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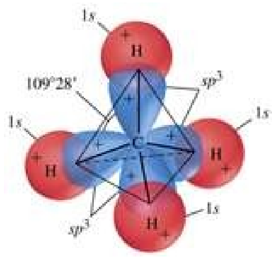


CHAPTER 6 Chemical Bonding

SECTION 1 Introduction to Chemical Bonding

OBJECTIVES

1. **Define Chemical bond.**
2. **Explain why most atoms form chemical bonds.**
3. **Describe ionic and covalent bonding..**
4. **Explain why most chemical bonding is neither purely ionic or purley**
5. **Classify bonding type according to electronegativity differences.**

SECTION 2 Covalent Bonding and Molecular Compounds

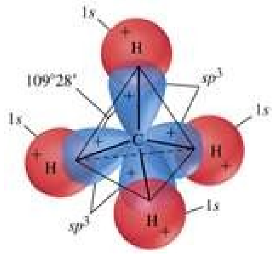
OBJECTIVES

1. **Define molecule and molecular formula.**
2. **Explain the relationships among potential energy, distance between approaching atoms, bonds length and bond energy.**
3. **State the octet rule.**
4. **List the six basic steps used in writing Lewis dot Structures.**
5. **Explain how to determine Lewis structures for molecules containing single bonds, multiple bonds or both.**
6. **Explain why scientists use resonance structures to represent some molecules.**

SECTION 3 Ionic Bonding and Ionic compounds

OBJECTIVES

1. **Compare and contrast the chemical formula for a molecular compound wit one for an ionic compound.**
2. **Discuss the arrangements of ions in crystals.**
3. **Define lattice energy and explain its significance.**
4. **List and compare the distinctive properties of ionic and molecular compounds.**
5. **Write the Lewis structure for a polyatomic ion given the identity of the atoms combined and other appropriate information.**



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SECTION 4 Metallic Bonding

OBJECTIVES

1. **Describe the electron-sea model of metallic bonding, and explain why metals are good electrical conductors.**
2. **Explain why metal surfaces are shinny.**
3. **Explain why metals are malleable and ductile but ionic crystalline compounds are not.**

SECTION 5 Molecular Geometry

OBJECTIVES

1. **Explain VSEPR theory.**
2. **Predict the shapes of molecules or polyatomic ionis using VSEPR theory.**
3. **Explain how the shapes of molecules are accounted for by hybridization theory.**
4. **Describe dipole – dipole forces, hydrogen bonding, induced dipoles and London Dispersion Forces and their effects on properties such as boiling point and melting point.**
5. **Explain what determines molecular polarity.**

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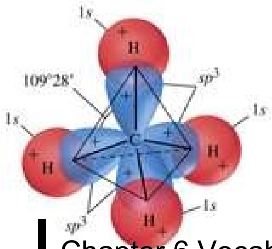


Chapter 6 Vocabulary

Chemical Bonding

Sec 6-1 Introduction to Chemical Bonding

Chemical bond
Ionic bonding
Covalent bonding
Nonpolar covalent bond
Polar
Polar covalent bond

Sec 6-2 Covalent Bonding and Molecular Compounds

Molecule
Molecular compound
Chemical formula
Molecular formula
Bond energy
Electron dot notation
Lewis structure
Single bond
Multiple bond
resonance

Sec 6-3 Ionic Bonding and Ionic Compounds

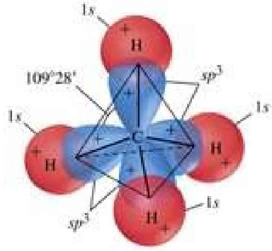
Ionic compound
Formula unit
Lattice energy
Polyatomic ion

Sec 6-4 Metallic Bonding

Metallic bonding
Malleability
ductility

Sec 6-5 Molecular Geometry

VSEPR theory
Hybridization
hybrid orbitals
Intramolecular force
Intermolecular force
dipole
hydrogen bonding
London dispersion forces



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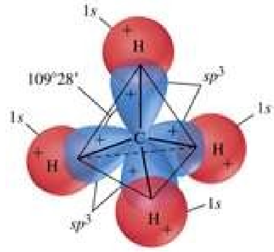

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CHAPTER 6 STUDY GUIDE

Chemical Bonding

SECTION 1 INTRODUCTION TO CHEMICAL BONDING

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.

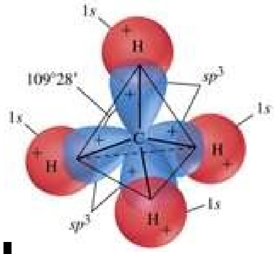
5. _____ Atoms with a strong attraction for electrons they share with another atom exhibit
   (a) zero electronegativity. (c) high electronegativity.
   (b) low electronegativity. (d) Lewis electronegativity.

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 ionic bond negative ions is called a(n) ______________________.



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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 _______________________.
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:

    Answers will vary.

    ____________________ 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?

    ____

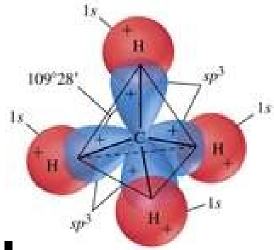
    ____

    ____

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

    ____

    ____



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CHAPTER 6 STUDY GUIDE

Chemical Bonding

SECTION 2 COVALENT BONDING AND MOLECULAR COMPOUNDS

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.

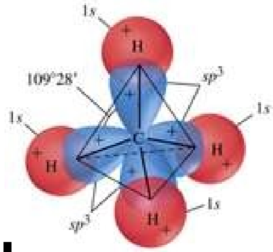
2. Name two elements that form compounds that can be exceptions to the octet rule.

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



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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:

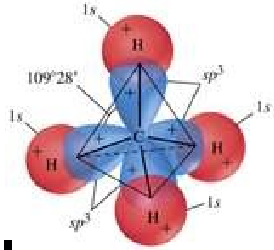
a. NH_3

b. H_2O

c. CH_4

d. C_2H_2

e. CH_2O



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CHAPTER 6 STUDY GUIDE

Chemical Bonding

SECTION 3 IONIC BONDING AND IONIC COMPOUNDS

SHORT ANSWER Answer the following questions in the space provided.

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

   ____

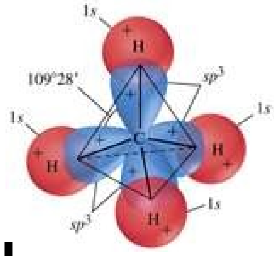
   ____

   ____

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

   ____

   ____



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SECTION 3 continued

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

____

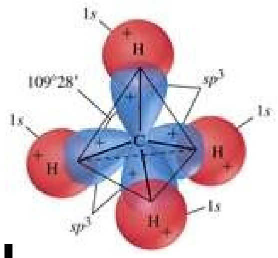
____

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^-.



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CHAPTER 6 STUDY GUIDE

Chemical Bonding

SECTION 4 METALLIC BONDING

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.

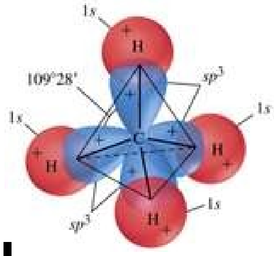
7. Use the concept of electron configurations to explain why the number of valence electrons in metals tends to be less than the number in most nonmetals.

____

____

____

____



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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?

____

____

____

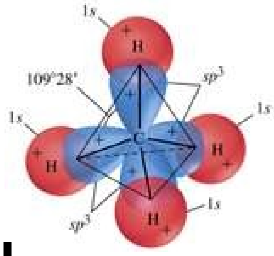
10. Draw two diagrams of a metallic bond. In the first diagram, draw a weak metallic bond; in the second, show a metallic bond that would be stronger. Be sure to include nuclear charge and number of electrons in your illustrations.

a. b.

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:

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



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CHAPTER 6 STUDY GUIDE

Chemical Bonding

SECTION 5 MOLECULAR GEOMETRY

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.

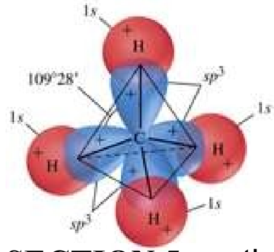
   _______ hydrogen bonding

   _______ ionic

   _______ dipole-dipole

   _______ London dispersion

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



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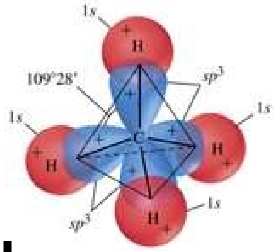
SECTION 5 continued

7. Complete the following table:

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



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CHAPTER 6 STUDY GUIDE

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

________________ b. Na—Cl

__________________c.__ H—O

_____________________ d. H—H

_____________________ e. H—C

_____________________ f. H—N

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

____

____

____

____

____

____

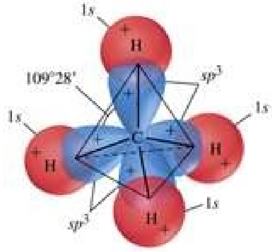
____

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.

____

____

____



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MIXED REVIEW continued

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