SECTION 1

SHORT ANSWER   Answer the following questions in the space provided.

1. _____ A chemical bond between atoms results from the attraction between the valence electrons and ____ of different atoms.
   (a) nuclei
   (b) inner electrons
   (c) isotopes
   (d) Lewis structures

2. _____ A covalent bond consists of
   (a) a shared electron.
   (b) a shared electron pair.
   (c) two different ions.
   (d) an octet of electrons.

3. _____ If two covalently bonded atoms are identical, the bond is identified as
   (a) nonpolar covalent.
   (b) polar covalent.
   (c) ionic.
   (d) dipolar.

4. _____ A covalent bond in which there is an unequal attraction for the shared electrons is
   (a) nonpolar.
   (b) polar.
   (c) ionic.
   (d) dipolar.

5. _____ Atoms with a strong attraction for electrons they share with another atom exhibit
   (a) zero electronegativity.
   (b) low electronegativity.
   (c) high electronegativity.
   (d) Lewis electronegativity.

6. _____ Bonds that possess between 5% and 50% ionic character are considered to be
   (a) ionic.
   (b) pure covalent.
   (c) polar covalent.
   (d) nonpolar covalent.

7. _____ The greater the electronegativity difference between two atoms bonded together, the greater the bond’s percentage of
   (a) ionic character.
   (b) nonpolar character.
   (c) metallic character.
   (d) electron sharing.

8. The electrons involved in the formation of a chemical bond are called ______ valence electrons _______.

9. A chemical bond that results from the electrostatic attraction between positive and negative ions is called a(n) ______ ionic bond _______.

SECTION 1 continued

10. If electrons involved in bonding spend most of the time closer to one atom rather than
   the other, the bond is _________ polar covalent _________.

11. If a bond’s character is more than 50% ionic, then the bond is called
   a(n) _________ ionic bond _________.

12. A bond’s character is more than 50% ionic if the electronegativity difference between the two
   atoms is greater than _________ 1.7 _________.

13. Write the formula for an example of each of the following compounds:
   Answers will vary.
   
   _________ H2 _________ a. nonpolar covalent compound
   _________ HCl _________ b. polar covalent compound
   _________ NaCl _________ c. ionic compound

14. Describe how a covalent bond holds two atoms together.
   A pair of electrons is attracted to both nuclei of the two atoms bonded together.

15. What property of the two atoms in a covalent bond determines whether or not the bond will be
   polar?
   electronegativity

16. How can electronegativity be used to distinguish between an ionic bond and a covalent bond?
   The difference between the electronegativity of the two atoms in a bond will
determine whether the bond is ionic or covalent. If the difference in
   electronegativity is greater than 1.7, the bond is considered ionic.

17. Describe the electron distribution in a polar-covalent bond and its effect on the partial charges of
   the compound.
   The electron density is greater around the more electronegative atom, giving that
   part of the compound a partial negative charge. The other part of the compound has an equal partial positive charge.
SECTION 2

SHORT ANSWER  Answer the following questions in the space provided.

1. Use the concept of potential energy to describe how a covalent bond forms between two atoms.

   As the atoms involved in the formation of a covalent bond approach each other, the electron-proton attraction is stronger than the electron-electron and proton-proton repulsions. The atoms are drawn to each other and their potential energy decreases. Eventually, a distance is reached at which the repulsions between the like charges equals the attraction of the opposite charges. At this point, potential energy is at a minimum and a stable molecule forms.

2. Name two elements that form compounds that can be exceptions to the octet rule.

   Choose from hydrogen, boron, beryllium, phosphorus, sulfur, and xenon.

3. Explain why resonance structures are used instead of Lewis structures to correctly model certain molecules.

   Resonance structures show that one Lewis structure cannot correctly represent the location of electrons in a bond. Resonance structures show delocalized electrons, while Lewis structures depict electrons in a definite location.

4. Bond energy is related to bond length. Use the data in the tables below to arrange the bonds listed in order of increasing bond length, from shortest bond to longest.

   a.

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—F</td>
<td>569</td>
</tr>
<tr>
<td>H—I</td>
<td>299</td>
</tr>
<tr>
<td>H—Cl</td>
<td>432</td>
</tr>
<tr>
<td>H—Br</td>
<td>366</td>
</tr>
</tbody>
</table>

   H—F, H—Cl, H—Br, H—I
b. Bond energy (kJ/mol)

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C—C</td>
<td>346</td>
</tr>
<tr>
<td>C≡C</td>
<td>835</td>
</tr>
<tr>
<td>C≡C</td>
<td>612</td>
</tr>
<tr>
<td>C≡C, C≡C, C—C</td>
<td></td>
</tr>
</tbody>
</table>

5. Draw Lewis structures to represent each of the following formulas:

a. \( \text{NH}_3 \)

```
H — N — H
  |       |
  H
```

b. \( \text{H}_2\text{O} \)

```
H — O —·
  |       |
  H
```

c. \( \text{CH}_4 \)

```
  H
  |
H — C — H
  |
  H
```

d. \( \text{C}_2\text{H}_2 \)

```
  H — C≡C — H
```

e. \( \text{CH}_2\text{O} \)

```
H
  |
H — C≡O
```

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 3

SHORT ANSWER  Answer the following questions in the space provided.

1. **a** The notation for sodium chloride, NaCl, stands for one
   (a) formula unit.   (c) crystal.
   (b) molecule.      (d) atom.

2. **d** In a crystal of an ionic compound, each cation is surrounded by a number of
   (a) molecules.     (c) dipoles.
   (b) positive ions. (d) negative ions.

3. **b** Compared with the neutral atoms involved in the formation of an ionic compound, the
   crystal lattice that results is
   (a) higher in potential energy. (c) equal in potential energy.
   (b) lower in potential energy. (d) unstable.

4. **b** The lattice energy of compound A is greater in magnitude than that of compound B. What
   can be concluded from this fact?
   (a) Compound A is not an ionic compound.
   (b) It will be more difficult to break the bonds in compound A than those in compound B.
   (c) Compound B has larger crystals than compound A.
   (d) Compound A has larger crystals than compound B.

5. **b** The forces of attraction between molecules in a molecular compound are generally
   (a) stronger than the attractive forces among formula units in ionic bonding.
   (b) weaker than the attractive forces among formula units in ionic bonding.
   (c) approximately equal to the attractive forces among formula units in ionic bonding.
   (d) equal to zero.

6. Describe the force that holds two ions together in an ionic bond.

   **The force of attraction between unlike charges holds a negative ion and a positive ion together in an ionic bond.**

7. What type of energy best represents the strength of an ionic bond?

   **lattice energy**
8. What types of bonds are present in an ionic compound that contains a polyatomic ion?

The atoms in a polyatomic ion are held together with covalent bonds, but polyatomic ions combine with ions of opposite charge to form ionic compounds.

9. Arrange the ionic bonds in the table below in order of increasing strength from weakest to strongest.

<table>
<thead>
<tr>
<th>Ionic bond</th>
<th>Lattice energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>-787</td>
</tr>
<tr>
<td>CaO</td>
<td>-3384</td>
</tr>
<tr>
<td>KCl</td>
<td>-715</td>
</tr>
<tr>
<td>MgO</td>
<td>-3760</td>
</tr>
<tr>
<td>LiCl</td>
<td>-861</td>
</tr>
</tbody>
</table>

KCl, NaCl, LiCl, CaO, MgO

10. Draw Lewis structures for the following polyatomic ions:

a. \( \text{NH}_4^+ \)

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{N} \\
\text{H} \\
\text{H}
\end{array}
\]

b. \( \text{SO}_4^{2-} \)

\[
\begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array}
\]

11. Draw the two resonance structures for the nitrite anion, \( \text{NO}_2^- \).

\[
\begin{array}{c}
\text{O} \\
\text{N} \\
\text{O}
\end{array} \quad \leftrightarrow \quad \begin{array}{c}
\text{O} \\
\text{N} \\
\text{O}
\end{array}
\]
CHAPTER 6 REVIEW

Chemical Bonding

SECTION 4

SHORT ANSWER  Answer the following questions in the space provided.

1. ____ b ____ In metals, the valence electrons are considered to be
   (a) attached to particular positive ions.   (c) immobile.
   (b) shared by all surrounding atoms.      (d) involved in covalent bonds.

2. ____ a ____ The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their
   (a) chemical bonds.      (c) enthalpies of vaporization.
   (b) London forces.       (d) polarity.

3. ____ d ____ As light strikes the surface of a metal, the electrons in the electron sea
   (a) allow the light to pass through.
   (b) become attached to particular positive ions.
   (c) fall to lower energy levels.
   (d) absorb and re-emit the light.

4. ____ d ____ Mobile electrons in the metallic bond are responsible for
   (a) luster.               (c) electrical conductivity.
   (b) thermal conductivity. (d) All of the above.

5. ____ a ____ In general, the strength of the metallic bond ____ moving from left to right on any row of the periodic table.
   (a) increases
   (b) decreases
   (c) remains the same
   (d) varies

6. ____ c ____ When a metal is drawn into a wire, the metallic bonds
   (a) break easily.            (c) do not break.
   (b) break with difficulty.  (d) become ionic bonds.

7. Use the concept of electron configurations to explain why the number of valence electrons in metals tends to be less than the number in most nonmetals.
   Most metals have their outer electrons in s orbitals, while nonmetals have their
   outer electrons in p orbitals.
SECTION 4 continued

8. How does the behavior of electrons in metals contribute to the metal’s ability to conduct electricity and heat?

   The mobility of electrons in a network of metal atoms contributes to the metal’s ability to conduct electricity and heat.

9. What is the relationship between the enthalpy of vaporization of a metal and the strength of the bonds that hold the metal together?

   The amount of energy required to vaporize a metal is a measure of the strength of the bonds that hold the metal together. The greater a metal’s enthalpy of vaporization, the stronger the metallic bond.

10. Draw two diagrams of a metallic bond. In the first diagram, draw a weak metallic bond; in the second, show a metallic bond that would be stronger. Be sure to include nuclear charge and number of electrons in your illustrations.

   a. weak bond
   b. strong bond

   Note: In the strong bond, the charge on the nucleus and the number of electrons must be greater than in the weak bond.

11. Complete the following table:

<table>
<thead>
<tr>
<th></th>
<th>Metals</th>
<th>Ionic Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Components</td>
<td>atoms</td>
<td>ions</td>
</tr>
<tr>
<td>Overall charge</td>
<td>neutral</td>
<td>neutral</td>
</tr>
<tr>
<td>Conductive in the solid state</td>
<td>yes</td>
<td>no</td>
</tr>
<tr>
<td>Melting point</td>
<td>low to high</td>
<td>high</td>
</tr>
<tr>
<td>Hardness</td>
<td>soft to hard</td>
<td>hard</td>
</tr>
<tr>
<td>Malleable</td>
<td>yes</td>
<td>no</td>
</tr>
<tr>
<td>Ductile</td>
<td>yes</td>
<td>no</td>
</tr>
</tbody>
</table>
CHAPTER 6 REVIEW

Chemical Bonding

SECTION 5

SHORT ANSWER  Answer the following questions in the space provided.

1. Identify the major assumption of the VSEPR theory, which is used to predict the shape of atoms.

Pairs of valence electrons repel one another.

2. In water, two hydrogen atoms are bonded to one oxygen atom. Why isn’t water a linear molecule?

The electron pairs that are not involved in bonding also take up space, creating a tetrahedron of electron pairs and making the water molecule angular or bent.

3. What orbitals combine together to form $sp^3$ hybrid orbitals around a carbon atom?

the $s$ orbital and all three $p$ orbitals from the second energy level

4. What two factors determine whether or not a molecule is polar?

electronegativity difference and molecular geometry or unshared electron pairs

5. Arrange the following types of attractions in order of increasing strength, with 1 being the weakest and 4 the strongest.

3. hydrogen bonding
4. ionic
2. dipole-dipole
1. London dispersion

6. How are dipole-dipole attractions, London dispersion forces, and hydrogen bonding similar?

They are all forces of attraction between molecules. In all cases there is an attraction between the slightly negatively-charged portion of one molecule and the slightly positively charged portion of another molecule.
7. Complete the following table:

<table>
<thead>
<tr>
<th>Formula</th>
<th>Lewis structure</th>
<th>Geometry</th>
<th>Polar</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂S</td>
<td></td>
<td>bent</td>
<td>yes</td>
</tr>
<tr>
<td></td>
<td><img src="image1" alt="Lewis structure of H₂S" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CCl₄</td>
<td>:Cl: :Cl: :Cl: :Cl:</td>
<td>tetrahedral</td>
<td>no</td>
</tr>
<tr>
<td></td>
<td><img src="image2" alt="Lewis structure of CCl₄" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td><img src="image3" alt="Lewis structure of BF₃" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td><img src="image4" alt="Lewis structure of H₂O" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>PCl₅</td>
<td>:Cl: :Cl: :Cl: :Cl:</td>
<td>trigonal bipyramidal</td>
<td>no</td>
</tr>
<tr>
<td></td>
<td><img src="image5" alt="Lewis structure of PCl₅" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>BeF₂</td>
<td>:F: :Be: :F:</td>
<td>linear</td>
<td>no</td>
</tr>
<tr>
<td></td>
<td><img src="image6" alt="Lewis structure of BeF₂" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td><img src="image7" alt="Lewis structure of SF₆" /></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
CHAPTER 6 REVIEW

Chemical Bonding

MIXED REVIEW

SHORT ANSWER   Answer the following questions in the space provided.

1. Name the type of energy that is a measure of strength for each of the following types of bonds:
   - lattice energy
   - bond energy
   - enthalpy of vaporization
     a. ionic bond
     b. covalent bond
     c. metallic bond

2. Use the electronegativity values shown in Figure 20, on page 161 of the text, to determine whether each of the following bonds is nonpolar covalent, polar covalent, or ionic.
   - ionic  a. H—F
   - ionic  b. Na—Cl
   - polar covalent  c. H—O
   - nonpolar covalent  d. H—H
   - polar covalent  e. H—C
   - polar covalent  f. H—N

3. How is a hydrogen bond different from an ionic or covalent bond?
   A hydrogen bond is a dipole-dipole attraction between a partially positive hydrogen atom and the unshared electron pair of a strongly electronegative atom such as O, N, or F. Unlike ionic or covalent bonds, in which electrons are given up or shared, the hydrogen bond is a weaker attraction. Hydrogen bonds are generally intermolecular, while ionic and covalent bonds occur between ions or atoms respectively.

4. H₂S and H₂O have similar structures and their central atoms belong to the same group. Yet H₂S is a gas at room temperature and H₂O is a liquid. Use bonding principles to explain why this is.
   Oxygen has higher electronegativity than sulfur, which creates a highly polar bond. Increased polarity in H₂O bonds means a stronger intermolecular attraction, making water a liquid at room temperature. Hydrogen bonding exists between water molecules, but not between hydrogen sulfide molecules.
MIXED REVIEW continued

5. In what way is a polar-covalent bond similar to an ionic bond?

There is a difference between the electronegativities of the two atoms in both types of bonds that results in electrons being more closely associated with the more electronegative atom.

6. Draw a Lewis structure for each of the following formulas. Determine whether the molecule is polar or nonpolar.

- **polar**  
  a. H₂S

\[
\begin{array}{c}
\text{H} \\
\text{S} \\
\text{H}
\end{array}
\]

- **polar**  
  b. COCl₂

\[
\begin{array}{c}
\text{Cl} \\
\text{C} = \text{O} \\
\text{Cl}
\end{array}
\]

- **polar**  
  c. PCl₃

\[
\begin{array}{c}
\text{Cl} \\
\text{P} \\
\text{Cl}
\end{array}
\]

- **polar**  
  d. CH₂O

\[
\begin{array}{c}
\text{H} \\
\text{H} \text{C} = \text{O}
\end{array}
\]