Name _____ Date _____ Class _____

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 1

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

1. a A chemical bond between atoms results from the attraction between the valence electrons and ____ of different atoms.

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

2. b A covalent bond consists of

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

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

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

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

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

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

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

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

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

7. a The greater the electronegativity difference between two atoms bonded together, the greater the bond's percentage of

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

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

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

Name _____ Date _____ Class _____

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 polar covalent.

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

12. A bond's character is more than 50% ionic if the electronegativity difference between the two atoms is greater than 1.7.

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

Answers will vary.

H_2 **a.** nonpolar covalent compound

HCl **b.** polar covalent compound

NaCl **c.** ionic compound

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

A pair of electrons is attracted to both nuclei of the two atoms bonded together.

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

electronegativity

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

The difference between the electronegativity of the two atoms in a bond will determine whether the bond is ionic or covalent. If the difference in electronegativity is greater than 1.7, the bond is considered ionic.

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

The electron density is greater around the more electronegative atom, giving that part of the compound a partial negative charge. The other part of the compound has an equal partial positive charge.

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.

As the atoms involved in the formation of a covalent bond approach each other, the electron-proton attraction is stronger than the electron-electron and proton-proton repulsions. The atoms are drawn to each other and their potential energy decreases. Eventually, a distance is reached at which the repulsions between the like charges equals the attraction of the opposite charges. At this point, potential energy is at a minimum and a stable molecule forms.

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

Choose from hydrogen, boron, beryllium, phosphorus, sulfur, and xenon.

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

Resonance structures show that one Lewis structure cannot correctly represent the location of electrons in a bond. Resonance structures show delocalized electrons, while Lewis structures depict electrons in a definite location.

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

H—F, H—Cl, H—Br, H—I

Name _____ Date _____ Class _____

SECTION 2 continued

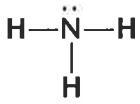
b.

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

C≡C, C=C, C—C

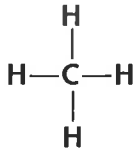
5. Draw Lewis structures to represent each of the following formulas:

a. NH_3

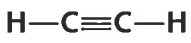


b. H_2O

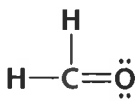
c. CH_4



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. a The notation for sodium chloride, NaCl, stands for one

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

2. d 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. b 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. b 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. b 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.

The force of attraction between unlike charges holds a negative ion and a positive ion together in an ionic bond.

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

lattice energy

Name _____ Date _____ Class _____

SECTION 3 continued

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

The atoms in a polyatomic ion are held together with covalent bonds, but polyatomic ions combine with ions of opposite charge to form ionic compounds.

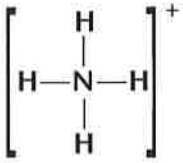
9. Arrange the ionic bonds in the table below in order of increasing strength from weakest to strongest.

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

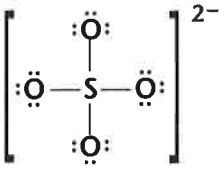
KCl, NaCl, LiCl, CaO, MgO

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. b 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. a 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. d 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. d Mobile electrons in the metallic bond are responsible for

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

5. a 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. c 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.

Most metals have their outer electrons in *s* orbitals, while nonmetals have their outer electrons in *p* orbitals.

Name _____ Date _____ Class _____

SECTION 4 continued

8. How does the behavior of electrons in metals contribute to the metal's ability to conduct electricity and heat?

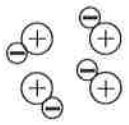
The mobility of electrons in a network of metal atoms contributes to the metal's ability to conduct electricity and heat.

9. What is the relationship between the enthalpy of vaporization of a metal and the strength of the bonds that hold the metal together?

The amount of energy required to vaporize a metal is a measure of the strength of the bonds that hold the metal together. The greater a metal's enthalpy of vaporization, the stronger the metallic bond.

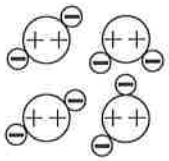
10. Draw two diagrams of a metallic bond. In the first diagram, draw a weak metallic bond; in the second, show a metallic bond that would be stronger. Be sure to include nuclear charge and number of electrons in your illustrations.

a.



weak bond

b.



strong bond

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	atoms	ions
Overall charge	neutral	neutral
Conductive in the solid state	yes	no
Melting point	low to high	high
Hardness	soft to hard	hard
Malleable	yes	no
Ductile	yes	no

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.

Pairs of valence electrons repel one another.

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

The electron pairs that are not involved in bonding also take up space, creating a tetrahedron of electron pairs and making the water molecule angular or bent.

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

the *s* orbital and all three *p* orbitals from the second energy level

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

electronegativity difference and molecular geometry or unshared electron pairs

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

3 hydrogen bonding

4 ionic

2 dipole-dipole

1 London dispersion

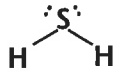
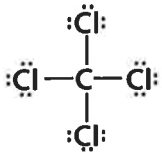
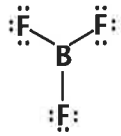
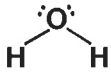
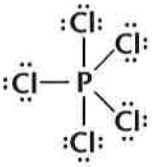
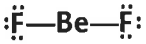
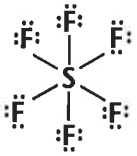
6. How are dipole-dipole attractions, London dispersion forces, and hydrogen bonding similar?

They are all forces of attraction between molecules. In all cases there is an attraction between the slightly negatively-charged portion of one molecule and the slightly positively charged portion of another molecule.

Name _____ Date _____ Class _____

SECTION 5 continued

7. Complete the following table:

Formula	Lewis structure	Geometry	Polar
H_2S		bent	yes
CCl_4		tetrahedral	no
BF_3		trigonal planar	no
H_2O		bent	yes
PCl_5		trigonal bipyramidal	no
BeF_2		linear	no
SF_6		octahedral	no

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:

lattice energy — **a.** ionic bond

bond energy — **b.** covalent bond

enthalpy of vaporization — **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.

ionic — **a.** H—F

ionic — **b.** Na—Cl

polar covalent — **c.** H—O

nonpolar covalent — **d.** H—H

polar covalent — **e.** H—C

polar covalent — **f.** H—N

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

A hydrogen bond is a dipole-dipole attraction between a partially positive hydrogen atom and the unshared electron pair of a strongly electronegative atom such as O, N, or F. Unlike ionic or covalent bonds, in which electrons are given up or shared, the hydrogen bond is a weaker attraction. Hydrogen bonds are generally intermolecular, while ionic and covalent bonds occur between ions or atoms respectively.

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.

Oxygen has higher electronegativity than sulfur, which creates a highly polar bond. Increased polarity in H_2O bonds means a stronger intermolecular attraction, making water a liquid at room temperature. Hydrogen bonding exists between water molecules, but not between hydrogen sulfide molecules.

Name _____ Date _____ Class _____

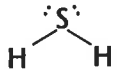
MIXED REVIEW continued

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

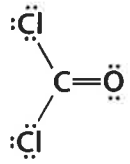
There is a difference between the electronegativities of the two atoms in both types of bonds that results in electrons being more closely associated with the more electronegative atom.

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

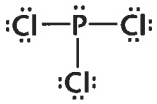
polar **a.** H_2S



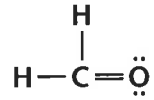
polar **b.** $COCl_2$



polar **c.** PCl_3



polar **d.** CH_2O



Name _____ Date _____ Class _____

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 1

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

1. c In a Stock system name such as iron(III) sulfate, the Roman numeral tells us

(a) how many atoms of Fe are in one formula unit.
(b) how many sulfate ions can be attached to the iron atom.
(c) the charge on each Fe ion.
(d) the total positive charge of the formula unit.

2. c Changing a subscript in a correctly written chemical formula

(a) changes the number of moles represented by the formula.
(b) changes the charges on the other ions in the compound.
(c) changes the formula so that it no longer represents the compound it previously represented.
(d) has no effect on the formula.

3. The explosive TNT has the molecular formula $C_7H_5(NO_2)_3$.

4 elements **a.** How many elements make up this compound?

6 oxygen atoms **b.** How many oxygen atoms are present in one molecule of $C_7H_5(NO_2)_3$?

21 atoms **c.** How many atoms in total are present in one molecule of $C_7H_5(NO_2)_3$?

4.2×10^{24} atoms **d.** How many atoms are present in a sample of 2.0×10^{23} molecules of $C_7H_5(NO_2)_3$?

4. How many atoms are present in each of these formula units?

11 atoms **a.** $Ca(HCO_3)_2$

45 atoms **b.** $C_{12}H_{22}O_{11}$

10 atoms **c.** $Fe(ClO_2)_3$

9 atoms **d.** $Fe(ClO_3)_2$

5. N_2O_5 **a.** What is the formula for the compound dinitrogen pentoxide?

iron(II) oxide **b.** What is the Stock system name for the compound FeO?

H_2SO_3 **c.** What is the formula for sulfurous acid?

phosphoric acid **d.** What is the name for the acid H_3PO_4?

Name _____ Date _____ Class _____

SECTION 1 continued

6. Some binary compounds are ionic, others are covalent. The type of bond favored partially depends on the position of the elements in the periodic table. Label each of these claims as True or False; if False, specify the nature of the error.

a. Covalently bonded binary molecular compounds are typically composed of nonmetals.

True

b. Binary ionic compounds are composed of metals and nonmetals, typically from opposite sides of the periodic table.

True

7. Refer to **Table 2** on page 226 of the text and **Table 5** on page 230 of the text for examples of names and formulas for polyatomic ions and acids.

a. Derive a generalization for determining whether an acid name will end in the suffix *-ic* or *-ous*.

In general, if the anion name ends in *-ate*, the corresponding acid name will end in a suffix of *-ic*. In general, if the anion name ends in *-ite*, the corresponding acid name will end in a suffix of *-ous*.

b. Derive a generalization for determining whether an acid name will begin with the prefix *hydro-* or not.

In general, if the anion name ends in *-ide*, the corresponding acid name will end in a suffix of *-ic* and begin with a prefix of *hydro-*. The prefix *hydro-* is never used for anions ending in *-ate* or *-ite*.

8. Fill in the blanks in the table below.

Compound name	Formula
Aluminum sulfide	Al_2S_3
Cesium carbonate	Cs_2CO_3
Lead(II) chloride	$PbCl_2$
Ammonium phosphate	$(NH_4)_3PO_4$
Hydroiodic acid	**HI**

Name _____ Date _____ Class _____

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 2

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

1. Assign the oxidation number to the specified element in each of the following examples:

+4 **a.** S in H_2SO_3

+6 **b.** S in $MgSO_4$

−2 **c.** S in K_2S

+1 **d.** Cu in Cu_2S

+6 **e.** Cr in Na_2CrO_4

+5 **f.** N in HNO_3

+4 **g.** C in $(HCO_3)^-$

−3 **h.** N in $(NH_4)^+$

2. SCl_2 **a.** What is the formula for the compound sulfur(II) chloride?

nitrogen(IV) oxide **b.** What is the Stock system name for NO_2?

3. fluorine **a.** Use electronegativity values to determine the one element that always has a negative oxidation number when it appears in any binary compound.

0; F_2 **b.** What is the oxidation number and formula for the element described in part **a** when it exists as a pure element?

4. Tin has possible oxidation numbers of +2 and +4 and forms two known oxides. One of them has the formula SnO_2.

tin(IV) oxide **a.** Give the Stock system name for SnO_2.

SnO **b.** Give the formula for the other oxide of tin.

5. Scientists think that two separate reactions contribute to the depletion of the ozone, O_3, layer. The first reaction involves oxides of nitrogen. The second involves free chlorine atoms. The equations that represent the reactions follow. When a compound is not stated as a formula, write the correct formula in the blank beside its name.

a. NO (nitrogen monoxide) + $O_3 \rightarrow$ NO_2 (nitrogen dioxide) + O_2

Name _____ Date _____ Class _____

SECTION 2 continued

b. $Cl + O_3 \rightarrow$ ClO (chlorine monoxide) $+ O_2$

6. Consider the covalent compound dinitrogen trioxide when answering the following:

N_2O_3 **a.** What is the formula for dinitrogen trioxide?

+3 **b.** What is the oxidation number assigned to each nitrogen atom in this compound? Explain your answer.

The three oxygen atoms have oxidation states of −6 total, and because the algebraic sum of the oxidation states in a neutral compound must be zero, the two nitrogen atoms must have oxidation states of +6 total, therefore +3 each.

nitrogen(III) oxide **c.** Give the Stock name for dinitrogen trioxide.

7. The oxidation numbers assigned to the atoms in some organic compounds have unexpected values. Assign oxidation numbers to each atom in the following compounds: (Note: Some oxidation numbers may not be whole numbers.)

a. CO_2

Carbon is +4 and each oxygen is −2.

b. CH_4 (methane)

Carbon is −4 and each hydrogen is +1.

c. $C_6H_{12}O_6$ (glucose)

Each carbon is 0, each hydrogen is +1, and each oxygen is −2.

d. C_3H_8 (propane gas)

Each carbon is −8/3 and each hydrogen is +1.

8. Assign oxidation numbers to each element in the compounds found in the following situations:

a. Rust, Fe_2O_3, forms on an old nail.

Each iron is +3 and each oxygen is −2.

b. Nitrogen dioxide, NO_2, pollutes the air as a component of smog.

Nitrogen is +4 and each oxygen is −2.

c. Chromium dioxide, CrO_2, is used to make recording tapes.

Chromium is +4 and each oxygen is −2.

Name _____ Date _____ Class _____

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 3

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

1. Label each of the following statements as True or False:

True — **a.** If the formula mass of one molecule is x amu, the molar mass is x g/mol.

False — **b.** Samples of equal numbers of moles of two different chemicals must have equal masses as well.

True — **c.** Samples of equal numbers of moles of two different molecular compounds must have equal numbers of molecules as well.

2. How many moles of each element are present in a 10.0 mol sample of $Ca(NO_3)_2$?

10 mol of calcium, 20 mol of nitrogen, 60 mol of oxygen

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

3. Consider a sample of 10.0 g of the gaseous hydrocarbon C_3H_4 to answer the following questions.

0.250 mol — **a.** How many moles are present in this sample?

1.50×10^{23} molecules — **b.** How many molecules are present in the C_3H_4 sample?

4.51×10^{23} carbon atoms — **c.** How many carbon atoms are present in this sample?

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SECTION 3 continued

10.1% **d.** What is the percentage composition of hydrogen in the sample?

4. One source of aluminum metal is alumina, Al_2O_3.

52.9% **a.** Determine the percentage composition of Al in alumina.

2100 lb **b.** How many pounds of aluminum can be extracted from 2.0 tons of alumina?

5. Compound A has a molar mass of 20 g/mol, and compound B has a molar mass of 30 g/mol.

20 g **a.** What is the mass of 1.0 mol of compound A, in grams?

0.17 mol **b.** How many moles are present in 5.0 g of compound B?

4.0 mol **c.** How many moles of compound B are needed to have the same mass as 6.0 mol of compound A?

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CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 4

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

1. Write empirical formulas to match the following molecular formulas:

CH_3O_2 **a.** $C_2H_6O_4$

N_2O_5 **b.** N_2O_5

HgCl **c.** Hg_2Cl_2

CH_2 **d.** C_6H_{12}

2. C_4H_8 A certain hydrocarbon has an empirical formula of CH_2 and a molar mass of 56.12 g/mol. What is its molecular formula?

3. A certain ionic compound is found to contain 0.012 mol of sodium, 0.012 mol of sulfur, and 0.018 mol of oxygen.

$Na_2S_2O_3$ **a.** What is its empirical formula?

neither **b.** Is this compound a sulfate, sulfite, or neither?

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

4. Water of hydration was discussed in **Sample Problem K** on pages 243–244 of the text. Strong heating will drive off the water as a vapor in hydrated copper(II) sulfate. Use the data table below to answer the following:

Mass of the empty crucible	4.00 g
Mass of the crucible plus hydrate sample	4.50 g
Mass of the system after heating	4.32 g
Mass of the system after a second heating	4.32 g

36% **a.** Determine the mass percentage of water in the original sample.

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SECTION 4 continued

5 **b.** The compound has the formula $CuSO_4 \cdot xH_2O$. Determine the value of x.

c. What might be the purpose of the second heating?

The second heating is to ensure that all the water in the sample has been driven off. If the mass is less after the second heating, water was still present after the first heating.

5. Gas X is found to be 24.0% carbon and 76.0% fluorine by mass.

CF_2 **a.** Determine the empirical formula of gas X.

C_4F_8 **b.** Given that the molar mass of gas X is 200.04 g/mol, determine its molecular formula.

6. A compound is found to contain 43.2% copper, 24.1% chlorine, and 32.7% oxygen by mass.

$CuClO_3$ **a.** Determine its empirical formula.

b. What is the correct Stock system name of the compound in part **a**?

copper(I) chlorate

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CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

MIXED REVIEW

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

1. Write formulas for the following compounds:

$CuCO_3$ **a.** copper(II) carbonate

Na_2SO_3 **b.** sodium sulfite

$(NH_4)_3PO_4$ **c.** ammonium phosphate

SnS_2 **d.** tin(IV) sulfide

HNO_2 **e.** nitrous acid

2. Write the Stock system names for the following compounds:

magnesium perchlorate **a.** $Mg(ClO_4)_2$

iron(II) nitrate **b.** $Fe(NO_3)_2$

iron(III) nitrite **c.** $Fe(NO_2)_3$

cobalt(II) oxide **d.** CoO

nitrogen(V) oxide **e.** dinitrogen pentoxide

3. 13 atoms **a.** How many atoms are represented by the formula $Ca(HSO_4)_2$?

4.0 mol **b.** How many moles of oxygen atoms are in a 0.50 mol sample of this compound?

+6 **c.** Assign the oxidation number to sulfur in the HSO_4^- anion.

4. Assign the oxidation number to the element specified in each of the following:

+1 **a.** hydrogen in H_2O_2

−1 **b.** hydrogen in MgH_2

0 **c.** sulfur in S_8

+4 **d.** carbon in $(CO_3)^{2-}$

+6 **e.** chromium in $Na_2Cr_2O_7$

+4 **f.** nitrogen in NO_2

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MIXED REVIEW continued

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

5. c, b, d, a

Following are samples of four different compounds. Arrange them in order of increasing mass, from smallest to largest.

a. 25 g of oxygen gas
b. 1.00 mol of H_2O
c. 3×10^{23} molecules of C_2H_6
d. 2×10^{23} molecules of $C_2H_6O_2$

6. NaOH

a. What is the formula for sodium hydroxide?

40.00 g/mol

b. What is the formula mass of sodium hydroxide?

10. g

c. What is the mass in grams of 0.25 mol of sodium hydroxide?

7. 80% C, 20% H

What is the percentage composition of ethane gas, C_2H_6, to the nearest whole number?

8. $C_5H_{10}O_5$

Ribose is an important sugar (part of RNA),with a molar mass of 150.15 g/mol. If its empirical formula is CH_2O, what is its molecular formula?

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MIXED REVIEW continued

9. Butane gas, C_4H_{10}, is often used as a fuel.

174 g **a.** What is the mass in grams of 3.00 mol of butane?

1.81×10^{24} molecules **b.** How many molecules are present in that 3.00 mol sample?

C_2H_5 **c.** What is the empirical formula of the gas?

10. $C_{10}H_8$ Naphthalene is a soft covalent solid that is often used in mothballs. Its molar mass is 128.18 g/mol and it contains 93.75% carbon and 6.25% hydrogen. Determine the molecular formula of napthalene from this information.

11. Nicotine has the formula $C_xH_yN_z$. To determine its composition, a sample is burned in excess oxygen, producing the following results:

1.0 mol of CO_2
0.70 mol of H_2O
0.20 mol of NO_2

Assume that all the atoms in nicotine are present as products.

1.0 mol **a.** Determine the number of moles of carbon present in the products of this combustion.

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MIXED REVIEW continued

1.40 mol **b.** Determine the number of moles of hydrogen present in the combustion products.

0.20 mol **c.** Determine the number of moles of nitrogen present in the combustion products.

C_5H_7N **d.** Determine the empirical formula of nicotine based on your calculations.

162 g/mol **e.** In a separate experiment, the molar mass of nicotine is found to be somewhere between 150 and 180 g/mol. Calculate the molar mass of nicotine to the nearest gram.

12. When $MgCO_3(s)$ is strongly heated, it produces solid MgO as gaseous CO_2 is driven off.

52.2% **a.** What is the percentage loss in mass as this reaction occurs?

Mg is +2, C is +4, and O is −2 **b.** Assign the oxidation number to each atom in $MgCO_3$.

No **c.** Does the oxidation number of carbon change upon the formation of CO_2?

5 The Perdioc Law

Section: History of the Periodic Table

1. d	**2.** a
3. d	**4.** a
5. b	**6.** c
7. c	**8.** b
9. b	**10.** d

Section: Electron Configuration and the Periodic Table

1. b	**2.** d
3. b	**4.** a
5. b	**6.** b
7. b	**8.** d
9. c	**10.** c

Section: Electron Configuration and Periodic Properties

1. a	**2.** c
3. c	**4.** a
5. b	**6.** a
7. c	**8.** a
9. a	**10.** d

6 Chemical Bonding

Section: Introduction to Chemical Bonding

1. c	**2.** b
3. b	**4.** a
5. a	**6.** b
7. d	**8.** c
9. a	**10.** b

Section: Covalent Bonding and Molecular Compounds

1. c	**2.** c
3. c	**4.** b
5. a	**6.** b
7. d	**8.** d
9. a	**10.** a

Section: Ionic Bonding and Ionic Compounds

1. d	**2.** c
3. c	**4.** a
5. d	**6.** d
7. d	**8.** b
9. a	**10.** b

Section: Metallic Bonding

1. d	**2.** a
3. c	**4.** c
5. b	**6.** a
7. d	**8.** b
9. c	**10.** b

Section: Molecular Geometry

1. d	**2.** a
3. b	**4.** d
5. b	**6.** c
7. b	**8.** c
9. a	**10.** d

7 Chemical Formulas and Chemical Compounds

Section: Chemical Names and Formulas

1. b	**2.** c
3. c	**4.** a
5. d	**6.** c
7. b	**8.** a
9. d	**10.** c

Section: Oxidation Numbers

1. c	**2.** b
3. a	**4.** c
5. b	**6.** d
7. d	**8.** b
9. c	**10.** a

Section: Using Chemical Formulas

1. a	**2.** a
3. d	**4.** b
5. d	**6.** c
7. a	**8.** c
9. b	**10.** b

Section: Determining Chemical Formulas

1. c	**2.** a
3. b	**4.** d
5. c	**6.** c
7. a	**8.** c
9. b	**10.** b

8 Chemical Equations and Reactions

Section: Describing Chemical Reactions

1. b	**2.** d
3. d	**4.** c
5. c	**6.** d
7. d	**8.** c
9. d	**10.** d