CHAPTER 3 Review

Reflecting on Chapter 3

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Elements combine to form a wide variety of compounds.
- Ionic compounds and covalent compounds have characteristic properties. You can use these properties to classify various compounds.
- You can use the concepts of electron arrangement and forces in atoms to explain the periodic trend of electronegativity.
- The electronegativity difference between elements can be used to predict what kinds of compounds the elements will form.
- Lewis structures can represent the formation of ionic and covalent compounds according to the octet rule.
- You can explain the conductivity of covalent and ionic compounds using an understanding of covalent and ionic bonding.
- By predicting the shapes of molecules, you can predict their polarity.
- The polarity of molecules can be used to explain the range of boiling points and melting points among compounds that contain molecules with similar masses.
- There is a methodical way to unambiguously name compounds and write their chemical formulas.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

chemical bonds	ionic bond
covalent bond	electronegativity
octet rule	isoelectronic
pure covalent bond	diatomic elements
double bond	triple bond
molecular compounds	intramolecular forces
intermolecular forces	metallic bond
alloy	polar covalent bond
lone pairs	bonding pairs
polar molecule	dipolar molecules
non-polar molecule	chemical formula
valence	polyatomic ions
zero sum rule	chemical nomenclature
binary compound	Stock system
tertiary compounds	

Knowledge/Understanding

1. Both electronegativity and electron affinity describe electron attraction. Explain how they are different.

2. Calculate ΔEN for each bond.
  (a) Zn—O
  (b) Mg—I
  (c) Co—Cl
  (d) N—O

3. Indicate whether each bond in question 2 is ionic or covalent.

4. Name three characteristics of covalent compounds and three characteristics of ionic compounds.

5. Give two examples of ionic compounds and two examples of covalent compounds.

6. How does a property of noble gases lead to the octet rule?

7. Draw a Lewis structure to represent each ionic compound.
  (a) potassium bromide
  (b) calcium fluoride
  (c) magnesium oxide
  (d) lithium oxide

8. Draw a Lewis structure to represent each covalent compound.
  (a) CO_2
  (b) NaH
  (c) NF_3

9. Describe, in detail, what happens when an ionic bond forms between calcium and chlorine. Use Lewis structures to illustrate your description.

10. Diatomic elements (such as oxygen, nitrogen, and chlorine) tend to exist at room temperature as gases. Explain why this is true using your understanding of bonding.

11. A solid covalent compound has both intermolecular and intramolecular forces. Do solid ionic compounds contain intermolecular forces? Explain your answer.

12. Distinguish between a non-polar covalent bond and a polar covalent bond.

13. Distinguish between an ionic bond and a polar covalent bond.

14. Without calculating ΔEN, arrange each set of bonds from most polar to least polar. Then calculate ΔEN for each bond to check your arrangement.
   (a) Mn—O, Mn—N, Mn—F
   (b) Be—F, Be—Cl, Be—Br
   (c) Ti—Cl, Fe—Cl, Cu—Cl, Ag—Cl, Hg—Cl
15. What kind of diagram would you use (Lewis structure, structural diagram, ball-and-stick model, or space-filling model) to illustrate each idea?
   (a) When sodium and chlorine form an ionic bond, sodium loses an electron and chlorine gains an electron.
   (b) Water, H_2O, is a molecule that has a bent shape.
   (c) Each oxygen atom in carbon dioxide, CO_2, has two bonding pairs and two non-bonding pairs.
   (d) Silicon tetrachloride, $SiCl_4$, is a tetrahedral molecule.
16. Write the valences for the elements in each compound. If the compound is ionic, indicate the charge that is associated with each valence.
   (a) AgCl (b) Mn_3P_2
   (c) PCl_5 (d) CH_4
   (e) TiO_2 (f) HgF_2
   (g) CaO (h) FeS
17. Write the formula of each compound.
   (a) tin(II) fluoride
   (b) barium sulfate
   (c) hydrogen cyanide
   (d) cesium bromide
   (e) ammonium hydrogen phosphate
   (f) sodium periodate
   (g) potassium bromate
   (h) sodium cyanate
18. Write the name of each compound.
   (a) HIO_2 (b) $KClO_4$
   (c) CsF (d) $NiCl_2$
   (e) $NaHSO_4$ (f) $Al_2(SO_3)_3$
   (g) $K_2Cr_2O_7$ (h) $Fe(IO_4)_3$
19. Name each compound in two different ways.
   (a) FeO (b) $SnCl_4$
   (c) $CuCl_2$ (d) $CrBr_3$
   (e) PbO_2 (f) HgO

Inquiry

20. Suppose that you have two colourless compounds. You know that one is an ionic compound and the other is a covalent compound. Design an experiment to determine which compound is which. Describe the tests you would perform and the results you would expect.
21. You have two liquids, A and B. You know that one liquid contains polar molecules, and the other liquid contains non-polar molecules. You do not know which is which, however. You pour each liquid so that it falls in a steady, narrow stream. As you pour, you hold a negatively charged ebonite rod to the stream. The stream of liquid A is deflected toward the rod. The rod does not affect the stream of liquid B. Which liquid is polar? Explain your answer.

Communication

22. Explain how you would predict the most common valences of the elements of the second period (Li, Be, B, C, N, O, F, and Ne) if you did not have access to the periodic table. Use Lewis structures to illustrate your explanation.
23. Create a concept map to summarize what you learned in this chapter about the nature of bonding and the ways in which bonding models help to explain physical and chemical properties.
24. Compare and contrast ionic bonding and metallic bonding. Include the following ideas:
   - Metals do not bond to other metals in definite ratios. Metals do bond to non-metals in definite ratios.
   - Solid ionic compounds do not conduct electricity, but solid metals do.
25. Explain why it was important for chemists worldwide to decide on a system for naming compounds.

Making Connections

26. Chemists do not always agree on names, not just for compounds but even for elements. As new elements are synthesized in laboratories, they must be named. Until 1997, there was a controversy over the names of elements 104 to

109 (called the *transfermium* elements). The periodic table at the back of this textbook gives the names that have now been accepted. Do some research to find out what other names were proposed for those elements. Find out what justification was given for the alternative names and the accepted names. Then write an essay in which you evaluate the choice that was made. Do you agree or disagree? Justify your opinion.

Answers to Practice Problems and Short Answers to Section Review Questions:

Practice Problems: **1.(a)** 1.24, covalent **(b)** 0.50, covalent **(c)** 1.85, ionic **(d)** 1.94, ionic **(e)** 1.78, ionic **(f)** 0.49, covalent **(g)** 1.73, ionic **(h)** 2.03, ionic **2.(a)** 2.44, ionic **(b)** 2.34, ionic **(c)** 3.16, ionic **(d)** 3.00, ionic **(e)** 1.98, ionic **(f)** 2.55, ionic **3.(a)** one calcium atom gives up two electrons to one oxygen atom **(b)** one potassium atom gives up one electron to one chlorine atom **(c)** one potassium atom gives up one electron to one fluorine atom **(d)** one lithium atom gives up one electron to one fluorine atom **(e)** one lithium atom gives up one electron to one bromine atom **(f)** one barium atom gives up two electrons to one oxygen atom **4.(a)** 1.85 **(b)** 2.16 **(c)** 2.46 **(d)** 2.51 **(e)** 1.76 **(f)** 1.96 **5.(a)** one magnesium atom gives up one electron to each of two chlorine atoms **(b)** one calcium atom gives up one electron to each of two chlorine atoms **(c)** two lithium atoms each give up one electron to one oxygen atom **(d)** two sodium atoms each give up one electron to one oxygen atom **(e)** two potassium atoms each give up one electron to one sulfur atom **(f)** one calcium atom gives up one electron to each of two bromine atoms **6.(a)** Each iodine atom has seven electrons. Two iodine atoms bonded together share one pair of electrons so each has access to eight electrons. **(b)** Each bromine atom has seven electrons. Two bromine atoms bonded together share one pair of electrons so that each has access to eight electrons. **(c)** Each hydrogen atom has one electron. Two hydrogen atoms bonded together share one pair of electrons so that each has access to two electrons. **(d)** Each fluorine atom has seven electrons. Two fluorine atoms bonded together share one pair of electrons so that each has access to eight electrons. **7.(a)** One hydrogen atom bonds to one oxygen atom, sharing one electron pair. **(b)** Two chlorine atoms bond to one oxygen atom. Each chlorine atom shares one pair of electrons with the oxygen atom. **(c)** One carbon atom bonds to four hydrogen atoms. Each hydrogen atom shares one pair of electrons with the carbon atom. **(d)** One iodine atom bonds to one hydrogen atom. They share an electron pair. **(e)** One nitrogen atom bonds to three hydrogen atoms. Each hydrogen atom shares a pair of electrons with the nitrogen atom. **(f)** One hydrogen atom bonds to one rubidium atom. They share a pair of electrons. **8.** The carbon atom shares two pairs of electrons with each sulfur atom, so it has two double bonds. **9.** The carbon atom shares one pair of electrons with hydrogen, and three pairs of electrons with the nitrogen atom. **10.** The two carbon atoms share three pairs of electrons in a triple bond. Each carbon atom shares one pair of electrons with one hydrogen atom. **11.(a)** polar covalent **(b)** covalent **(c)** covalent **(d)** polar covalent **(e)** covalent **(f)** ionic **(g)** polar covalent **(h)** ionic **12.(a)** C δ^+, F δ^- **(d)** Cu δ^+, O δ^- **(g)** Fe δ^+, O δ^- **13.(a)** O—O, N—O, H—Cl, Na—Cl **(b)** N—N, P—O, C—Cl, Mg—C **14.** Group 13, 3, Group 16, –2, Group 17, –1 **15.** 0, do not tend to gain or lose electrons **16.(a)** Na_2S **(b)** CaS **(c)** BaS **(d)** Al_2S_3 **(e)** Rb_2S **(f)** H_2S **17.(a)** CaO **(b)** CaS **(c)** $CaCl_2$ **(d)** $CaBr_2$ **(e)** Ca_3P_2 **(f)** CaF_2 **18.(a)** $NaNO_3$ **(b)** Na_3PO_4 **(c)** Na_2SO_3 **(d)** $NaCH_3COO$ **(e)** $Na_2S_2O_3$ **(f)** Na_2CO_3 **19.** $Mg(NO_3)_2$ **(b)** $Mg_3(PO_4)_2$ **(c)** $MgSO_3$ **(d)** $Mg(CH_3COO)_2$ **(e)** MgS_2O_3 **20.(a)** aluminum oxide **(b)** calcium bromide **(c)** sodium phosphide **(d)** copper(I) sulfide **(e)** magnesium nitride **(f)** mercury(II) iodide **21.(a)** FeS **(b)** SnO (c) CrO **(d)** $CoCl_2$ **(e)** MnI_2 **(f)** ZnO **22.(a)** hydrogen selenide **(b)** hydrogen chloride **(c)** hydrogen fluoride **(d)** lithium hydride **(e)** calcium hydride **(f)** phosphorus(III) hydride **23.(a)** ammonium sulfite **(b)** aluminum nitrite **(c)** lithium carbonate **(d)** nickel(II) hydroxide **(e)** silver phosphate **(f)** copper(II) acetate **24.(a)** sulfur hexafluoride **(b)** dinitrogen pentoxide **(c)** phosphorus pentachloride **(d)** carbon tetrafluoride

Section Review: **3.1:** **4.(a)** Li, La, Zn, Si, Br **(b)** Cs, Y, Ga, P, Cl **5.(a)** 0.40, covalent **(b)** 1.89, ionic **(c)** 0.96, covalent **(d)** 2.16, ionic **6.(a)** low **(b)** covalent **3.2:** **1.(a)** one magnesium atom gives up one electron to each of two fluorine atoms **(b)** one potassium electron gives up one electron to one bromine atom **(c)** one rubidium atom gives up one electron to one chlorine atom **(d)** one calcium atom gives up two electrons to one oxygen atom **2.(a)** The hydrogen atom and chlorine atoms all bond to the carbon atom. **(b)** Each hydrogen and chlorine atom shares one electron pair with the carbon atom. **(c)** The two nitrogen atoms share three pairs of electrons in a triple covalent bond. **3.(a)** covalent **(b)** covalent **(c)** covalent **(d)** ionic **3.3:** **1.(a)** 1.94, ionic **(b)** 0.35, covalent **(c)** 2.23, ionic **(d)** 1.54, polar covalent **(e)** 0.86, polar covalent **(f)** 0.61, polar covalent **2.(d)** Si δ^+, O δ^- **(e)** S δ^+, O δ^- **(f)** C δ^+, Cl δ^- **4.(a)** O—F, H—Br, H—Cl, K—Br **(b)** C—H, C—Br, C—O, C—F **7.(a)** 1.26, non-polar molecule **(b)** 0.97, polar molecule **(c)** 1.55, non-polar molecule **3.4:** **1.(a)** potassium chromate **(b)** ammonium nitrate **(c)** sodium sulfate **(d)** strontium phosphate **(e)** potassium nitrite **(f)** barium hypochlorite **2.(a)** magnesium chloride **(b)** sodium oxide **(c)** iron(III) chloride **(d)** copper(II) oxide **(e)** zinc sulfide **(f)** aluminum bromide **3.(a)** $NaHCO_3$ **(b)** $K_2Cr_2O_7$ **(c)** NaClO **(d)** LiOH **(e)** $KMnO_4$ **(f)** NH_4Cl **(g)** $Ca_3(PO_4)_2$ **(h)** $Na_2S_2O_3$ **4.(a)** any two of: vanadium(II) oxide, VO, vanadium(III) oxide, V_2O_3, vanadium(IV) oxide, VO_2, vanadium(V) oxide, V_2O_5 **(b)** iron(II) sulfide, FeS, iron(I) sulfide, Fe_2S **(c)** nickel(II) oxide, NiO, nickel(III) oxide, Ni_2O_3