Name _____ Date _____ Class _____

 10 | # CHEMICAL QUANTITIES

SECTION 10.1 THE MOLE: A MEASUREMENT OF MATTER (pages 287–296)

This section defines the mole and explains how the mole is used to measure matter. It also teaches you how to calculate the mass of a mole of any substance.

▶ Measuring Matter (pages 287–289)

1. What do the questions "how much?" and "how many?" have in common?

_____

_____

2. List two or three ways to measure matter.

_____

▶ What Is a Mole? (pages 290–293)

3. Circle the letter of the term that is an SI unit for measuring the amount of a substance.

    a. dozen   **b.** ounce   **c.** pair   **d.** mole

4. What is Avogadro's number?

_____

5. Circle the letter of the term that is NOT a representative particle of a substance.

    a. molecule   **b.** atom   **c.** grain   **d.** formula unit

6. List the representative particle for each of the following types of substances.

    a. molecular compounds   _____

    b. ionic compounds   _____

    c. elements   _____

7. Is the following sentence true or false? To determine the number of

representative particles in a compound, you count the molecules by viewing

them under a microscope. _____

8. How can you determine the number of atoms in a mole of a molecular compound?

_____

_____

CHAPTER 10, Chemical Quantities *(continued)*

9. Complete the table about representative particles and moles.

Representative Particles and Moles			
	Representative Particle	**Chemical Formula**	**Representative Particles in 1.00 mol**
Atomic oxygen		0	
Oxygen gas	Molecule		
Sodium ion			
Sodium chloride			

▶ The Mass of a Mole of an Element (pages 293–294)

10. What is the atomic mass of an element?

_____

11. Circle the letter of the phrase that completes this sentence correctly.
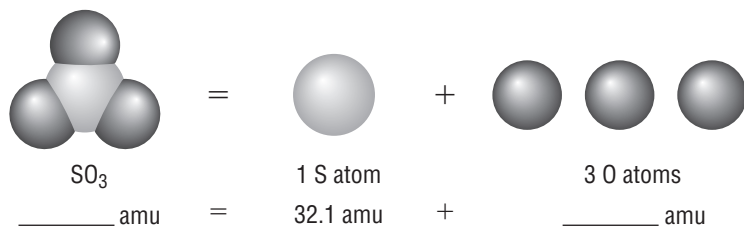The atomic masses of all elements

    a. are the same.

    b. are based on the mass of the carbon isotope C-12.

    c. are based on the mass of a hydrogen atom.

▶ The Mass of a Mole of a Compound (pages 295–296)

12. How do you determine the mass of a mole of a compound?

_____

_____

_____

13. Complete the labels on the diagram below.

SO_3 =   1 S atom   +   3 O atoms

_____ amu   =   32.1 amu   +   _____ amu

Name _____ Date _____ Class _____

14. What is the molar mass of a compound?

    _____

15. Is the following sentence true or false? Molar masses can be calculated

    directly from atomic masses expressed in grams. _____

SECTION 10.2 MOLE–MASS AND MOLE–VOLUME RELATIONSHIPS (pages 297–303)

This section explains how to use molar mass and molar volume to convert among measurements of mass, volume, and number of particles.

▶ **The Mole–Mass Relationship (pages 297–299)**

1. What is the molar mass of a substance?

   _____

2. What is the molar mass of KI (potassium iodide)?

   _____

▶ **The Mole–Volume Relationship (pages 300–302)**

3. Is the following sentence true or false? The volumes of one mole of different

   solid and liquid substances are the same. _____

4. Circle the letter of each term that can complete this sentence correctly.
   The volume of a gas varies with a change in

   a. temperature.   **c.** pressure.

   b. the size of the container.   **d.** the amount of light in the container.

5. Circle the letter of the temperature that is defined as standard temperature.

   a. 0 K   **c.** 0°C

   b. 100 K   **d.** 100°C

6. Is the following sentence true or false? Standard pressure is 101.3 kPa or

   1 atmosphere (atm). _____

7. What is the molar volume of a gas at standard temperature and pressure

   (STP)? _____

8. What units do you normally use to describe the density of a gas?

   _____

CHAPTER 10, Chemical Quantities *(continued)*

9. What is Avogadro's hypothesis?

_____

_____

10. Look at Figure 10.9 on page 300 to help you answer this question. Why is Avogadro's hypothesis reasonable?

_____

_____

_____

11. How many gas particles occupy a volume of 22.4 L at standard temperature

and pressure? _____

▶ The Mole Road Map (page 303)

12. The figure below shows how to convert from one unit to another unit. Write the missing conversion factors below.
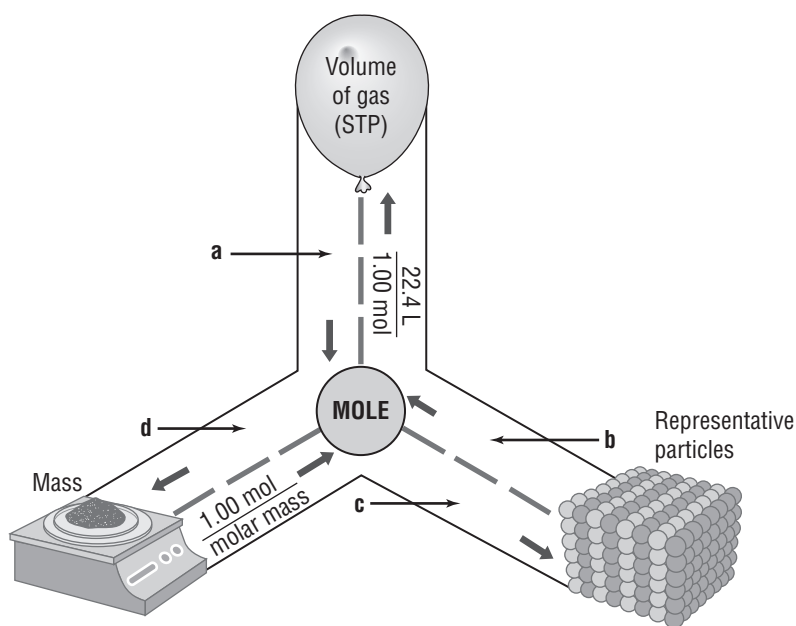
a. _____   **c.** _____

b. _____   **d.** _____

SECTION 10.3 PERCENT COMPOSITION AND CHEMICAL FORMULAS (pages 305–312)

This section explains how to calculate percent composition from chemical formulas or experimental data, and how to derive empirical and molecular formulas.

▶ Percent Composition of a Compound (pages 305–308)

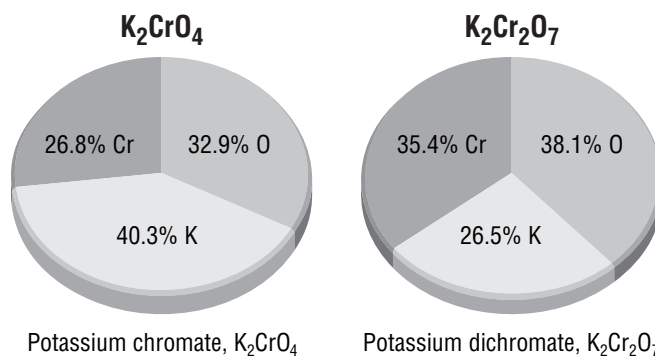1. How do you express relative amounts of each element in a compound?

   _____

2. Circle the letter of the phrase that completes this sentence correctly. The number of percent values in the percent composition of a compound is

   a. half as many as there are different elements in the compound.

   b. as many as there are different elements in the compound.

   c. twice as many as there are different elements in the compound.

3. What is the formula for the percent by mass of an element in a compound?

   _____

4. In the diagram below, which compound has a greater percent composition of

   chromium? _____

   How much greater is this percent? _____

K_2CrO_4   **$K_2Cr_2O_7$**

26.8% Cr   32.9% O   35.4% Cr   38.1% O

40.3% K   26.5% K

Potassium chromate, K_2CrO_4 Potassium dichromate, $K_2Cr_2O_7$

5. To calculate the percent composition of a known compound, start with the

   chemical formula of the compound and calculate the _____ ,

   which gives the mass of one mole of the compound.

6. Is the following sentence true or false? You can use percent composition

   to calculate the number of grams of an element in a given amount of a
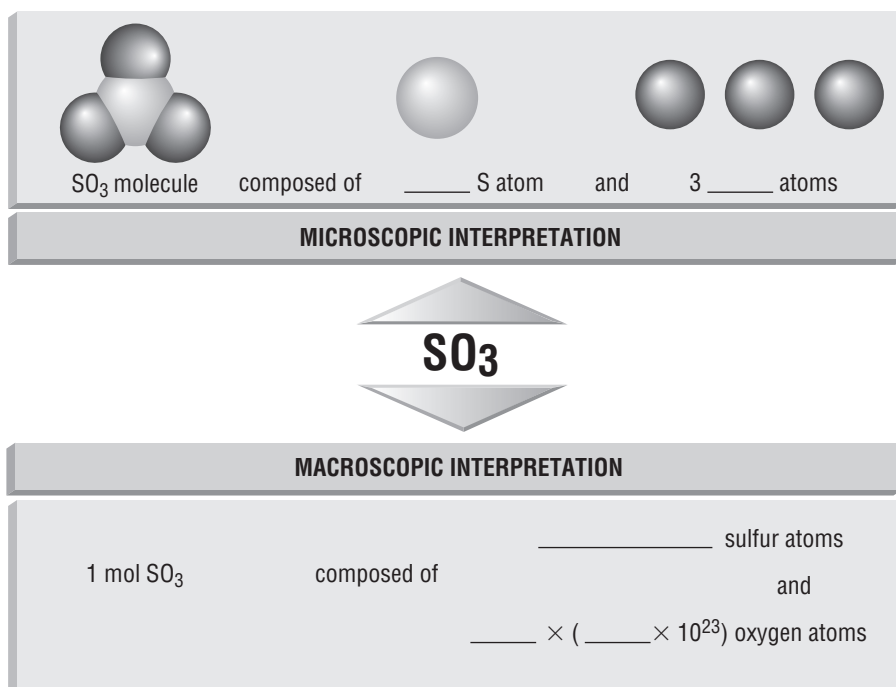
   compound. _____

CHAPTER 10, Chemical Quantities *(continued)*

7. How do you calculate the grams of an element in a specific amount of a compound?

_____

_____

▶ **Empirical Formulas** (pages 309–310)

8. An empirical formula of a compound gives the _____

whole-number ratio of the atoms of the elements in a compound.

9. Is the following sentence true or false? The empirical formula of a compound

is always the same as the molecular formula. _____

10. Look at Figure 10.16 and Table 10.3. Name three compounds that have an empirical formula of CH.

_____

11. Fill in the labels on the diagram below.

SO_3 molecule   composed of   _____ S atom   and   3 _____ atoms

MICROSCOPIC INTERPRETATION

SO_3

MACROSCOPIC INTERPRETATION

1 mol SO_3 composed of   _____ sulfur atoms

and

_____ × (_____ × 10^{23}) oxygen atoms

▶ Molecular Formulas (pages 311–312)

12. The molecular formula of a compound is either the same as its empirical

formula or a _____ of it.

13. What do you need to know to calculate the molecular formula of a compound?

_____

14. If you divide the molar mass of a compound by the empirical formula mass,
what is the result?

_____

15. What factor would you use to convert the empirical formula of a compound to
a molecular formula?

_____

📖 Reading Skill Practice

By looking carefully at photographs and illustrations in textbooks, you can better understand what
you have read. Look carefully at Figure 10.15 on page 309. What important idea does this illustration
communicate?

Name _____ Date _____ Class _____

GUIDED PRACTICE PROBLEMS

GUIDED PRACTICE PROBLEM 1 (page 289)

1. If 0.20 bushels is 1 dozen apples and a dozen apples has a mass of 2.0 kg, what is the mass of 0.50 bushel of apples?

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

_____

Unknown

_____

Use dimensional analysis to convert the number of bushels to the mass of apples, by following this sequence of conversions:

Number of bushels → dozens of apples → mass of apples

Calculate

Step 2. Solve for the unknown.
The first conversion factor is:

The second conversion factor is:

Multiplying the number of bushels by these two conversion factors gives the answer in kilograms.

The mass of 0.50 bushel of apples is _____.

Evaluate

Step 3. Does the result make sense?

_____

_____

_____

GUIDED PRACTICE PROBLEM 3 (page 291)

3. How many moles is 2.80×10^{24} atoms of silicon?

Step 1. List what you know. 2.80×10^{24} atoms of Si

[   ] atoms in one mole

Step 2. Multiply the atoms of silicon by a mol/atoms conversion factor.

2.80×10^{24} atoms Si $\times$ $\dfrac{1 \text{ mol}}{[\qquad] \text{ atoms Si}}$

Step 3. Divide.   [   ] mol

GUIDED PRACTICE PROBLEM 5 (page 292)

5. How many atoms are in 1.14 mol SO_3?

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

_____

Unknown

_____

Calculate

Step 2. Solve for the unknown.
The first conversion factor is $\dfrac{6.02 \times 10^{23} \text{ molecules of water}}{1 \text{ mol water}}$

The second conversion factor is $\dfrac{4 \text{ atoms}}{1 \text{ molecule } SO_3}$.

Multiply moles of SO_3 by these conversion factors:

number of atoms = 1.14 ~~mol SO_3~~ $\times$ $\dfrac{6.02 \times 10^{23} \text{ ~~molecules of } SO_3~~}}{\text{~~1 mol } SO_3~~}$ $\times$ $\dfrac{4 \text{ atoms}}{\text{~~1 molecule } SO_3~~}$.

= _____

Evaluate

Step 3. Does the result make sense?

_____

_____

CHAPTER 10, Chemical Quantities *(continued)*

EXTRA PRACTICE (similar to Practice Problem 5, page 292)

4. How many molecules is 0.360 mol of water?

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

Unknown

_____

Calculate

Step 2. Solve for the unknown.
The conversion factor is $\dfrac{6.02 \times 10^{23} \text{ molecules of water}}{1 \text{ mol water}}$

Multiplying mols of water by this conversion factor will give the answer

molecules of water = 0.360 ~~mol water~~ $\times \dfrac{6.02 \times 10^{23} \text{ molecules of water}}{1 ~~\text{mol water}~~}$

= _____

Evaluate

Step 3. Does the result make sense?

_____

_____

GUIDED PRACTICE PROBLEM 7 (page 296)

7. Find the molar mass of PCl_3.

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

_____

Name _____ Date _____ Class _____

Unknown

_____

Calculate

Step 2. Solve for the unknown.
Convert moles of phosphorus and chlorine to grams of phosphorus and chlorine. Then add to get the results.

$1 \; \cancel{mol \; P} \times \dfrac{31.0 \text{ g P}}{1 \; \cancel{mol \; P}} = 31.0 \text{ g P}$

$3 \; \cancel{mol \; Cl} \times \dfrac{35.5 \text{ g Cl}}{1 \; \cancel{mol \; Cl}} = 106.5 \text{ g Cl}$

molar mass of PCl_3 = _____

Evaluate

Step 3. Does the result make sense?

_____

_____

EXTRA PRACTICE (similar to Practice Problem 5, page 292)

5. How many atoms are there in 2.00 moles of SO_3?

EXTRA PRACTICE (similar to Practice Problem 7, page 296)

7. Find the molar mass of table salt (sodium chloride).

EXTRA PRACTICE (similar to Practice Problem 8, page 296)

8. What is the mass of 1 mole of ozone (O_3)?

Name _____ Date _____ Class _____

GUIDED PRACTICE PROBLEM 16 (page 298)

16. Find the mass, in grams, of 4.52×10^{-3} mol $C_{20}H_{42}$.

Analyze

Step 1. List the known and the unknown.
Known

_____

Unknown

_____

Calculate

Step 2. Solve for the unknown.
Determine the molar mass of $C_{20}H_{42}$:

1 mol $C_{20}H_{42}$ = 20 $\times$ 12.0 g + 42 $\times$ 1.0 g = 282 g

Multiply the given number of moles by the conversion factor:

mass = 4.52×10^{-3} ~~mol $C_{20}H_{42}$~~ $\times$ $\dfrac{282 \text{ g } C_{20}H_{42}}{1 \text{ mol } C_{20}H_{42b}}$ = _____

Evaluate

Step 3. Does the result make sense?

_____

_____

EXTRA PRACTICE (similar to Practice Problem 17, page 298)

17. Calculate the mass, in grams, of 10 mol of sodium sulfate (Na_2SO_4).

Calculate the mass, in grams, of 10 mol of iron(II) hydroxide ($Fe(OH)_2$).

GUIDED PRACTICE PROBLEM 18 (page 299)

18. Find the number of moles in 3.70×10^{-1} g of boron.

Analyze

Step 1. List the known and the unknown.
Known

_____

Unknown

_____

The unknown number of moles is calculated by converting the known mass to the number of moles using a conversion factor of mass → moles.

Calculate

Step 2. Solve for the unknown.
Determine the molar mass of boron: 1 mol B = 10.8 g B

Multiply the given mass by the conversion factor relating mass of boron to moles of boron:

$$\text{mass} = 3.70 \times 10^{-1} \, \text{g B} \times \frac{1 \, \text{mol B}}{10.8 \, \text{g B}}$$

= _____

Evaluate

Step 3. Does the result make sense?

_____

_____

GUIDED PRACTICE PROBLEM 20 (page 301)

20. What is the volume of these gases at STP?
   a. 3.20×10^{-3} mol CO_2
   b. 3.70 mol N_2

a. 3.20×10^{-3} mol CO_2
Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

CHAPTER 10, Chemical Quantities *(continued)*

Unknown

_____

To convert moles to liters, use the relationship 1 mol CO_2 = 22.4 L CO_2 (at STP).

Calculate

Step 2. Solve for the unknown.
Multiply the given number of moles of CO_2 by the conversion factor to give the result.

volume = 3.20×10^{-3} ~~mol CO₂~~ $\times \dfrac{22.4 \text{ L } CO_2}{1 \text{ ~~mol CO₂~~}}$

= _____

Evaluate

Step 3. Does the result make sense?

_____

_____

b. 3.70 mol N_2

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

Unknown

_____

Use the relationship 1 mol N_2 = 22.4 L N_2 (at STP) to convert moles to liters.

Calculate

Step 2. Solve for the unknown.
Multiply the given number of moles of N_2 by the conversion factor to give the result.

volume = 3.70 ~~mol N₂~~ $\times \dfrac{22.4 \text{ L } N_2}{1 \text{ ~~mol N₂~~}}$

= _____

Name _____ Date _____ Class _____

Evaluate

Step 3. Does the result make sense?

_____

_____

GUIDED PRACTICE PROBLEM 22 (page 302)

22. A gaseous compound composed of sulfur and oxygen, which is linked to the formation of acid rain, has a density of 3.58 g/L at STP. What is the molar mass of this gas?

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

Unknown

_____

To convert density (g/L) to molar mass (g/mol), a conversion factor of L/mol is needed.

Calculate

Step 2. Solve for the unknown.
Multiply the density by the conversion factor relating liters and moles.

$$\text{molar mass} = \frac{3.58 \text{ g}}{1 \text{ L}} \times \frac{22.4 \text{ L}}{1 \text{ mol}}$$

= _____

Evaluate

Step 3. Does the result make sense?

_____

_____

CHAPTER 10, Chemical Quantities *(continued)*

GUIDED PRACTICE PROBLEM 32 (page 306)

32. A compound is formed when 9.03 g Mg combines with 3.48 g N. What is the percent composition of this compound?

Analyze

Step 1. List the known and the unknowns.
Knowns

_____ _____ ___ _____

_____ _____

_____

Unknowns

_____ _____

_____

The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To be expressed as a percentage, the ratio must be multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent Mg} = \frac{9.03 \text{ g Mg}}{12.51 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

$$\text{percent N} = \frac{3.48 \text{ g N}}{12.51 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

Evaluate

Step 3. Does the result make sense?

_____

_____

GUIDED PRACTICE PROBLEM 34 (page 307)

34. Calculate the percent composition of these compounds.
a. ethane (C_2H_6)
b. sodium bisulfate ($NaHSO_4$)

a. ethane (C_2H_6)

Analyze

Step 1. List the knowns and the unknowns.
Knowns

_____

_____

_____

Unknowns

_____

_____

Because no masses are given, the percent composition can be determined based on the molar mass of the substance. The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To express the ratio as a percent, the ratio is multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent C} = \frac{24.0 \text{ g C}}{30.0 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

$$\text{percent H} = \frac{6.0 \text{ g H}}{30.0 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

Evaluate

Step 3. Does the result make sense?

_____

_____

CHAPTER 10, Chemical Quantities *(continued)*

b. sodium bisulfate (NaHSO$_4$)
Analyze

Step 1. List the knowns and the unknowns.
Knowns

_____

_____

_____

_____

_____

Unknowns

_____

_____

_____

_____

Because no masses are given, the percent composition can be determined based on the molar mass of the substance. The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To express the ratio as a percent, the ratio is multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent Na} = \frac{23.0 \text{ g Na}}{120.1 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

$$\text{percent H} = \frac{1.0 \text{ g H}}{120.1 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

$$\text{percent S} = \frac{32.1 \text{ g S}}{120.1 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

$$\text{percent O} = \frac{64.0 \text{ g O}}{120.1 \text{ g compound}} \times 100\% = \underline{\hspace{3cm}}$$

Evaluate

Step 3. Does the result make sense?

_____

_____

GUIDED PRACTICE PROBLEM 36 (page 310)

36. Calculate the empirical formula of each compound.
   a. 94.1% O, 5.9% H
   b. 67.6% Hg, 10.8% S, 21.6% O

a. 94.1% O, 5.9% H
Analyze

Step 1. List the knowns and the unknown.
Knowns

_____

_____

_____

Unknown

_____

Use the percent composition to convert to mass, recalling that percent means parts per hundred. Then use the molar mass to convert to number of moles. Finally, determine whole number ratios based on the number of moles of each element per 100 grams of compound.

Calculate

Step 2. Solve for the unknown.
One hundred grams of compound has 5.9 g H and 94.1 g O.
Multiply by conversion factors relating moles of the elements to grams.

$$5.9 \; \cancel{g \, H} \times \frac{1 \; mol \; H}{1.0 \; \cancel{g \, H}} = 5.9 \; mol \; H$$

$$94.1 \; \cancel{g \, O} \times \frac{1 \; mol \; O}{16.0 \; \cancel{g \, O}} = 5.88 \; mol \; O$$

So the mole ratio for 100 g of the compound is $H_{5.9}O_{5.9}$. But formulas must have whole number subscripts. Divide each molar quantity by the smaller number of moles. This will give 1 mol for the element with the smaller number of moles. In this case, the ratio is one-to-one and so the empirical formula is simply H_1O_1. However, a subscript of one is never written, so the answer is _____.

Evaluate

Step 3. Does the result make sense?

_____

_____

CHAPTER 10, Chemical Quantities *(continued)*

b. 67.6% Hg, 10.8% S, 21.6% O

Analyze

Step 1. List the knowns and the unknown.

Knowns

_____

_____

_____

_____

Unknown

_____

Use the percent composition to convert to mass. Then use molar mass to convert to number of moles. Finally, determine whole number ratios based on the number of moles of each element per 100 grams of compound.

Calculate

Step 2. Solve for the unknown.

One hundred grams of compound has 67.6 g Hg, 10.8 g S, and 21.6 g O. Multiply by a conversion factor relating moles to grams.

$$67.6 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} = 0.337 \text{ mol Hg}$$

$$10.8 \text{ g S} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} = 0.336 \text{ mol S}$$

$$21.6 \text{ g O} \times \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 1.35 \text{ mol O}$$

So the mole ratio for 100 g of the compound is $Hg_{0.34}S_{0.34}O_{1.35}$.
Divide each molar quantity by the smaller number of moles.

$$\frac{0.34 \text{ mol Hg}}{0.34} = 1 \text{ mol Hg}$$

$$\frac{0.34 \text{ mol S}}{0.34} = 1 \text{ mol S}$$

$$\frac{1.35 \text{ mol O}}{1.35} = 4 \text{ mol O}$$

The empirical formula is _____ .

Evaluate

Step 3. Does the result make sense?

_____

_____

Name _____ Date _____ Class _____

GUIDED PRACTICE PROBLEM 38 (page 312)

38. Find the molecular formula of ethylene glycol, which is used as antifreeze. The molar mass is 62 g/mol and the empirical formula is CH_3O.

Analyze

Step 1. List the knowns and the unknown.
Knowns

_____ _____

_____

Unknown

_____

Calculate

Step 2. Solve for the unknown.
First, calculate the empirical formula mass (efm):

$1 \text{ mol C} \times \dfrac{12 \text{ g C}}{1 \text{ mol C}} = 12 \text{ g C}$

$3 \text{ mol H} \times \dfrac{1.0 \text{ g H}}{1 \text{ mol H}} = 3 \text{ g H}$

$1 \text{ mol O} \times \dfrac{16 \text{ g O}}{1 \text{ mol O}} = 16 \text{ g O}$

So efm = 12g + 3 g + 16 g = 31 g.

Divide the molar mass by the empirical formula mass:
Molar mass/efm = 62 g/31 g = 2

Multiply subscripts in the empirical formula by this value.

The molecular formula is _____ .

Evaluate

Step 3. Does the result make sense?

_____

_____