Chemistry CP

Name:

Review Sheet — Bonding (Chapter 6)

After studying chapter 6, you should be able to:

- Infer the number of valence electrons in an atom of a main-group element, and then construct its Lewis dot structure.
- Explain why most atoms form chemical bonds.
- Describe the formation of a cation from a metallic element and the formation of an anion from a non-metallic element.
- State the octet rule.
- List the characteristics of a covalent bond.
- Create Lewis structures for covalent compounds containing single, double, and triple bonds.
- Explain the modern interpretation of resonance bonding.
- List the characteristics of an ionic bond.
- Explain the properties of metals by using the concept of metallic bonding.
- Compare and contrast the properties of ionic and molecular compounds.
- Classify bonds as ionic, covalent, or polar based on electronegativity differences.
- Use electron dot formulas to determine chemical formulas for binary ionic compounds.
- Describe the shapes of simple covalently bonded molecules using VSEPR theory.
- Use the bond types and geometry of molecules to determine molecular polarity.
- Explain the relationship between bond length and bond energy.

Problems for you to try:

1. Explain what is meant by the term "chemical bond."

2. What is the difference between a polar bond and a nonpolar bond? (p. 162)

3. Explain how you can use the periodic table to predict relative polarity of bonds. For example, how do you know by looking at the periodic table that a C-O bond is more polar than an N-O bond?

4. Calculate the electronegativity difference for the atoms that are bonded in the following diatomic molecules. Then tell whether the bond is nonpolar covalent, polar covalent, or ionic, and which atom has the greater share of the bonding electrons. (pp. 162-163)

Formula	Electronegativity Difference	Type of Bond	Atom With Greater Electron Share
F_2			
$MgBr_2$			
H_2S			
NaF			

5. State the octet rule. (p. 169)

6. What is resonance? Give an example of a molecule with resonance. Explain your choice. (p. 175)

7. Develop your own analogy for metallic bonding. Explain how this analogy explains the properties of metals. (pp. 181-182)

8. Use Lewis structures to show the transfer of electrons to form ionic compounds from the following combinations of atoms. State the formula of the compound that forms. (p. 177)

   a. Mg and F

   b. Na and Br

   c. Ca and O

   d. Li and Se

9. True or false: If we cannot draw an acceptable Lewis structure for a molecule, that molecule is not stable. Explain.

10. Why do we consider only valence electrons when drawing Lewis diagrams?

11. Does a Lewis structure indicate which electrons came from which atoms? Explain.

12. Why must a Lewis structure for a molecule be drawn before we can determine its geometry?

13. What is the main idea in VSEPR theory?

14. According to the VSEPR model, the arrangements of electron pairs around NH_3 and CH_4

a. are different because there is a different number of atoms around the central atom.
b. are different because there is a different number of electron pairs around the central atom.
c. are the same because both nitrogen and carbon are in the second period.
d. are the same because there is the same number of electron pairs around the central atom.
e. Both a and b are correct.

Provide support for your answer.

15. Draw Lewis structures for the following covalent compounds. (pp. 170-174)

a. $C\ell F$

b. $C\ell_2O$

c. $C_2C\ell_2$

d. PH_3

e. C_2Br_4

f. SiO_2

Methane (CH_4), ammonia (NH_3), and water (H_2O) are common molecular compounds that have hydrogen atoms bonded to atoms of second period elements. Use the table below to answer the next two questions. (pp. 167-168)

Bond	Bond dissociation energy (kJ/mol)
C-H	414
N-H	391
O-H	463

16. Review the relationship between bond length and bond energy to predict which molecules have the longest bonds.

17. Which has the strongest (most stable) bonds?

18. Identify the following shapes of molecules. (pp. 183-185)

a) ______________ b) ______________

c) ______________ d) ______________

e) ______________ f) ______________

g) ______________ h) ______________

19. Draw each molecule using lines to represent bonds and dots to represent lone pairs. Indicate the geometry of the molecule. Use electronegativity differences (see p. 151 in your textbook) of the atoms involved and determine if whether each bond is polar. If it is polar, add an arrowhead to the lines representing the bonds, pointing toward the more electronegative atom. Finally, decide whether the molecule as a whole is polar. (Hint: Do the arrows cancel out or not?) If it is, draw a large arrow near the molecule to indicate the direction of polarity. The first substance has been done as an example. (p. 191)

Formula	Representation	Geometry	Polarity of Bonds?	Polarity of Molecule?
NI_3				
NCl_3				
BCl_3 (note: does not obey octet rule for B)				
CCl_4				
CH_3Cl				

20. You are given an unknown solid. What tests could you perform to determine the types of bonding present in the sample? Be specific. (p. 179)