Name ________________ Date ________ Class ________

CHAPTER 6 REVIEW

Chemical Bonding

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
(c) isotopes
(b) inner electrons
(d) Lewis structures

2. ______ A covalent bond consists of

(a) a shared electron.
(c) two different ions.
(b) a shared electron pair.
(d) an octet of electrons.

3. ______ If two covalently bonded atoms are identical, the bond is identified as

(a) nonpolar covalent.
(c) ionic.
(b) polar covalent.
(d) dipolar.

4. ______ A covalent bond in which there is an unequal attraction for the shared electrons is

(a) nonpolar.
(c) ionic.
(b) polar.
(d) dipolar.

5. ______ Atoms with a strong attraction for electrons they share with another atom exhibit

(a) zero electronegativity.
(c) high electronegativity.
(b) low electronegativity.
(d) Lewis electronegativity.

6. ______ Bonds that possess between 5% and 50% ionic character are considered to be

(a) ionic.
(c) polar covalent.
(b) pure 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.
(c) metallic character.
(b) nonpolar character.
(d) electron sharing.

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

________________________.

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

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

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

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

13. Write the formula for an example of each of the following compounds:

_______________ **a.** nonpolar covalent compound

_______________ **b.** polar covalent compound

_______________ **c.** ionic compound

14. Describe how a covalent bond holds two atoms together.

____

____

____

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

____

____

16. How can electronegativity be used to distinguish between an ionic bond and a covalent bond?

____

____

____

____

17. Describe the electron distribution in a polar-covalent bond and its effect on the partial charges of the compound.

____

____

____

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

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.

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

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

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.

Bond	Bond energy (kJ/mol)
H—F	569
H—I	299
H—Cl	432
H—Br	366

b.

Bond	Bond energy (kJ/mol)
C—C	346
C≡C	835
C=C	612

______

______

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

a. NH_3

b. H_2O

c. CH_4

d. C_2H_2

e. CH_2O

Name ______________________ Date ____________ Class ____________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 3

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ The notation for sodium chloride, NaCl, stands for one

(a) formula unit.
(b) molecule.
(c) crystal.
(d) atom.

2. ______ In a crystal of an ionic compound, each cation is surrounded by a number of

(a) molecules.
(b) positive ions.
(c) dipoles.
(d) negative ions.

3. ______ Compared with the neutral atoms involved in the formation of an ionic compound, the crystal lattice that results is

(a) higher in potential energy.
(b) lower in potential energy.
(c) equal in potential energy.
(d) unstable.

4. ______ 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. ______ 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.

____

____

____

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

____

____

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

____

____

____

____

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

Ionic bond	Lattice energy (kJ/mol)
NaCl	−787
CaO	−3384
KCl	−715
MgO	−3760
LiCl	−861

____

____

10. Draw Lewis structures for the following polyatomic ions:

a. NH_4^+

b. SO_4^{2-}

11. Draw the two resonance structures for the nitrite anion, NO_2^-.

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 4

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ In metals, the valence electrons are considered to be

(a) attached to particular positive ions.
(b) shared by all surrounding atoms.
(c) immobile.
(d) involved in covalent bonds.

2. ______ The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their

(a) chemical bonds.
(b) London forces.
(c) enthalpies of vaporization.
(d) polarity.

3. ______ 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. ______ Mobile electrons in the metallic bond are responsible for

(a) luster.
(b) thermal conductivity.
(c) electrical conductivity.
(d) All of the above.

5. ______ 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. ______ When a metal is drawn into a wire, the metallic bonds

(a) break easily.
(b) break with difficulty.
(c) do not break.
(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.

____

____

____

____

SECTION 4 continued

8. How does the behavior of electrons in metals contribute 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?

____

____

____

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

11. Complete the following table:

	Metals	Ionic Compounds
Components		
Overall charge		
Conductive in the solid state		
Melting point		
Hardness		
Malleable		
Ductile		

Name ______________________________ Date ______________ Class ______________

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.

____

____

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

____

____

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

____

____

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

____

____

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

   _______ hydrogen bonding

   _______ ionic

   _______ dipole-dipole

   _______ London dispersion

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

____

____

____

____

7. Complete the following table:

Formula	Lewis structure	Geometry	Polar
H_2S			
CCl_4			
BF_3			
H_2O			
PCl_5			
BeF_2			
SF_6			

Name ______________________ Date ______________ Class ______________

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:

______________________ **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.

______________ **a.** H—F ______________ **d.** H—H

______________ **b.** Na—Cl ______________ **e.** H—C

______________ **c.** H—O ______________ **f.** H—N

3. How is a hydrogen bond different from an ionic or covalent bond?

4. H_2S and H_2O have similar structures and their central atoms belong to the same group. Yet H_2S is a gas at room temperature and H_2O is a liquid. Use bonding principles to explain why this is.

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

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

_______________ **a.** H_2S

_______________ **b.** $COCl_2$

_______________ **c.** PCl_3

_______________ **d.** CH_2O