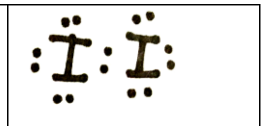
Chapter 16 – Covalent Bonding

Chapter 16: 1 – 26; 28, 30, 31, 35-37, 40, 43-46, *Extra Credit*: 50-53, 55, 56, 58, 59, 62-67

Section 16.1 – The Nature of Covalent Bonding

Practice Problems

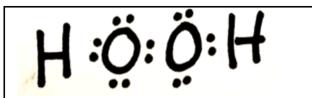
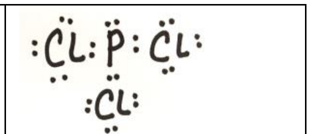
1. Draw electron dot structures for each molecule.

		
a. chlorine	b. bromine	c. iodine

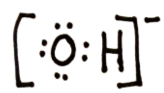
What do you observe about the three structures?

They all look similar with respect to the electrons.

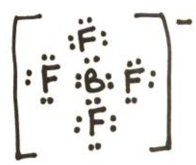
2. The following molecules have single covalent bonds. Draw an electron dot structure for each.

	
a. H_2O_2	b. PCl_3

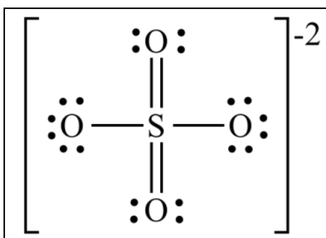
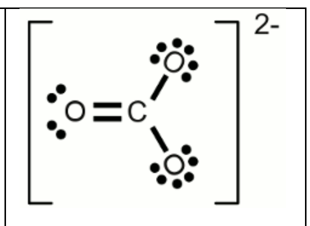
3. Draw the electron dot structure of the hydroxide ion (OH^-).



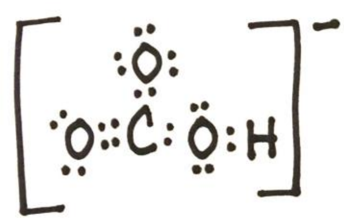
4. Draw the electron dot structure of the polyatomic boron tetrafluoride anion (BF_4^-).



5. Draw the electron dot structures for sulfate (SO_4^{2-}) and carbonate (CO_3^{2-}). Sulfur and carbon are the central atoms, respectively.

	
sulfate	carbonate

6. Draw the electron dot structure for the hydrogen carbonate ion (HCO_3^-). Carbon is the central atom, and hydrogen is attached to oxygen in this polyatomic anion.



Section Review 16.1

7. How are single, double, and triple bonds indicated in electron dot structures?

A single bond is indicated by two dots (:) between two atoms, representing a pair of electrons being shared, or by a single line (-).

A double bond is represented by four dots (::) between two atoms, indicating two pairs of shared electrons or by a double line (=).

A triple bond is represented by six dots (:::) between atoms, or three lines (≡), indicating three pairs of electrons being shared.

8. Provide an example of each of the following – you do not have to draw the structure:

a. coordinate covalent bonding

carbon monoxide (CO)

b. resonance structures

ozone (O_3)

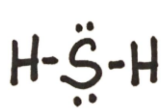
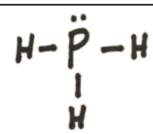
c. exceptions to the octet rule

nitrogen dioxide (NO_2), or phosphorus pentachloride (PCl_5)

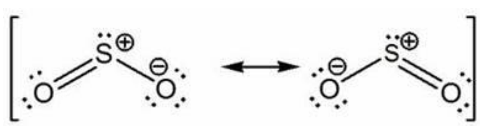
9. What kinds of information does a structural formula reveal about the compound it represents?

The structural formula identifies the atoms in the compound, and their respective number and arrangement within the molecule.

10. Draw electron dot structures for the following molecules, which have only single covalent bonds.

		
a. H_2S	b. PH_3	c. ClF

11. Draw the resonance structures for sulfur dioxide (SO_2). S is the central atom.



Note on the two resonance structures, the partial positive charge by the S, and the partial negative charges on the O's. The two resonance structures then give rise to a *hybrid structure* which is the best representation of the molecule with respect to charge separations.

12. How many kilojoules are required to dissociate all the C-H single bonds in 0.1 mol of methane (CH_4)? Assume that the bond dissociation energy is the same for each bond. Refer to Table 16.3.

$$(4\ C-H\ bonds)\left(393\,\frac{kJ}{mol}\right)(0.1\,mol) = 157\ kJ$$

Section 16.2 – Bonding Theories

Section Review 16.2

13. Use the molecular orbital theory to describe covalent bonding. What occurs during hybridization?

When two atoms combine in a bonding situation, their atomic orbitals overlap to produce molecular orbitals. In hybridization, several atomic orbitals mix to form the same total number of equivalent hybrid orbitals.

14. Explain how the VSEPR theory can be used to predict bond angles in the following covalently bonded molecules.

a. methane

In a methane molecule, the four valence electron pairs repel each other, forming the corners of a tetrahedron in which the pairs are equidistant from each other with bond angles of 109.5 degrees.

b. ammonia

In an ammonia molecule, the four valance electron pairs repel each other, but the unshared pair repels the bonded pairs even more strongly, with bond angles of 107 degrees.

c. water

In a water molecule, the four valence electron pairs repel each other, but the two unshared pairs repel the bonded pairs even more strongly, with bond angles of 105 degrees.

15. What shape would you expect a simple carbon-containing compound to have if the carbon atom has the following hybridizations.

a. sp^2 trigonal planar b. sp^3 tetrahedral c. *sp* linear

16. What is a sigma bond? Describe, with the aid of a diagram, how the overlap of two half-filled *1s* orbitals produces a sigma bond.

A sigma bond is formed by the overlap of two s orbitals, the overlap of an s orbital with a p orbital, or the end-to-end overlap of two p orbitals. Refer to figure 16.11.

17. How many sigma and how many pi bonds are in an ethyne molecule (C_2H_2)? Draw the Lewis structure.

3 sigma bonds, and 2 pi bonds.

18. The BF_3 molecule is planar. The attachment of a fluoride ion to the boron in BF_3, through a coordinate covalent bond, creates the BF_4^- ion. What is the geometric shape of this ion? Draw the Lewis structure.

Tetrahedral.

Section 16.3 – Polar Bonds and Molecules

Practice Problems

19. Identify the bonds between atoms of each pair of elements as (1) nonpolar covalent, (2) moderately polar covalent, (3) very polar covalent, or (4) ionic. Refer to Table 16.4, and 14.2 (p. 405).

a. H and Br	moderately polar covalent	d. Cl and F	moderately polar covalent
b. K and Cl	very polar covalent	e. Li and O	ionic
c. C and O	moderately polar covalent	f. Br and Br	nonpolar covalent

20. Order the following covalent bonds from least to most polar:

a. H-Cl 0.9 b. H-Br 0.7 c. H-S 0.4 d. H-C 0.4 e. F-F 0

e, d, c, b, a

Section Review 16.3

21. Explain how you can use electronegativity values to classify a bond as nonpolar covalent, polar covalent, or ionic.

Find the difference in electronegativity values for the two atoms. Then use Table 16.4 to determine the most likely type.

22. Describe the three kinds of attractive forces that hold groups of molecules together. Rank these forces from weakest to strongest.

a. Dispersion forces – the weakest of the three, are caused by the motion of electrons.

b. Dipole interactions are the attractions between the oppositely charged ends of polar molecules.

c. Hydrogen bonding – the strongest of the three, occurs when a hydrogen atom bonded to a more electronegative atom is attracted to another highly electronegative atom.

23. Not every molecule with polar bonds is polar. Explain this statement, using CCl_4 as an example.

The atoms in CCl_4 are oriented so that the bond polarities cancel.

24. Draw the electron dot structure for each molecule below. Identify polar covalent bonds by assigning slightly positive ($\delta +$) and slightly negative ($\delta -$) symbols to the appropriate atoms.

a. HOOH	b. BrCl	c. HBr	d. H_2O

25. How does a network solid differ from most other covalent compounds?

The atoms in a network solid are covalently bonded in a large array (or crystal), which can be thought of as a single molecule.

26. Which of the following are characteristic of most covalent compounds?

a. high melting points
b. shared bonding electrons
c. low water solubility
d. existence as molecules
e. composed of a metal and a nonmetal

Chapter 16 Review

Concept Practice

28. Classify the following compounds as ionic or covalent. 16.1

a. $MgCl_2$ Ionic   b. Na_2S Ionic   c. H_2O covalent   d. H_2S covalent

30. How many electrons do atoms in a double covalent bond share? How many in a triple bond? 16.1

Atoms in a double covalent bond share four electrons; in a triple bond they share six.

31. Based upon the examples provided in Section 16.1, state a general rule for determining which atom is the central one in a binary molecule compound. 16.1

A single atom (count of only 1), is typically the central atom. Water, with two hydrogens, and only a single atom of oxygen, has a central atom of oxygen.

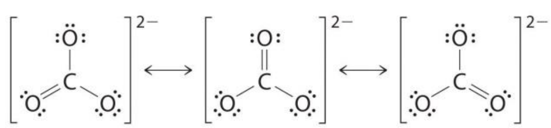
35. Explain why compounds containing C-N and C-O single bonds can form coordinate covalent bonds with H^+ but compounds containing only C-H and C-C bonds cannot. 16.1

An unshared pair of electrons is needed for a coordinate covalent bond. There are no unshared pairs in C-H and C-C bonds.

36. What is true for the electron dot structures of all compounds that exhibit resonance? 16.1

The molecules of each compound can be described by more than one Lewis structure.

37. Draw resonance structures for the carbonate ion (CO_3^{2-}). Each oxygen is attached to the carbon. 16.1. There are three of these structures.



39. How can you experimentally determine whether a substance is paramagnetic? 16.1

The measure mass of a paramagnetic substance appears greater when measured in the presence of a magnetic field than when measured in the absence of one.

40. Predict whether the following species are diamagnetic or paramagnetic. 16.1

a. BF_3 diamagnetic b. O_2^- paramagnetic c. NO_2 paramagnetic d. F_2 diamagnetic

43. What is the relationship between the magnitude of a molecule's bond dissociation energy and its expected chemical reactivity? 16.1

Increasing bond dissociation energy is linked to lower chemical reactivity.

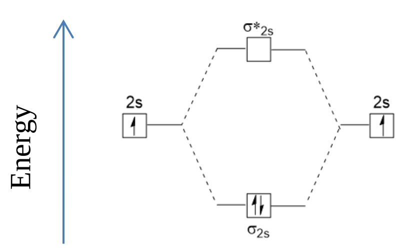
44. Explain what is meant by *bond dissociation energy*.

Bond dissociation energy is defined as the energy needed to break one covalent bond.

45. Assume the total bond energy in a molecule is the sum of the individual bond energies. Calculate the total bond energy in a mole of ethyne (C_2H_2). *Hint*: Write the electron dot structure to determine the kinds of bonds. Then refer to Table 16.3.

2 H-C bonds (393 kJ/ mol) + 1 C≡C bond (908 kJ/ mol) = 1,694 kJ/ mol total

46. Draw the molecular orbital diagrams for the possible diatomic molecule Li_2. Would you expect Li_2 to exist as a stable molecule? 16.2



Extra Credit. Answer the following questions below, points awarded extra credit are indicated and will be applied towards your homework grade. These questions are recommended for students planning on taking organic chemistry anytime in the future.

50. What types of hybrid orbitals are involved in the bonding of the carbon atoms in the following molecules? Draw their Lewis structures. 16.2 (8 possible points)

a. CH_4

b. $H_2C{=}CH_2$

c. $HC{\equiv}CH$

d, $N{\equiv}C{-}C{\equiv}N$

51. What must always be true if a covalent bond is to be polar? 16.3 (1 point)

52. The bonds between the following pairs of elements are covalent. Arrange them according to polarity, naming the most polar bond first. 16.3 (2 points)

a. H-Cl   b. H-C   c. H-F   d. H-O   e. H-H   f. S-Cl

53. Arrange the following bonds in order of increasing ionic character. 16.3 (2 points)

a. Cl-F   b. N-N   c. K-O   d. C-H   e. S-O   f. Li-F

55. Depict (with a drawing) the hydrogen bonding between two ammonia molecules; then depict the bonding between one ammonia and one water molecule. 16.3 (2 points)

Between two ammonia	Between one ammonia and one water

56. Circle the compound in each pair that exhibits the stronger intermolecular hydrogen bonding. Explain your rationale behind your answer. 16.3 (4 points)

a. H_2S, H_2O

b. HCl, HF

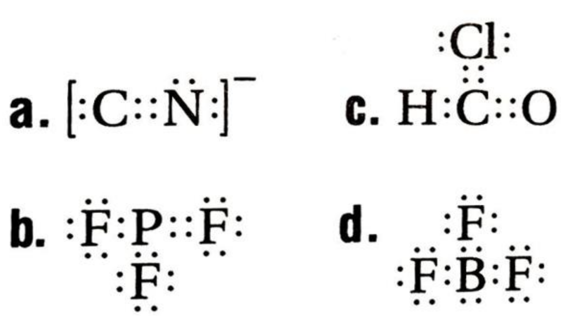
c. HBr, HCl

d. NH_3, H_2O

58. Explain why compounds with strong intermolecular attractive forces have higher boiling points than compounds with weak intermolecular attractive forces. 16.3 (2 points)

59. Using Figures 16.17 through 16.19 as an example, devise a hybridization scheme for PCl_3 and predict the molecular shape based on this scheme. (3 points)

62. Explain why each Lewis structure below is incorrect. Replace each structure with one that is more acceptable. (8 points)

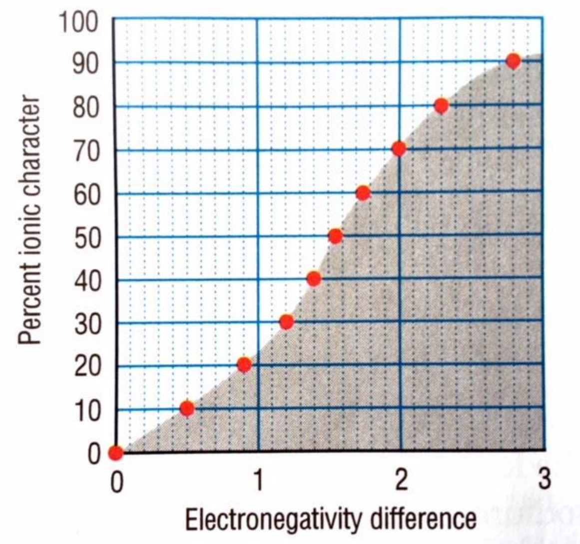


63. Use VSEPR theory to predict the geometry of each of the following: (4 points)

a. $SiCl_4$ b. CO_3^{2-}

c. CCl_4 d. SCl_2

64. The following graph shows how the percent ionic character of a single bond varies according to the difference in electronegativity between the two elements forming the bond. Answer the following questions, using this graph, and Table 14.2. (6 points)



a. What is the relationship between the percent ionic character of single bonds and the electronegativity difference of their elements?

b. What electronegativity difference will result in a bond with a 50% ionic character?

c. Estimate the percent ionic character of the bonds formed between:

(1) lithium and oxygen

(2) nitrogen and oxygen

(3) magnesium and chlorine

(4) nitrogen and fluorine

65. Using bond dissociation energies, estimate Δ*H* for the following reaction. (2 points)

$CO(g) + 2\ H_2(g) \rightarrow CH_3OH(g)$

66. Give the angles between the orbitals of each hybrid. (3 points)

a. sp^3 hybrids b. sp^2 hybrids c. sp hybrids

67. Describe the different between a bonding molecular orbital and an antibonding molecular orbital. How do the energies of these orbitals compare? (3 points)