.

<table>
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<tr>
<td>b. Br and I</td>
<td>0.3</td>
<td>nonpolar-covalent</td>
<td>Br</td>
</tr>
<tr>
<td>c. S and Cl</td>
<td>0.5</td>
<td>polar-covalent</td>
<td>Cl</td>
</tr>
</tbody>
</table>

64. a. Ï€:Cl−P−Cl:
    Ï€:Cl:

b. :Cl:
    :F:C:Cl:
    :F:

c.  H:→C:→N:→H
    H:→H

65. Ï€:Cl−Be−Cl:

66. a. [O::N::O]−
b. [O::N::O]−
    trigonal-planar
c. [H
    H::N::H
    H]
    tetrahedral

67. Most atoms are bonded to other atoms in nature because bonding lowers their potential energy.

68. The carbon-carbon triple bond in C₂H₂ is the strongest of the bonds. Therefore, it is the shortest because stronger bonds are shorter bonds. The carbon-carbon single bond in C₂H₆ is the weakest of the bonds, so it is the longest. From Table 6-2, students should surmise that the lengths of the single, double, and triple carbon-carbon bonds are about 154 pm, 134 pm, and 120 pm, respectively.

69. a. C, C and O, O
    b. polar covalent
    c. ionic

70. Lithium, sodium, potassium, rubidium, cesium, barium, and radium are listed as bcc.

71. The electrons in metals require very little energy to enter the conduction band, which is a set of overlapping orbitals. Thus, electrons can move freely through a metal with a small applied voltage. Refer to the One-Stop Planner CD-ROM for appropriate scoring rubrics for items 72 and 73. Look for these points in each report:

72. a. Students should mention Dr. Pauling’s role in establishing the electronegativity scale and in discovering the hydrogen bond. For this work, Dr. Pauling received a Nobel Prize.
   b. Some scientists criticize Pauling for being an advocate of something as trendy as vitamin supplements. Be sure students examine real scientific data, rather than political responses, to make their decision.

73. Students should discuss some properties of metalloids.

74. Students might use either physical or chemical properties to distinguish between these substances. Ionic substances will be solid at room temperature, eliminating water. They will not conduct electricity, eliminating Cu. And they will be composed of more than one type of atom, eliminating C. Ionic solids are brittle and have high melting points.

75. Students’ answers will vary greatly. Some possible responses include plastics (covalent), paper (covalent), wood (covalent), dishes (ionic or covalent), pots (metallic), and cloth (covalent).