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# The Mole

## Section 10.1 Measuring Matter

page 320–324

### Practice Problems

pages 323–324

1. Zinc (Zn) is used to form a corrosion-inhibiting surface on galvanized steel. Determine the number of Zn atoms in 2.50 mol of Zn.

$$2.50 \text{ mol Zn} \times \frac{6.02 \times 10^{23} \text{ atoms Zn}}{1 \text{ mol Zn}} \\ = 1.51 \times 10^{24} \text{ atoms of Zn}$$

2. Calculate the number of molecules in 11.5 mol of water (H<sub>2</sub>O).

$$11.5 \text{ mol H}_2\text{O} \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \\ = 6.92 \times 10^{24} \text{ molecules of H}_2\text{O}$$

3. Silver nitrate (AgNO<sub>3</sub>) is used to make several different silver halides used in photographic films. How many formula units of AgNO<sub>3</sub> are there in 3.25 mol AgNO<sub>3</sub>?

$$3.25 \text{ mol AgNO}_3 \times \frac{6.02 \times 10^{23} \text{ formula units AgNO}_3}{1 \text{ mol AgNO}_3} \\ = 1.96 \times 10^{24} \text{ formula units of AgNO}_3$$

4. **Challenge** Calculate the number of oxygen atoms in 5.00 mol of oxygen molecules. Oxygen is a diatomic molecule, O<sub>2</sub>.

$$5.00 \text{ mol O}_2 \times \frac{6.02 \times 10^{23} \text{ molecules O}_2}{1 \text{ mol O}_2} \\ \times \frac{2 \text{ O atoms}}{1 \text{ molecule O}_2} = 6.02 \times 10^{24} \text{ atoms O}$$

5. How many moles contain each of the following?

a.  $5.75 \times 10^{24}$  atoms Al

$$5.75 \times 10^{24} \text{ atoms Al} \times \frac{1 \text{ mol Al}}{6.02 \times 10^{23} \text{ atoms Al}} \\ = 9.55 \text{ mol Al}$$

b.  $2.50 \times 10^{20}$  atoms Fe

$$2.50 \times 10^{20} \text{ atoms Fe} \times \frac{1 \text{ mol Fe}}{6.02 \times 10^{23} \text{ atoms Fe}} \\ = 4.15 \times 10^{-4} \text{ mol Fe}$$

6. **Challenge** Identify the representative particle for each formula and convert the given number of representative particles to moles.

a.  $3.75 \times 10^{24}$  CO<sub>2</sub>

The representative particle is a molecule.

$$3.75 \times 10^{24} \text{ molecules CO}_2 \\ \times \frac{1 \text{ mol CO}_2}{6.02 \times 10^{23} \text{ molecules CO}_2} = 6.23 \text{ mol CO}_2$$

b.  $3.58 \times 10^{23}$  ZnCl<sub>2</sub>

The representative particle is a formula unit.

$$3.58 \times 10^{23} \text{ formula units ZnCl}_2 \\ \times \frac{1 \text{ mol ZnCl}_2}{6.02 \times 10^{23} \text{ formula units ZnCl}_2} \\ = 0.595 \text{ mol ZnCl}_2$$

## Section 10.1 Assessment

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7. **Explain** why chemists use the mole.

Chemists use the mole because it is a convenient way of knowing how many representative particles are in a sample.

8. **State** the mathematical relationship between Avogadro's number and 1 mol.

One mole contains  $6.02 \times 10^{23}$  representative particles.

9. **List** the conversion factors used to convert between particles and moles.

$$\frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mol}}, \\ \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ representative particles}}$$

**10. Explain** how a mole is similar to a dozen.

The mole is a unit for counting  $6.02 \times 10^{23}$  representative particles. The dozen is used to count 12 items.

**11. Apply** How does a chemist count the number of particles in a given number of moles of a substance?

They multiply the number moles by Avogadro's number.

**12. Calculate** the mass in atomic mass units of 0.25 mol of carbon-12 atoms.

$$0.25 \text{ mol carbon-12} \times \frac{6.02 \times 10^{23} \text{ carbon-12 atoms}}{1 \text{ mol carbon-12}} \\ \times \frac{12 \text{ amu}}{1 \text{ carbon-12 atom}} = 1.8 \times 10^{24} \text{ amu}$$

**13. Calculate** the number of representative particles of each substance.

a. 11.5 mol Ag

$$11.5 \text{ mol Ag} \times \frac{6.02 \times 10^{23} \text{ atoms Ag}}{1 \text{ mol Ag}} \\ = 6.92 \times 10^{24} \text{ atoms Ag}$$

b. 18.0 mol H<sub>2</sub>O

$$18.0 \text{ mol H}_2\text{O} \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \\ = 1.08 \times 10^{25} \text{ molecules H}_2\text{O}$$

c. 0.150 mol NaCl

$$0.150 \text{ mol NaCl} \times \frac{6.02 \times 10^{23} \text{ formula units NaCl}}{1 \text{ mol NaCl}} \\ = 9.03 \times 10^{22} \text{ formula units NaCl}$$

d.  $1.35 \times 10^{-2}$  mol CCH<sub>4</sub>

$$1.35 \times 10^{-2} \text{ mol CCH}_4 \\ \times \frac{6.02 \times 10^{23} \text{ molecules CCH}_4}{1 \text{ mol CCH}_4} \\ = 8.13 \times 10^{21} \text{ molecules CCH}_4$$

**14. Arrange** these three samples from smallest to largest in terms of number of representative particles:  $1.25 \times 10^{25}$  atoms of zinc (Zn), 3.56 mol of iron (Fe), and  $6.78 \times 10^{22}$  molecules of glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>).

$$3.56 \text{ mol Fe} \times \frac{6.02 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} \\ = 2.14 \times 10^{24} \text{ atoms Fe}$$

From smallest to largest:  $6.78 \times 10^{22}$  molecules glucose,  $2.14 \times 10^{24}$  atoms Fe,  $1.25 \times 10^{25}$  atoms Zn.

## Section 10.2 Mass and the Mole

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### Problem-Solving Lab

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Hydrogen - 1



Helium - 4

**1. Apply** What is the mass in grams of one helium atom? (The mass of a neutron is approximately the same as the mass of a proton.)

Because neutrons and protons have about the same mass, treat the two neutrons and two protons in the He atom as four protons.

$$\frac{1.672 \times 10^{-24} \text{ g}}{1 \text{ proton}} \times \frac{4 \text{ protons}}{\text{He atom}} \\ = 6.69 \times 10^{-24} \text{ g/He atom}$$

**2. Draw** Carbon-12 contains six protons and six neutrons. Draw the carbon-12 nucleus and calculate the mass of one atom in amu and g.

A correctly drawn model of the nucleus will have six protons and six neutrons packed in a tight spherical shape. Because neutrons and protons have about the same mass, treat the six neutrons and six protons in the C-12 atom as 12 protons.

$$\frac{1.672 \times 10^{-24} \text{ g}}{1 \text{ proton}} \times \frac{12 \text{ protons}}{\text{C-12 atom}} \\ = 2.01 \times 10^{-23} \text{ g/C-12 atom}$$

- 3. Apply** How many atoms of hydrogen-1 are in a 1.007-g sample? Recall that 1.007 amu is the mass of one atom of hydrogen-1. Round your answer to two significant digits.

$$1.007 \text{ g} \times \frac{1 \text{ H-1 atom}}{1.672 \times 10^{-24} \text{ g}} \\ = 6.023 \times 10^{23} \text{ H-1 atoms}$$

- 4. Apply** If you had samples of helium and carbon that contained the same number of atoms as you calculated in Question 1, what would be the mass in grams of each sample?

$$\text{He: } 6.023 \times 10^{23} \text{ He atoms} \times \frac{6.69 \times 10^{-24} \text{ g}}{\text{He atom}} \\ = 4.03 \text{ g}$$

$$\text{C: } 6.023 \times 10^{23} \text{ C atoms} \times \frac{2.01 \times 10^{-23} \text{ g}}{\text{C atom}} \\ = 12.1 \text{ g}$$

- 5. Conclude** What can you conclude about the relationship between the number of atoms and the mass of each sample?

The mass of one mole of any atom is the same value in grams as the mass of one atom in amu.

## Practice Problems

pages 328–331

- 15.** Determine the mass in grams of each of the following.

- a. 3.57 mol Al

$$3.57 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 96.3 \text{ g Al}$$

- b. 42.6 mol Si

$$42.6 \text{ mol Si} \times \frac{28.09 \text{ g Si}}{1 \text{ mol Si}} = 1.20 \times 10^3 \text{ g Si}$$

- 16. Challenge** Convert each given quantity in scientific notation to mass in grams expressed in scientific notation.

- a.  $3.45 \times 10^2$  mol Co

$$3.45 \times 10^2 \text{ mol Co} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} \\ = 2.03 \times 10^4 \text{ g Co}$$

- b.  $2.45 \times 10^{-2}$  mol Zn

$$2.45 \times 10^{-2} \text{ mol Zn} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 1.60 \text{ g Zn}$$

- 17.** Determine the number of moles in each of the following.

- a. 25.5 g Ag

$$25.5 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}} = 0.236 \text{ mol Ag}$$

- b. 300.0 g S

$$300.0 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 9.355 \text{ mol S}$$

- 18. Challenge** Convert each mass to moles. Express the answer in scientific notation.

- a.  $1.25 \times 10^3$  g Zn

$$1.25 \times 10^3 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} = 19.1 \text{ mol Zn} \\ = 1.91 \times 10^1 \text{ mol Zn}$$

- b. 1.00 kg Fe

$$1.00 \text{ kg Fe} \times \frac{1000 \text{ g Fe}}{1 \text{ kg Fe}} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \\ = 17.9 \text{ mol Fe} = 1.79 \times 10^1 \text{ mol Fe}$$

- 19.** How many atoms are in each of the following samples?

- a. 55.2 g Li

$$55.2 \text{ g Li} \times \frac{1 \text{ mol Li}}{6.94 \text{ g Li}} \times \frac{6.02 \times 10^{23} \text{ atoms Li}}{1 \text{ mol Li}} \\ = 4.79 \times 10^{24} \text{ atoms Li}$$

- b. 0.230 g Pb

$$0.230 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \times \frac{6.02 \times 10^{23} \text{ atoms Pb}}{1 \text{ mol Pb}} \\ = 6.68 \times 10^{20} \text{ atoms Pb}$$

- c. 11.5 g Hg

$$11.5 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} \times \frac{6.02 \times 10^{23} \text{ atoms Hg}}{1 \text{ mol Hg}} \\ = 3.45 \times 10^{22} \text{ atoms Hg}$$

20. What is the mass in grams of each of the following?

a.  $6.02 \times 10^{24}$  atoms Bi

$$6.02 \times 10^{24} \text{ atoms Bi} \times \frac{1 \text{ mol Bi}}{6.02 \times 10^{23} \text{ atoms Bi}} \times \frac{209.0 \text{ g Bi}}{1 \text{ mol Bi}} = 2.09 \times 10^3 \text{ g Bi}$$

b.  $1.00 \times 10^{24}$  atoms Mn

$$1.00 \times 10^{24} \text{ atoms Mn} \times \frac{1 \text{ mol Mn}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{54.94 \text{ g Mn}}{1 \text{ mol Mn}} = 91.3 \text{ g Mn}$$

c.  $3.40 \times 10^{22}$  atoms He

$$3.40 \times 10^{22} \text{ atoms He} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 0.226 \text{ g He}$$

d.  $1.50 \times 10^{15}$  atoms N

$$1.50 \times 10^{15} \text{ atoms N} \times \frac{1 \text{ mol N}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 3.49 \times 10^{-8} \text{ g N}$$

e.  $1.50 \times 10^{15}$  atoms U

$$1.50 \times 10^{15} \text{ atoms U} \times \frac{1 \text{ mol U}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{238.0 \text{ g U}}{1 \text{ mol U}} = 5.93 \times 10^{-7} \text{ g U}$$

21. **Challenge** Convert each given mass to number of representative particles. Identify the type of representative particle, and express the number in scientific notation.

a.  $4.56 \times 10^3$  g Si

$$4.56 \times 10^3 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 9.77 \times 10^{25} \text{ atoms Si}$$

b. 0.120 kg Ti

$$0.120 \text{ kg Ti} \times \frac{1000 \text{ g Ti}}{1 \text{ kg Ti}} \times \frac{1 \text{ mol Ti}}{47.87 \text{ g Ti}} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 1.51 \times 10^{24} \text{ atoms Ti}$$

## Section 10.2 Assessment

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22. **Summarize** in terms of particles and mass, one-mole quantities of two different monatomic elements.

Each one-mole quantity has  $6.02 \times 10^{23}$  particles (atoms), but they will have different masses.

23. **State** the conversion factors needed to convert between mass and moles of the element fluorine.

Mass-to-mole conversions use the conversion factor  $1 \text{ mol}/18.998 \text{ g}$ . Moles-to-mass conversions use the conversion factor  $18.998 \text{ g}/1 \text{ mol}$ .

24. **Explain** how molar mass relates the mass of an atom to the mass of a mole of atoms.

Molar mass is the mass in grams of one mole of any pure substance.

25. **Describe** the steps used to convert the mass of an element to the number of atoms of the element.

Multiply the mass by the inverse of molar mass, and then multiply by Avogadro's number.

26. **Arrange** these quantities from smallest to largest in terms of mass: 1.0 mol of Ar,  $3.0 \times 10^{24}$  atoms of Ne, and 20 g of Kr.

$$1.0 \text{ mol Ar} \times \frac{39.95 \text{ g Ar}}{1 \text{ mol Ar}} = 39.95 \text{ g Ar}$$

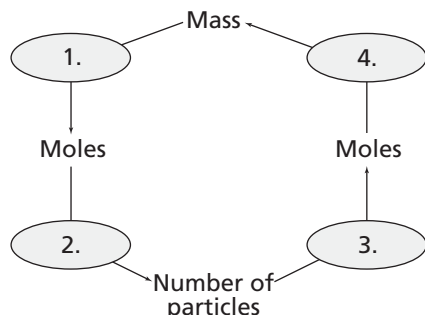
$$3.0 \times 10^{24} \text{ atoms Ne} \times \frac{1 \text{ mol Ne}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} = 101 \text{ g Ne}$$

20 g Kr, 1.0 mol Ar,  $3.0 \times 10^{24}$  atoms Ne

27. **Identify** the quantity calculated by dividing the molar mass of an element by Avogadro's number.

Because the molar mass is a ratio of grams per mole and Avogadro's number is a ratio of particles per mole, dividing the molar mass of an element by Avogadro's number yields the mass of a single representative particle of that element.

- 28. Design** a concept map that shows the conversion factors needed to convert among mass, moles, and number of particles.



Student concept maps will vary, but should show the correct sets of conversion factors needed to convert among mass, moles, and number of particles.

## Section 10.3 Moles of Compounds

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### Practice Problems

pages 335–339

- 29.** Zinc chloride ( $\text{ZnCl}_2$ ) is used in soldering flux, an alloy used to join two metals together. Determine the moles of  $\text{Cl}^-$  ions in 2.50 mol  $\text{ZnCl}_2$ .

$$2.50 \text{ mol ZnCl}_2 \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol ZnCl}_2} = 5.00 \text{ mol Cl}^-$$

- 30.** Plants and animals depend on glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) as an energy source. Calculate the number of moles of each element in 1.25 mol  $\text{C}_6\text{H}_{12}\text{O}_6$ .

$$1.25 \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{6 \text{ mol C}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = 7.50 \text{ mol C}$$

$$1.25 \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{12 \text{ mol H}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = 15.0 \text{ mol H}$$

$$1.25 \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{6 \text{ mol O}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = 7.50 \text{ mol O}$$

- 31.** Iron(III) sulfate [ $\text{Fe}_2(\text{SO}_4)_3$ ] is sometimes used in the water purification process. Determine the number of moles of sulfate ions present in 3.00 mol of  $\text{Fe}_2(\text{SO}_4)_3$ .

$$3.00 \text{ mol Fe}_2(\text{SO}_4)_3 \times \frac{3 \text{ mol SO}_4^{2-}}{1 \text{ mol Fe}_2(\text{SO}_4)_3} = 9.00 \text{ mol SO}_4^{2-}$$

- 32.** How many moles of oxygen atoms are present in 5.00 mol diphosphorus pentoxide ( $\text{P}_2\text{O}_5$ )?

$$5.00 \text{ mol P}_2\text{O}_5 \times \frac{5 \text{ mol O}}{1 \text{ mol P}_2\text{O}_5} = 25.0 \text{ mol O}$$

- 33. Challenge** Calculate the number of moles of hydrogen atoms in  $1.15 \times 10^1$  mol of water. Express the answer in scientific notation.

$$1.15 \times 10^1 \text{ mol H}_2\text{O} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 23.0 \text{ mol H} = 2.30 \times 10^1 \text{ mol H}$$

- 34.** Determine the molar mass of each ionic compound.

- a. NaOH

$$1 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 22.99 \text{ g Na}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$1 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.008 \text{ g H}$$

$$\text{molar mass NaOH} = 40.00 \text{ g/mol}$$

- b.  $\text{CaCl}_2$

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g Ca}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass CaCl}_2 = 110.98 \text{ g/mol}$$

- c.  $\text{KC}_2\text{H}_3\text{O}_2$

$$1 \text{ mol K} \times \frac{39.10 \text{ g K}}{1 \text{ mol K}} = 39.10 \text{ g K}$$

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g C}$$

$$3 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 3.024 \text{ g H}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass KC}_2\text{H}_3\text{O}_2 = 98.14 \text{ g/mol}$$

35. Calculate the molar mass of each molecular compound.

a.  $C_2H_5OH$

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g C}$$

$$6 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 6.048 \text{ g H}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{16.00 \text{ g O}}$$

$$\text{molar mass } C_2H_5OH = 46.07 \text{ g/mol}$$

b. HCN

$$1 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.008 \text{ g H}$$

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$1 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = \underline{14.01 \text{ g N}}$$

$$\text{molar mass HCN} = 27.03 \text{ g/mol}$$

c.  $CCl_4$

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$4 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = \underline{141.80 \text{ g Cl}}$$

$$\text{molar mass } CCl_4 = 153.81 \text{ g/mol}$$

36. **Challenge** Identify each substance as a molecular compound or an ionic compound, and then calculate its molar mass.

a.  $Sr(NO_3)_2$

ionic compound;

$$1 \text{ mol Sr} \times \frac{87.62 \text{ g Sr}}{1 \text{ mol Sr}} = 87.62 \text{ g Sr}$$

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 28.02 \text{ g N}$$

$$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{96.00 \text{ g O}}$$

$$\text{molar mass } Sr(NO_3)_2 = 211.64 \text{ g/mol}$$

b.  $(NH_4)_3PO_4$

ionic compound;

$$3 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 42.03 \text{ g N}$$

$$12 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 12.096 \text{ g H}$$

$$1 \text{ mol P} \times \frac{30.97 \text{ g P}}{1 \text{ mol P}} = 30.97 \text{ g P}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{64.00 \text{ g O}}$$

$$\text{molar mass } (NH_4)_3PO_4 = 149.10 \text{ g/mol}$$

c.  $C_{12}H_{22}O_{11}$

molecular compound;

$$12 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 144.12 \text{ g C}$$

$$22 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 22.176 \text{ g H}$$

$$11 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{176.00 \text{ g O}}$$

$$\text{molar mass } C_{12}H_{22}O_{11} = 342.30 \text{ g/mol}$$

37. The United States chemical industry produces more sulfuric acid ( $H_2SO_4$ ) in terms of mass, than any other chemical. What is the mass of 3.25 mol of  $H_2SO_4$ ?

Step 1: Find the molar mass of  $H_2SO_4$ .

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{64.00 \text{ g O}}$$

$$\text{molar mass } H_2SO_4 = 98.09 \text{ g/mol}$$

Step 2: Make mole  $\rightarrow$  mass conversion.

$$3.25 \text{ mol } H_2SO_4 \times \frac{98.09 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4} = 319 \text{ g } H_2SO_4$$

38. What is the mass of  $4.35 \times 10^{-2}$  moles of zinc chloride ( $\text{ZnCl}_2$ )?

Step 1: Find the molar mass of  $\text{ZnCl}_2$ .

$$1 \text{ mol Zn} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 65.38 \text{ g Zn}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass ZnCl}_2 = 136.28 \text{ g/mol}$$

Step 2: Make mole  $\rightarrow$  mass conversion.

$$4.35 \times 10^{-2} \text{ mol ZnCl}_2 \times \frac{136.28 \text{ g ZnCl}_2}{1 \text{ mol ZnCl}_2} = 5.93 \text{ g ZnCl}_2$$

39. **Challenge** Write the chemical formula for potassium permanganate, and then calculate the mass in grams of 2.55 mol of the compound.

Potassium permanganate has a formula of  $\text{KMnO}_4$ .

Step 1: Find the molar mass of  $\text{KMnO}_4$ .

$$1 \text{ mol K} \times \frac{39.10 \text{ g K}}{1 \text{ mol K}} = 39.10 \text{ g K}$$

$$1 \text{ mol Mn} \times \frac{54.94 \text{ g Mn}}{1 \text{ mol Mn}} = 54.94 \text{ g Mn}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass KMnO}_4 = 158.04 \text{ g/mol}$$

Step 2: Make mole  $\rightarrow$  mass conversion.

$$2.55 \text{ mol KMnO}_4 \times \frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} = 403 \text{ g KMnO}_4$$

40. Determine the number of moles present in each compound.

- a. 22.6 g  $\text{AgNO}_3$

Step 1: Find the molar mass of  $\text{AgNO}_3$ .

$$1 \text{ mol Ag} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 107.9 \text{ g Ag}$$

$$1 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 14.01 \text{ g N}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass AgNO}_3 = 169.9 \text{ g/mol}$$

Step 2: Make mass  $\rightarrow$  mole conversion.

$$22.6 \text{ g AgNO}_3 \times \frac{1 \text{ mol AgNO}_3}{169.9 \text{ g AgNO}_3} = 0.133 \text{ mol AgNO}_3$$

- b. 6.50 g  $\text{ZnSO}_4$

Step 1: Find the molar mass of  $\text{ZnSO}_4$ .

$$1 \text{ mol Zn} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 65.38 \text{ g Zn}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass ZnSO}_4 = 161.46 \text{ g/mol}$$

Step 2: Make mass  $\rightarrow$  mole conversion.

$$6.50 \text{ g ZnSO}_4 \times \frac{1 \text{ mol ZnSO}_4}{161.46 \text{ g ZnSO}_4} = 0.0403 \text{ mol ZnSO}_4$$

- c. 35.0 g  $\text{HCl}$

Step 1: Find the molar mass of  $\text{HCl}$ .

$$1 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.008 \text{ g H}$$

$$1 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 35.45 \text{ g Cl}$$

$$\text{molar mass HCl} = 36.46 \text{ g/mol}$$

Step 2: Make mass  $\rightarrow$  mole conversion.

$$35.0 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 0.960 \text{ mol HCl}$$



**41. Challenge** Identify each as an ionic or molecular compound and convert the given mass to moles. Express your answers in scientific notation.

a. 2.50 kg Fe<sub>2</sub>O<sub>3</sub>

ionic compound; Step 1: Find the molar mass of Fe<sub>2</sub>O<sub>3</sub>.

$$2 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 111.70 \text{ g Fe}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

molar mass Fe<sub>2</sub>O<sub>3</sub> = 159.70 g/mol

Step 2: Make mass → mole conversion.

$$2.50 \text{ kg Fe}_2\text{O}_3 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.70 \text{ g Fe}_2\text{O}_3}$$

$$= 15.7 \text{ mol Fe}_2\text{O}_3 = 1.57 \times 10^1 \text{ mol Fe}_2\text{O}_3$$

b. 25.4 mg PbCl<sub>4</sub>

ionic compound; Step 1: Find the molar mass of PbCl<sub>4</sub>.

$$1 \text{ mol Pb} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 207.2 \text{ g Pb}$$

$$4 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 141.80 \text{ g Cl}$$

molar mass PbCl<sub>4</sub> = 349.0 g/mol

Step 2: Make mass → mole conversion.

$$25.4 \text{ mg PbCl}_4 \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol PbCl}_4}{349.0 \text{ g PbCl}_4}$$

$$= 7.28 \times 10^{-5} \text{ mol PbCl}_4$$

**42.** Ethanol (C<sub>2</sub>H<sub>5</sub>OH), a domestically produced fuel source, is often blended with gasoline. A sample of ethanol has a mass of 45.6 g.

a. How many carbon atoms does the sample contain?

Step 1: Find the molar mass of C<sub>2</sub>H<sub>5</sub>OH.

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g}$$

$$6 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 6.048 \text{ g}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g}$$

molar mass C<sub>2</sub>H<sub>5</sub>OH = 46.07 g/mol

Step 2: Make mass → mole conversion.

$$45.6 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.07 \text{ g C}_2\text{H}_5\text{OH}}$$

$$= 0.990 \text{ mol C}_2\text{H}_5\text{OH}$$

Step 3: Make mole → molecule conversion.

$$0.990 \text{ mol C}_2\text{H}_5\text{OH} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}}$$

$$= 5.96 \times 10^{23} \text{ molecules C}_2\text{H}_5\text{OH}$$

Step 4: Determine the number of carbon atoms.

$$5.96 \times 10^{23} \text{ molecules C}_2\text{H}_5\text{OH}$$

$$\times \frac{2 \text{ C atoms}}{1 \text{ molecule C}_2\text{H}_5\text{OH}} = 1.19 \times 10^{24} \text{ C atoms}$$

b. How many hydrogen atoms are present?

$$5.96 \times 10^{23} \text{ molecules C}_2\text{H}_5\text{OH}$$

$$\times \frac{6 \text{ H atoms}}{1 \text{ molecule C}_2\text{H}_5\text{OH}} = 3.58 \times 10^{24} \text{ H atoms}$$

c. How many oxygen atoms are present?

$$5.96 \times 10^{23} \text{ molecules C}_2\text{H}_5\text{OH}$$

$$\times \frac{1 \text{ O atoms}}{1 \text{ molecule C}_2\text{H}_5\text{OH}} = 5.96 \times 10^{23} \text{ O atoms}$$

**43.** A sample of sodium sulfite (Na<sub>2</sub>SO<sub>3</sub>) has a mass of 2.25 g.

a. How many Na<sup>+</sup> ions are present?

Step 1: Find the molar mass of Na<sub>2</sub>SO<sub>3</sub>

$$2 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 45.98 \text{ g}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g}$$

molar mass Na<sub>2</sub>SO<sub>3</sub> = 126.05 g/mol

Step 2: Make mass → mole conversion

$$2.25 \text{ g Na}_2\text{SO}_3 \times \frac{1 \text{ mol Na}_2\text{SO}_3}{126.05 \text{ g Na}_2\text{SO}_3}$$

$$= 0.0179 \text{ mol Na}_2\text{SO}_3$$

Step 3: Make mole → formula unit conversion

$$0.0179 \text{ mol Na}_2\text{SO}_3 \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \text{ mol}}$$

$$= 1.08 \times 10^{22} \text{ formula units Na}_2\text{SO}_3$$

Step 4: Determine the number of Na<sup>+</sup> ions.

$$1.08 \times 10^{22} \text{ formula units Na}_2\text{SO}_3 \times \frac{2 \text{ Na}^+ \text{ ions}}{1 \text{ formula unit Na}_2\text{SO}_3} = 2.16 \times 10^{22} \text{ Na}^+ \text{ ions}$$

b. How many SO<sub>3</sub><sup>2-</sup> ions are present?

$$1.08 \times 10^{22} \text{ formula units Na}_2\text{SO}_3 \times \frac{2 \text{ SO}_3^{2-} \text{ ions}}{1 \text{ formula unit Na}_2\text{SO}_3} = 1.08 \times 10^{22} \text{ SO}_3^{2-} \text{ ions}$$

c. What is the mass in grams of one formula unit of Na<sub>2</sub>SO<sub>3</sub>?

$$\frac{126.08 \text{ g Na}_2\text{SO}_3}{1 \text{ mol Na}_2\text{SO}_3} \times \frac{1 \text{ mol Na}_2\text{SO}_3}{6.02 \times 10^{23} \text{ formula units}} = 2.09 \times 10^{-22} \text{ g Na}_2\text{SO}_3/\text{formula unit}$$

44. A sample of carbon dioxide (CO<sub>2</sub>) has a mass of 52.0 g.

a. How many carbon atoms are present?

Step 1: Find the molar mass of CO<sub>2</sub>.

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g}$$

$$\text{molar mass CO}_2 = 44.01 \text{ g/mol}$$

Step 2: Make mass → mole conversion.

$$52.0 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} = 1.18 \text{ mol CO}_2$$

Step 3: Make mole → molecule conversion.

$$1.18 \text{ mol CO}_2 \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}}$$

$$= 7.11 \times 10^{23} \text{ molecules CO}_2$$

Step 4: Determine the number of carbon atoms.

$$7.11 \times 10^{23} \text{ molecules CO}_2 \times \frac{1 \text{ C atom}}{1 \text{ molecule CO}_2} = 7.11 \times 10^{23} \text{ C atoms}$$

b. How many oxygen atoms are present?

$$7.11 \times 10^{23} \text{ molecules CO}_2 \times \frac{2 \text{ O atoms}}{1 \text{ molecule CO}_2} = 1.42 \times 10^{24} \text{ O atoms}$$

c. What is the mass in grams of one molecule of CO<sub>2</sub>?

$$\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molecules}} = 7.31 \times 10^{-23} \text{ g CO}_2/\text{molecule}$$

45. What mass of sodium chloride (NaCl) contains 4.59 × 10<sup>24</sup> formula units?

Step 1: Find the number of moles of NaCl.

$$4.59 \times 10^{24} \text{ formula units NaCl} \times \frac{1 \text{ mol NaCl}}{6.02 \times 10^{23} \text{ formula units}} = 7.62 \text{ mol NaCl}$$

Step 2: Find the molar mass of NaCl.

$$1 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 22.99 \text{ g}$$

$$1 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 35.45 \text{ g}$$

$$\text{molar mass NaCl} = 58.44 \text{ g/mol}$$

Step 3: Make mole → mass conversion.

$$7.62 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 445 \text{ g NaCl}$$

46. **Challenge** A sample of silver chromate has a mass of 25.8 g.

a. Write the formula for silver chromate.

The formula for silver chromate is Ag<sub>2</sub>CrO<sub>4</sub>.

b. How many cations are present in the sample?

Step 1: Find the molar mass of Ag<sub>2</sub>CrO<sub>4</sub>.

$$2 \text{ mol Ag} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 215.8 \text{ g}$$

$$1 \text{ mol Cr} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}} = 52.00 \text{ g}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g}$$

$$\text{molar mass Ag}_2\text{CrO}_4 = 331.8 \text{ g/mol}$$

Step 2: Make mass → mole conversion.

$$25.8 \text{ g Ag}_2\text{CrO}_4 \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{331.8 \text{ g Ag}_2\text{CrO}_4}$$

$$= 0.0778 \text{ mol Ag}_2\text{CrO}_4$$

Step 3: Make mole → formula unit conversion.

$$0.0778 \text{ mol Ag}_2\text{CrO}_4 \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \text{ mol}}$$

$$= 4.68 \times 10^{22} \text{ formula units Ag}_2\text{CrO}_4$$

Step 4: Determine the number of cations (Ag<sup>+</sup> ions):

$$4.68 \times 10^{22} \text{ formula units Ag}_2\text{CrO}_4 \times \frac{2 \text{ Ag}^+ \text{ ions}}{1 \text{ formula unit Ag}_2\text{CrO}_4} = 9.36 \times 10^{22} \text{ Ag}^+ \text{ ions}$$

- c. How many anions are present in the sample?

The anions are CrO<sub>4</sub><sup>2-</sup> ions:

$$4.68 \times 10^{22} \text{ formula units Ag}_2\text{CrO}_4 \times \frac{1 \text{ CrO}_4^{2-} \text{ ions}}{1 \text{ formula unit Ag}_2\text{CrO}_4}$$

$$= 4.68 \times 10^{22} \text{ CrO}_4^{2-} \text{ ions}$$

- d. What is the mass in grams of one formula unit of silver chromate?

$$\frac{331.8 \text{ g Ag}_2\text{CrO}_4}{1 \text{ mol Ag}_2\text{CrO}_4} \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{6.02 \times 10^{23} \text{ formula units}}$$

$$= 5.51 \times 10^{-22} \text{ g Ag}_2\text{CrO}_4/\text{formula unit}$$

## Section 10.3 Assessment

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47. **Describe** how to determine the molar mass of a compound.

Multiply the mass of one mole of each element by the ratio of that element to one mole of the compound. Add the resulting masses.

48. **Identify** the conversion factors needed to convert between the number of moles and the mass of a compound.

$$\frac{\text{Number of grams}}{1 \text{ mol}}, \frac{1 \text{ mol}}{\text{number of grams}}$$

49. **Explain** how you can determine the number of atoms or ions in a given mass of a compound.

Convert the mass to moles, multiply the number of moles by the ratio of the number of atoms or ions to one mole, multiply by Avogadro's number.

50. **Apply** How many moles of K, C, and O atoms are there in 1 mol of K<sub>2</sub>C<sub>2</sub>O<sub>4</sub>?

2 mol K, 2 mol C, 4 mol O

51. **Calculate** the molar mass of MgBr<sub>2</sub>.

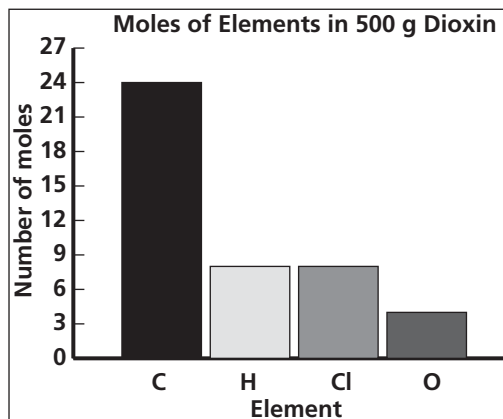
$$\text{molar mass MgBr}_2 = 24.305 \text{ g/mol} \times 2(79.904 \text{ g/mol}) = 184.113 \text{ g/mol}$$

52. **Calculate** Calcium carbonate is the calcium source for many vitamin tablets. The recommended daily allowance of calcium is 1000 mg of Ca<sup>2+</sup> ions. How many moles of Ca<sup>2+</sup> does 1000 mg represent?

$$1000 \text{ mg Ca}^{2+} \times \frac{1 \text{ g Ca}^{2+}}{10^3 \text{ mg Ca}^{2+}} \times \frac{1 \text{ mol Ca}^{2+}}{40.08 \text{ g Ca}^{2+}}$$

$$= 0.02 \text{ mol Ca}^{2+}$$

53. **Design** a bar graph that will show the number of moles of each element present in 500 g of a particular form of dioxin (C<sub>12</sub>H<sub>4</sub>Cl<sub>4</sub>O<sub>2</sub>), a powerful poison.



$$12 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 144.12 \text{ g C}$$

$$4 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 4.032 \text{ g H}$$

$$4 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 141.80 \text{ g Cl}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{32.00 \text{ g O}}$$

$$\text{molar mass} = 321.96 \text{ g/mol}$$

$$500 \text{ g C}_{12}\text{H}_4\text{Cl}_4\text{O}_2 \times \frac{1 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2}{321.96 \text{ g C}_{12}\text{H}_4\text{Cl}_4\text{O}_2}$$

$$= 2 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2$$

Student bar graphs should show the following molar quantities:

$$2 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2 \times \frac{12 \text{ mol C}}{1 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2} = 24 \text{ mol C}$$

$$2 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2 \times \frac{4 \text{ mol H}}{1 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2} = 8 \text{ mol H}$$

$$2 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2 \times \frac{4 \text{ mol Cl}}{1 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2} = 8 \text{ mol Cl}$$

$$2 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2 \times \frac{2 \text{ mol O}}{1 \text{ mol C}_{12}\text{H}_4\text{Cl}_4\text{O}_2} = 4 \text{ mol O}$$

## Section 10.4 Empirical and Molecular Formulas

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### Practice Problems

pages 344–350

54. What is the percent composition of phosphoric acid ( $\text{H}_3\text{PO}_4$ )?

Steps 1 and 2: Assume 1 mole; calculate molar mass of  $\text{H}_3\text{PO}_4$ .

$$3 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 3.024 \text{ g H}$$

$$1 \text{ mol P} \times \frac{30.97 \text{ g P}}{1 \text{ mol P}} = 30.97 \text{ g P}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{64.00 \text{ g O}}$$

$$\text{molar mass H}_3\text{PO}_4 = 97.99 \text{ g/mol}$$

Step 3: Determine percent by mass of each element.

$$\text{percent H} = \frac{3.024 \text{ g H}}{97.99 \text{ g H}_3\text{PO}_4} \times 100 = 3.08\% \text{ H}$$

$$\text{percent P} = \frac{30.97 \text{ g P}}{97.99 \text{ g H}_3\text{PO}_4} \times 100 = 31.61\% \text{ P}$$

$$\text{percent O} = \frac{64.00 \text{ g O}}{97.99 \text{ g H}_3\text{PO}_4} \times 100 = 65.31\% \text{ O}$$

55. Which has the larger percent by mass of sulfur,  $\text{H}_2\text{SO}_3$  or  $\text{H}_2\text{S}_2\text{O}_8$ ?

Steps 1 and 2: Assume 1 mole; calculate molar mass of  $\text{H}_2\text{SO}_3$ .

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$1 \text{ mol S} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = 32.06 \text{ g S}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{48.00 \text{ g O}}$$

$$\text{molar mass H}_2\text{SO}_3 = 82.08 \text{ g/mol}$$

Step 3: Determine percent by mass of S.

$$\text{percent S} = \frac{32.06 \text{ g S}}{82.08 \text{ g H}_2\text{SO}_3} \times 100 = 39.06\% \text{ S}$$

Repeat steps 1 and 2 for  $\text{H}_2\text{S}_2\text{O}_8$ . Assume 1 mole; calculate molar mass of  $\text{H}_2\text{S}_2\text{O}_8$ .

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$2 \text{ mol S} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = 64.12 \text{ g S}$$

$$8 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{128.00 \text{ g O}}$$

$$\text{molar mass H}_2\text{S}_2\text{O}_8 = 194.14 \text{ g/mol}$$

Step 3: Determine percent by mass of S.

$$\text{percent S} = \frac{64.12 \text{ g S}}{194.14 \text{ g H}_2\text{S}_2\text{O}_8} \times 100$$

$$= 33.03\% \text{ S}$$

$\text{H}_2\text{SO}_3$  has a larger percent by mass of S.

56. Calcium chloride ( $\text{CaCl}_2$ ) is sometimes used as a de-icer. Calculate the percent by mass of each element in  $\text{CaCl}_2$ .

Steps 1 and 2: Assume 1 mole; calculate molar mass of  $\text{CaCl}_2$ .

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g Ca}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = \underline{70.90 \text{ g Cl}}$$

$$\text{molar mass CaCl}_2 = 110.98 \text{ g/mol}$$

Step 3: Determine percent by mass of each element.

$$\begin{aligned} \text{percent Ca} &= \frac{40.08 \text{ g Ca}}{110.98 \text{ g CaCl}_2} \times 100 \\ &= 36.11\% \text{ Ca} \end{aligned}$$

$$\begin{aligned} \text{percent Cl} &= \frac{70.90 \text{ g Cl}}{110.98 \text{ g CaCl}_2} \times 100 \\ &= 63.89\% \text{ Cl} \end{aligned}$$

**57. Challenge** Sodium sulfate is used in the manufacture of detergents.

- a. Identify each of the component elements of sodium sulfate, and write the compound's chemical formula.

sodium, sulfur, and oxygen;  $\text{Na}_2\text{SO}_4$

- b. Identify the compound as ionic or covalent.  
ionic

- c. Calculate the percent by mass of each element in sodium sulfate.

Steps 1 and 2: Assume 1 mole; calculate molar mass of  $\text{Na}_2\text{SO}_4$ .

$$2 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 45.98 \text{ g Na}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass Na}_2\text{SO}_4 = 142.05 \text{ g/mol}$$

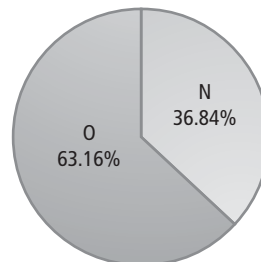
Step 3: Determine percent by mass of each element.

$$\begin{aligned} \text{percent Na} &= \frac{45.98 \text{ g Na}}{142.05 \text{ g Na}_2\text{SO}_4} \times 100 \\ &= 32.37\% \text{ Na} \end{aligned}$$

$$\begin{aligned} \text{percent S} &= \frac{32.07 \text{ g S}}{142.05 \text{ g Na}_2\text{SO}_4} \times 100 \\ &= 22.58\% \text{ S} \end{aligned}$$

$$\begin{aligned} \text{percent O} &= \frac{64.00 \text{ g O}}{142.05 \text{ g Na}_2\text{SO}_4} \times 100 \\ &= 45.05\% \text{ O} \end{aligned}$$

- 58.** The circle graph at the right gives the percent composition for a blue solid. What is the empirical formula for this solid?



Step 1: Assume 100 g sample; calculate moles of each element.

$$36.84 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 2.630 \text{ mol N}$$

$$63.16 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.948 \text{ mol O}$$

Step 2: Calculate mole ratios.

$$\frac{2.630 \text{ mol N}}{2.630 \text{ mol N}} = \frac{1.000 \text{ mol N}}{1.000 \text{ mol N}} = \frac{1 \text{ mol N}}{1 \text{ mol N}}$$

$$\frac{3.948 \text{ mol O}}{2.630 \text{ mol O}} = \frac{1.500 \text{ mol O}}{1.000 \text{ mol N}} = \frac{1.5 \text{ mol O}}{1 \text{ mol N}}$$

The simplest ratio is 1 mol N: 1.5 mol O.

Step 3: Convert decimal fraction to whole number.

In this case, multiply by 2 because  $1.5 \times 2 = 3$ . Therefore, the empirical formula is  $\text{N}_2\text{O}_3$ .

- 59.** Determine the empirical formula for a compound that contains 35.98% aluminum and 64.02% sulfur.

Step 1: Assume 100 g sample; calculate moles of each element.

$$35.98 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 1.334 \text{ mol Al}$$

$$64.02 \text{ g S} \times \frac{1 \text{ mol S}}{32.06 \text{ g S}} = 1.996 \text{ mol S}$$

Step 2: Calculate mole ratios.

$$\frac{1.334 \text{ mol Al}}{1.334 \text{ mol Al}} = \frac{1.000 \text{ mol Al}}{1.000 \text{ mol Al}} = \frac{1 \text{ mol Al}}{1 \text{ mol Al}}$$

$$\frac{1.996 \text{ mol S}}{1.334 \text{ mol Al}} = \frac{1.500 \text{ mol S}}{1.000 \text{ mol Al}} = \frac{1.5 \text{ mol S}}{1 \text{ mol Al}}$$

The simplest ratio is 1 mol Al: 1.5 mol S.

Step 3: Convert decimal fraction to whole number.

In this case, multiply by 2 because  $1.5 \times 2 = 3$ . Therefore, the empirical formula is  $\text{Al}_2\text{S}_3$ .

60. Propane is a hydrocarbon, a compound composed only of carbon and hydrogen. It is 81.82% carbon and 18.18% hydrogen. What is the empirical formula?

Step 1: Assume 100 g sample; calculate moles of each element.

$$81.82 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.813 \text{ mol C}$$

$$18.18 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 18.04 \text{ mol H}$$

Step 2: Calculate mole ratios.

$$\frac{6.813 \text{ mol C}}{6.813 \text{ mol C}} = \frac{1.000 \text{ mol C}}{1.000 \text{ mol C}} = \frac{1 \text{ mol C}}{1 \text{ mol C}}$$

$$\frac{18.04 \text{ mol H}}{6.813 \text{ mol C}} = \frac{2.649 \text{ mol H}}{1.000 \text{ mol C}} = \frac{2.65 \text{ mol H}}{1 \text{ mol C}}$$

The simplest ratio is 1 mol: 2.65 mol H.

Step 3: Convert decimal fraction to whole number.

In this case, multiply by 3 because  $2.65 \times 3 = 7.95 \approx 8$ . Therefore, the empirical formula is  $\text{C}_3\text{H}_8$ .

61. **Challenge** Aspirin is the world's most-often used medication. The chemical analysis of aspirin indicates that the molecule is 60.00% carbon, 4.44% hydrogen, and 35.56% oxygen. Determine the empirical formula for aspirin.

Step 1: Assume 100 g sample; calculate moles of each element.

$$60.00 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.00 \text{ mol C}$$

$$4.44 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.40 \text{ mol H}$$

$$35.56 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.22 \text{ mol O}$$

Step 2: Calculate mole ratios.

$$\frac{5.00 \text{ mol C}}{2.22 \text{ mol O}} = \frac{2.25 \text{ mol C}}{1.00 \text{ mol O}} = \frac{2.25 \text{ mol C}}{1 \text{ mol O}}$$

$$\frac{4.40 \text{ mol C}}{2.22 \text{ mol O}} = \frac{1.98 \text{ mol H}}{1.00 \text{ mol O}} = \frac{2 \text{ mol H}}{1 \text{ mol O}}$$

$$\frac{2.22 \text{ mol O}}{2.22 \text{ mol O}} = \frac{1.00 \text{ mol O}}{1.00 \text{ mol O}} = \frac{1 \text{ mol O}}{1 \text{ mol O}}$$

The simplest ratio is 2.25 mol C: 2 mol H: 1 mol O.

Step 3: Convert decimal fraction to whole number.

In this case, multiply by 4 because  $2.25 \times 4 = 9$ . Therefore, the empirical formula is  $\text{C}_9\text{H}_8\text{O}_4$ .

62. A compound was found to contain 49.98 g of carbon and 10.47 g of hydrogen. The molar mass of the compound is 58.12 g/mol. Determine the molecular formula.

Step 1: Assume 100 g sample; calculate moles of each element.

$$49.98 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.162 \text{ mol C}$$

$$10.47 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 10.39 \text{ mol H}$$

Step 2: Calculate mole ratios.

$$\frac{4.162 \text{ mol C}}{4.162 \text{ mol C}} = \frac{1.000 \text{ mol C}}{1.000 \text{ mol C}} = \frac{1 \text{ mol C}}{1 \text{ mol C}}$$

$$\frac{10.39 \text{ mol H}}{4.162 \text{ mol C}} = \frac{2.50 \text{ mol H}}{1.000 \text{ mol C}} = \frac{2.5 \text{ mol H}}{1 \text{ mol C}}$$

The simplest ratio is 1 mol C: 2.5 mol H.

Because  $2.5 \times 2 = 5$ , the empirical formula is  $\text{C}_2\text{H}_5$ .

Step 3: Calculate the molar mass of the empirical formula.

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g}$$

$$5 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 5.040 \text{ g}$$

$$\text{molar mass } \text{C}_2\text{H}_5 = 29.06 \text{ g/mol}$$

Step 4: Determine whole number multiplier.

$$\frac{58.12 \text{ g/mol}}{29.06 \text{ g/mol}} = 2.000$$

The molecular formula is  $\text{C}_4\text{H}_{10}$ .

63. A colorless liquid composed of 46.68% nitrogen and 53.32% oxygen has a molar mass of 60.01 g/mol. What is the molecular formula?

Step 1: Assume 100 g sample; calculate moles of each element.

$$46.68 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 3.332 \text{ mol N}$$

$$53.32 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.333 \text{ mol O}$$

Step 2: Calculate mole ratios.

$$\frac{3.332 \text{ mol N}}{3.332 \text{ mol N}} = \frac{1.000 \text{ mol N}}{1.000 \text{ mol N}} = \frac{1 \text{ mol N}}{1 \text{ mol N}}$$

$$\frac{3.333 \text{ mol O}}{3.332 \text{ mol N}} = \frac{1.000 \text{ mol O}}{1.000 \text{ mol N}} = \frac{1 \text{ mol O}}{1 \text{ mol N}}$$

The simplest ratio is 1 mol N: 1 mol O.

The empirical formula is NO.

Step 3: Calculate the molar mass of the empirical formula.

$$1 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 14.01 \text{ g}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g}$$

molar mass NO = 30.01 g/mol

Step 4: Determine whole number multiplier.

$$\frac{60.01 \text{ g/mol}}{30.01 \text{ g/mol}} = 2.000$$

The molecular formula is N<sub>2</sub>O<sub>2</sub>.

64. When an oxide of potassium is decomposed, 19.55 g K and 4.00 g O are obtained. What is the empirical formula for the compound?

Step 1: Calculate moles of each element.

$$19.55 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 0.5000 \text{ mol K}$$

$$4.00 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.250 \text{ mol O}$$

Step 2: Calculate mole ratios.

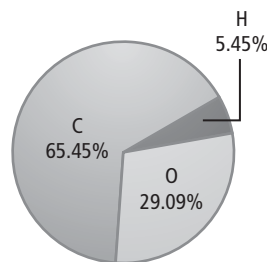
$$\frac{0.5000 \text{ mol K}}{0.250 \text{ mol O}} = \frac{2.00 \text{ mol K}}{1.00 \text{ mol O}} = \frac{2 \text{ mol K}}{1 \text{ mol O}}$$

$$\frac{0.250 \text{ mol O}}{0.250 \text{ mol O}} = \frac{1.00 \text{ mol O}}{1.00 \text{ mol O}} = \frac{1 \text{ mol O}}{1 \text{ mol O}}$$

The simplest ratio is 2 mol K: 1 mol O.

The empirical formula is K<sub>2</sub>O.

65. **Challenge** Analysis of a chemical used in photographic developing fluid yielded the percent composition data shown in the circle graph to the right. If the chemical's molar mass is 110.0 g/mol, what is its molecular formula?



Step 1: Assume 100 g sample; calculate moles of each element.

$$65.45 \text{ g C} \times \frac{1 \text{ mol C}}{12.04 \text{ g C}} = 5.450 \text{ mol C}$$

$$5.45 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.41 \text{ mol H}$$

$$29.09 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.818 \text{ mol O}$$

Step 2: Calculate mole ratios.

$$\frac{5.450 \text{ mol C}}{1.818 \text{ mol O}} = \frac{3.000 \text{ mol C}}{1.000 \text{ mol O}} = \frac{3 \text{ mol C}}{1 \text{ mol O}}$$

$$\frac{5.41 \text{ mol H}}{1.818 \text{ mol O}} = \frac{2.97 \text{ mol H}}{1.00 \text{ mol O}} = \frac{3 \text{ mol H}}{1 \text{ mol O}}$$

$$\frac{1.818 \text{ mol O}}{1.818 \text{ mol O}} = \frac{1.000 \text{ mol O}}{1.000 \text{ mol O}} = \frac{1 \text{ mol O}}{1 \text{ mol O}}$$

The empirical formula is C<sub>3</sub>H<sub>3</sub>O.

Step 3: Calculate the molar mass of the empirical formula.

$$3 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 36.03 \text{ g C}$$

$$3 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 3.024 \text{ g H}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

molar mass C<sub>3</sub>H<sub>3</sub>O = 55.05 g/mol

Step 4: Determine whole number multiplier.

$$\frac{110.0 \text{ g/mol}}{55.05 \text{ g/mol}} = 1.998, \text{ or } 2$$

The molecular formula is C<sub>6</sub>H<sub>6</sub>O<sub>2</sub>.

- 66. Challenge** Analysis of the pain reliever morphine yielded the data shown in the table. Determine the empirical formula of morphine.

| Element  | Mass (g) |
|----------|----------|
| carbon   | 17.900   |
| hydrogen | 1.680    |
| oxygen   | 4.225    |
| nitrogen | 1.228    |

Step 1: Calculate moles of each element.

$$17.900 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.490 \text{ mol C}$$

$$1.680 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.667 \text{ mol H}$$

$$4.225 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.2641 \text{ mol O}$$

$$1.228 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.08765 \text{ mol N}$$

Step 2: Calculate mole ratios.

$$\frac{0.08765 \text{ mol N}}{0.08765 \text{ mol N}} = \frac{1.000 \text{ mol N}}{1.000 \text{ mol N}} = \frac{1 \text{ mol N}}{1 \text{ mol N}}$$

$$\frac{1.490 \text{ mol C}}{0.08765 \text{ mol N}} = \frac{17.00 \text{ mol C}}{1.000 \text{ mol N}} = \frac{17 \text{ mol C}}{1 \text{ mol N}}$$

$$\frac{1.667 \text{ mol H}}{0.08765 \text{ mol N}} = \frac{19.02 \text{ mol H}}{1.000 \text{ mol N}} = \frac{19 \text{ mol H}}{1 \text{ mol N}}$$

$$\frac{0.2641 \text{ mol O}}{0.08765 \text{ mol N}} = \frac{3.013 \text{ mol O}}{1.000 \text{ mol N}} = \frac{3 \text{ mol O}}{1 \text{ mol N}}$$

The simplest ratio is 17 mol C: 19 mol H: 3 mol O: 1 mol N.

The empirical formula is  $\text{C}_{17}\text{H}_{19}\text{O}_3\text{N}$ .

## Section 10.4 Assessment

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- 67. Assess** A classmate tells you that experimental data shows a compound's molecular formula to be 2.5 times its empirical formula. Is he correct? Explain.

No, he is in error because the molecular formula must be a whole-number multiple of the empirical formula.

- 68. Calculate** Analysis of a compound composed of iron and oxygen yields 174.86 g of Fe and 75.14 g of O. What is the empirical formula for this compound?

Step 1: Calculate moles of each element

$$174.86 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 3.131 \text{ mol Fe}$$

$$75.14 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.696 \text{ mol O}$$

Step 2: Calculate mole ratios.

$$\frac{3.131 \text{ mol Fe}}{3.131 \text{ mol Fe}} = \frac{1.000 \text{ mol Fe}}{1.000 \text{ mol Fe}} = \frac{1 \text{ mol Fe}}{1 \text{ mol Fe}}$$

$$\frac{4.696 \text{ mol O}}{3.131 \text{ mol Fe}} = \frac{1.500 \text{ mol O}}{1.000 \text{ mol Fe}} = \frac{1.5 \text{ mol O}}{1 \text{ mol Fe}}$$

The simplest ratio is 1 mol Fe: 1.5 mol O.

Because  $1.5 \times 2 = 3$ , the empirical formula is  $\text{Fe}_2\text{O}_3$ .

- 69. Calculate** An oxide of aluminum contains 0.545 g of Al and 0.485 g of O. Find the empirical formula for the oxide.

Step 1: Calculate moles of each element.

$$0.545 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.0202 \text{ mol Al}$$

$$0.485 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0303 \text{ mol O}$$

Step 2: Calculate mole ratios.

$$\frac{0.0202 \text{ mol Al}}{0.0202 \text{ mol Al}} = \frac{1.00 \text{ mol Al}}{1.00 \text{ mol Al}} = \frac{1 \text{ mol Al}}{1 \text{ mol Al}}$$

$$\frac{0.0303 \text{ mol O}}{0.0202 \text{ mol Al}} = \frac{1.50 \text{ mol O}}{1.00 \text{ mol Al}} = \frac{1.5 \text{ mol O}}{1 \text{ mol Al}}$$

The simplest ratio is 1 mol Al: 1.5 mol O.

Because  $1.5 \times 2 = 3$ , the empirical formula is  $\text{Al}_2\text{O}_3$ .

- 70. Explain** how percent composition data for a compound are related to the masses of the elements in the compound.

Percent composition is numerically equal to the mass in grams of each element in a 100.0-g sample.



- 71. Explain** how you can find the mole ratio in a chemical compound.

The mole ratio is determined by calculating the moles of each element in the compound and dividing each number of moles by the smallest number of moles. It is sometimes necessary to multiply the ratio by an integer to obtain whole numbers.

- 72. Apply** The molar mass of a compound is twice that of its empirical formula. How are the compound's molecular and empirical formulas related?

The molecular formula is equal to twice the empirical formula.

- 73. Analyze** Hematite ( $\text{Fe}_2\text{O}_3$ ) and magnetite ( $\text{Fe}_3\text{O}_4$ ) are two ores used as sources of iron. Which ore provides the greater percent of iron per kilogram?

$$2 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 111.70 \text{ g Fe}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass Fe}_2\text{O}_3 = 159.70 \text{ g/mol}$$

$$3 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 167.55 \text{ g Fe}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass Fe}_3\text{O}_4 = 231.55 \text{ g/mol}$$

$$\begin{aligned} \text{percent by mass} &= \frac{111.70 \text{ g Fe}}{159.70 \text{ g Fe}_2\text{O}_3} \times 100 \\ &= 69.94\% \text{ Fe in Fe}_2\text{O}_3 \end{aligned}$$

$$\begin{aligned} \text{percent by mass} &= \frac{167.55 \text{ g Fe}}{231.55 \text{ g Fe}_3\text{O}_4} \times 100 \\ &= 72.36\% \text{ Fe in Fe}_3\text{O}_4 \end{aligned}$$

Hematite is 69.94% Fe, magnetite is 72.36% Fe. Magnetite contains a greater percentage of iron per kilogram than hematite.

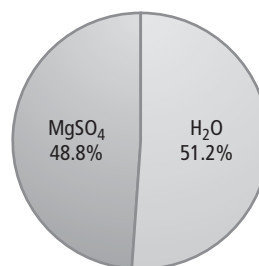
## Section 10.5 Formulas of Hydrates

pages 351–354

### Practice Problems

pages 353

- 74.** The composition of a hydrate is given in the pie chart shown at the right. What is the formula and name for this hydrate?



**Step 1:** Assume 100 g sample; calculate moles of each component.

$$48.8 \text{ g MgSO}_4 \times \frac{1 \text{ mol MgSO}_4}{120.38 \text{ g MgSO}_4} = 0.405 \text{ mol MgSO}_4$$

$$51.2 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 2.84 \text{ mol H}_2\text{O}$$

**Step 2:** Calculate mole ratios.

$$\begin{aligned} \frac{0.405 \text{ mol MgSO}_4}{0.405 \text{ mol MgSO}_4} &= \frac{1.00 \text{ mol MgSO}_4}{1.00 \text{ mol MgSO}_4} \\ &= \frac{1 \text{ mol MgSO}_4}{1 \text{ mol MgSO}_4} \end{aligned}$$

$$\begin{aligned} \frac{2.84 \text{ mol H}_2\text{O}}{0.405 \text{ mol MgSO}_4} &= \frac{7.01 \text{ mol H}_2\text{O}}{1.00 \text{ mol MgSO}_4} \\ &= \frac{7 \text{ mol H}_2\text{O}}{1 \text{ mol MgSO}_4} \end{aligned}$$

The formula of the hydrate is  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ . Its name is magnesium sulfate heptahydrate.

- 75. Challenge** An 11.75-g sample of a common hydrate of cobalt(II) chloride is heated. After heating, 0.0712 mol of anhydrous cobalt chloride remains. What is the formula and the name for this hydrate?

**Step 1:** Calculate the mass of  $\text{CoCl}_2$  remaining.

$$0.0712 \text{ mol CoCl}_2 \times \frac{129.83 \text{ g CoCl}_2}{1 \text{ mol CoCl}_2} = 9.24 \text{ g CoCl}_2$$

Step 2: Calculate the mass of water driven off.

mass of hydrated compound – mass of anhydrous compound remaining

$$= 11.75 \text{ g CoCl}_2 \cdot x\text{H}_2\text{O} - 9.24 \text{ g CoCl}_2 = 2.51 \text{ g H}_2\text{O}$$

Step 3: Calculate moles of each component.

$$9.24 \text{ g CoCl}_2 \times \frac{1 \text{ mol CoCl}_2}{129.83 \text{ g CoCl}_2}$$

$$= 0.0712 \text{ mol CoCl}_2$$

$$2.50 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.139 \text{ mol H}_2\text{O}$$

Step 4: Calculate mole ratios.

$$\frac{0.0712 \text{ mol CoCl}_2}{0.0712 \text{ mol CoCl}_2} = \frac{1.00 \text{ mol CoCl}_2}{1.00 \text{ mol CoCl}_2} = \frac{1 \text{ mol CoCl}_2}{1 \text{ mol CoCl}_2}$$

$$\frac{0.139 \text{ mol H}_2\text{O}}{0.0712 \text{ mol CoCl}_2} = \frac{1.95 \text{ mol H}_2\text{O}}{1.00 \text{ mol CoCl}_2} = \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CoCl}_2}$$

The formula of the hydrate is  $\text{CoCl}_2 \cdot 2\text{H}_2\text{O}$ . Its name is cobalt(II) chloride dihydrate.

## Section 10.5 Assessment

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**76. Summarize** the composition of a hydrate.

A hydrate is a solid ionic compound in which water molecules are trapped.

**77. Name** the compound that has the formula  $\text{SrCl}_2 \cdot 6\text{H}_2\text{O}$ .

strontium chloride hexahydrate

**78. Describe** the experimental procedure for determining the formula for a hydrate. Explain the reason for each step.

Mass an empty crucible. Add some hydrate and remass. Heat the crucible to drive out the water. Cool and remass. Determine the moles of the anhydrous compound. Subtract the mass of the crucible after heating from the mass of the crucible with the hydrate. The difference is the mass of the water lost. Determine the moles of water. Determine the simplest whole-number ratio of moles of water to moles of anhydrous compound, which will yield the formula of the hydrate.

**79. Apply** A hydrate contains 0.050 mol of  $\text{H}_2\text{O}$  to every 0.00998 mol of ionic compound. Write a generalized formula for the hydrate.

$\text{XY} \cdot 5\text{H}_2\text{O}$ , where XY represents the ionic compound

**80. Calculate** the mass of the water of hydration if a hydrate loses 0.025 mol of  $\text{H}_2\text{O}$  when heated.

$$\text{water of hydration} = 0.025 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.45 \text{ g H}_2\text{O}$$

**81. Arrange** these hydrates in order of increasing percent water content:  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ ,  $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ , and  $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ .

$$1 \text{ mol Mg} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 24.31 \text{ g Mg}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$11 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 176.00 \text{ g O}$$

$$14 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 14.112 \text{ g H}$$

$$\text{molar mass MgSO}_4 \cdot 7\text{H}_2\text{O} = 246.49 \text{ g}$$

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 137.33 \text{ g Ba}$$

$$10 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 160.00 \text{ g O}$$

$$18 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 18.144 \text{ g H}$$

$$\text{molar mass Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O} = 315.47 \text{ g}$$

$$1 \text{ mol Co} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} = 58.93 \text{ g Co}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 96.00 \text{ g O}$$

$$12 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 12.096 \text{ g H}$$

$$\text{molar mass CoCl}_2 \cdot 6\text{H}_2\text{O} = 237.93 \text{ g}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

molar mass  $\text{H}_2\text{O} = 18.02 \text{ g}$

$$\frac{7(18.02 \text{ g H}_2\text{O})}{246.49 \text{ g MgSO}_4 \cdot 7\text{H}_2\text{O}} \times 100$$

= 51.17%  $\text{H}_2\text{O}$  in  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

$$\frac{8(18.02 \text{ g H}_2\text{O})}{315.47 \text{ g Ba(OH)}_2 \cdot 8\text{H}_2\text{O}} \times 100$$

= 45.70%  $\text{H}_2\text{O}$  in  $\text{Ba(OH)}_2 \cdot 8\text{H}_2\text{O}$

$$\frac{6(18.02 \text{ g H}_2\text{O})}{237.93 \text{ g CoCl}_2 \cdot 6\text{H}_2\text{O}} \times 100$$

= 45.44%  $\text{H}_2\text{O}$  in  $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$

$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ ;  $\text{Ba(OH)}_2 \cdot 8\text{H}_2\text{O}$ ;  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

- 82. Apply** Explain how the hydrate in Figure 10.17 might be used as a means of roughly determining the probability of rain.

The hydrate is pink in moist air.

## Everyday Chemistry

page 355

### Writing in Chemistry

**Estimating** The estimating process used in this article is sometimes called a “back-of-the-envelope” calculation. Use this method to estimate the total mass of all of the students in your school.

Suppose a school has a student body of 500 and that the average mass of each student is about 60 kilograms. The total mass of the students is calculated as follows.

$$500 \text{ students} \times 60 \text{ kilograms/student} = 30,000 \text{ kg}$$

## Chapter 10 Assessment

pages 358–363

### Section 10.1

#### Mastering Concepts

- 83.** What is the numerical value of Avogadro’s number?

$$6.02 \times 10^{23}$$

- 84.** How many atoms of potassium does 1 mol of potassium contain?

$$6.02 \times 10^{23} \text{ potassium atoms}$$

- 85.** Compare a mole of Ag-108 and a mole of Pt-195 using atoms, protons, electrons, and neutrons.

Since they both contain one mole, you have  $6.02 \times 10^{23}$  particles of silver and  $6.02 \times 10^{23}$  particles of platinum. The difference would be in the number of protons, neutrons, and electrons present because silver contains 47 protons per atom, 47 electrons per atom and 61 neutrons per atom, while platinum contains 78 protons per atom, 78 electrons per atom and 117 neutrons per atom. A mole of platinum would contain more protons, electrons and neutrons but the same number of atoms.

- 86.** Why is the mole an important unit to chemists?

A mole allows a chemist to accurately measure the number of atoms, molecules, or formula units in a substance.

- 87. Currency** Examine the information in Table 10.2 and explain how rolls used to count pennies and dimes are similar to moles.

| Rolled-Coin Values |                          |
|--------------------|--------------------------|
| Coin               | Value of a Roll of Coins |
| Penny              | \$0.50                   |
| Dime               | \$5.00                   |

One roll of pennies always contains 50 pennies and one roll of dimes always contains 50 dimes. Each roll contains a specific number of coins. A mole also contains a specific number of particles,  $6.02 \times 10^{23}$  particles.

- 88.** Explain how Avogadro’s number is used as a conversion factor.

Avogadro’s number is the number of particles in one mole of a substance. It can be used to convert particles to moles or moles to particles.

**89. Conversion** Design a flowchart that could be used to help convert particles to moles or moles to particles.

Flowcharts will vary but should clearly show the use of proper conversion factors.

### Mastering Problems

**90.** Determine the number of representative particles in each substance.

a. 0.250 mol of silver

$$0.250 \text{ mol Ag} \times \frac{6.02 \times 10^{23} \text{ atoms Ag}}{1 \text{ mol Ag}} = 1.51 \times 10^{23} \text{ atoms Ag}$$

b.  $8.56 \times 10^{-3}$  mol of sodium chloride

$$8.56 \times 10^{-3} \text{ mol NaCl} \times \frac{6.02 \times 10^{23} \text{ formula units NaCl}}{1 \text{ mol NaCl}} = 5.15 \times 10^{21} \text{ formula units NaCl}$$

c. 35.3 mol of carbon dioxide

$$35.3 \text{ mol CO}_2 \times \frac{6.02 \times 10^{23} \text{ molecules CO}_2}{1 \text{ mol CO}_2} = 2.13 \times 10^{25} \text{ molecules CO}_2$$

d. 0.425 mol of nitrogen ( $\text{N}_2$ )

$$0.425 \text{ mol N}_2 \times \frac{6.02 \times 10^{23} \text{ molecules N}_2}{1 \text{ mol N}_2} = 2.56 \times 10^{23} \text{ molecules N}_2$$

**91.** Determine the number of representative particles in each substance.

a. 4.45 mol of  $\text{C}_6\text{H}_{12}\text{O}_6$

$$4.45 \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 2.68 \times 10^{24} \text{ molecules C}_6\text{H}_{12}\text{O}_6$$

b. 0.250 mol of  $\text{KNO}_3$

$$0.250 \text{ mol KNO}_3 \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \text{ mol}} = 1.51 \times 10^{23} \text{ formula units KNO}_3$$

c. 2.24 mol of  $\text{H}_2$

$$2.24 \text{ mol H}_2 \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.35 \times 10^{24} \text{ molecules of H}_2$$

d. 9.56 mol of Zn

$$9.56 \text{ mol Zn} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 5.76 \times 10^{24} \text{ atoms of Zn}$$

**92.** How many molecules are contained in each compound?

a. 1.35 mol of carbon disulfide ( $\text{CS}_2$ )

$$1.35 \text{ mol CS}_2 \times \frac{6.02 \times 10^{23} \text{ molecules CS}_2}{1 \text{ mol CS}_2} = 8.13 \times 10^{23} \text{ molecules CS}_2$$

b. 0.254 mol of diarsenic trioxide ( $\text{As}_2\text{O}_3$ )

$$0.254 \text{ mol As}_2\text{O}_3 \times \frac{6.02 \times 10^{23} \text{ molecules As}_2\text{O}_3}{1 \text{ mol As}_2\text{O}_3} = 1.53 \times 10^{23} \text{ molecules As}_2\text{O}_3$$

c. 1.25 mol of water

$$1.25 \text{ mol H}_2\text{O} \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 7.53 \times 10^{23} \text{ molecules H}_2\text{O}$$

d. 150.0 mol of HCl

$$150.0 \text{ mol HCl} \times \frac{6.02 \times 10^{23} \text{ molecules HCl}}{1 \text{ mol HCl}} = 9.030 \times 10^{25} \text{ molecules HCl}$$

**93.** Determine the number of moles in each substance.

a.  $3.25 \times 10^{20}$  atoms of lead

$$3.25 \times 10^{20} \text{ atoms Pb} \times \frac{1 \text{ mol Pb}}{6.02 \times 10^{23} \text{ atoms Pb}} = 5.39 \times 10^{-4} \text{ mol Pb}$$

b.  $4.96 \times 10^{24}$  molecules of glucose

$$4.96 \times 10^{24} \text{ molecules glucose} \times \frac{1 \text{ mol glucose}}{6.02 \times 10^{23} \text{ molecules glucose}} = 8.24 \text{ mol of glucose}$$

c.  $1.56 \times 10^{23}$  formula units of sodium hydroxide

$$1.56 \times 10^{23} \text{ formula units NaOH} \times \frac{1 \text{ mol NaOH}}{6.02 \times 10^{23} \text{ formula units NaOH}} = 2.59 \times 10^{-1} \text{ mol NaOH}$$

d.  $1.25 \times 10^{25}$  copper(II) ions

$$1.25 \times 10^{25} \text{ Cu}^{2+} \text{ ions} \times \frac{1 \text{ mol Cu}^{2+}}{6.02 \times 10^{23} \text{ Cu}^{2+} \text{ ions}} = 2.08 \times 10^1 \text{ mol Cu}^{2+} \text{ ions}$$

94. Perform the following conversions.

a.  $1.51 \times 10^{15}$  atoms of Si to mol of Si

$$1.51 \times 10^{15} \text{ atoms Si} \times \frac{1 \text{ mol Si}}{6.02 \times 10^{23} \text{ atoms Si}} = 2.51 \times 10^{-9} \text{ mol Si}$$

b.  $4.25 \times 10^{-2}$  mol of  $\text{H}_2\text{SO}_4$  to molecules of  $\text{H}_2\text{SO}_4$

$$4.25 \times 10^{-2} \text{ mol H}_2\text{SO}_4 \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 2.56 \times 10^{22} \text{ molecules H}_2\text{SO}_4$$

c.  $8.95 \times 10^{25}$  molecules of  $\text{CCl}_4$  to mol of  $\text{CCl}_4$

$$8.95 \times 10^{25} \text{ molecules CCl}_4 \times \frac{1 \text{ mol CCl}_4}{6.02 \times 10^{23} \text{ molecules CCl}_4} = 1.49 \times 10^2 \text{ mol CCl}_4$$

d. 5.90 mol of Ca to atoms of Ca

$$5.90 \text{ mol Ca} \times \frac{6.02 \times 10^{23} \text{ atoms Ca}}{1 \text{ mol Ca}} = 3.55 \times 10^{24} \text{ atoms Ca}$$

95. How many moles contain the given quantity?

a.  $1.25 \times 10^{15}$  molecules of carbon dioxide

$$1.25 \times 10^{15} \text{ molecules CO}_2 \times \frac{1 \text{ mol CO}_2}{6.02 \times 10^{23} \text{ molecules CO}_2} = 2.08 \times 10^{-9} \text{ mol CO}_2$$

b.  $3.59 \times 10^{21}$  formula units of sodium nitrate

$$3.59 \times 10^{21} \text{ formula units NaNO}_3 \times \frac{1 \text{ mol NaNO}_3}{6.02 \times 10^{23} \text{ formula units NaNO}_3} = 5.96 \times 10^{-3} \text{ mol NaNO}_3$$

c.  $2.89 \times 10^{27}$  formula units of calcium carbonate

$$2.89 \times 10^{27} \text{ formula units CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{6.02 \times 10^{23} \text{ formula units CaCO}_3} = 4.80 \times 10^3 \text{ mol CaCO}_3$$

96. **RDA of Selenium** The recommended daily allowance (RDA) of selenium in your diet is  $8.87 \times 10^{-4}$  mol. How many atoms of selenium is this?

$$8.87 \times 10^{-4} \text{ mol Se} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol Se}} = 5.34 \times 10^{20} \text{ atoms Se}$$



**Solution A**  
0.250 mol  
 $\text{Cu}^{2+}$  ions

**Solution B**  
0.130 mol  
 $\text{Ca}^{2+}$  ions

97. The two solutions shown in **Figure 10.19** are mixed. What is the total number of metal ions in the mixture?

$$0.250 \text{ mol Ca}^{2+} \text{ ions} + 0.130 \text{ mol Ca}^{2+} \text{ ion} = 0.380 \text{ mol metal ions}$$

$$0.380 \text{ mol metal ions} \times \frac{6.02 \times 10^{23} \text{ mol metal ions}}{1 \text{ mol metal ions}} = 2.29 \times 10^{23} \text{ Cu}^{2+} \text{ and Ca}^{2+} \text{ ions}$$

98. **Jewelry** A bracelet containing 0.200 mol metal atoms is 75% gold. How many particles of gold atoms are in the bracelet?

$$0.200 \text{ mol metal atoms} \times 0.75 = 0.150 \text{ mol Au}$$

$$0.150 \text{ mol Au} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \text{ mol Au}} = 9.03 \times 10^{22} \text{ atoms Au}$$

99. **Snowflakes** A snowflake contains  $1.9 \times 10^{18}$  molecules of water. How many moles of water does it contain?

$$1.9 \times 10^{18} \text{ molecules H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{6.02 \times 10^{23} \text{ molecules H}_2\text{O}} = 3.2 \times 10^{-6} \text{ mol H}_2\text{O}$$

- 100.** If you could count two atoms every second, how long would it take you to count a mole of atoms? Assume that you counted continually for 24 hours every day. How does the time you calculated compare with the age of Earth, which is estimated to be  $4.5 \times 10^9$  years old?

$$6.02 \times 10^{23} \text{ particles} \times \frac{1 \text{ s}}{2 \text{ particles}} = 3.01 \times 10^{23} \text{ s}$$

$$3.01 \times 10^{23} \text{ s} \times \frac{1 \text{ h}}{3600 \text{ s}} \times \frac{1 \text{ day}}{24 \text{ h}} \times \frac{1 \text{ y}}{365 \text{ day}}$$

$$= 2.1 \times 10^6$$

$$= 9.5 \times 10^{15} \text{ yr; about 2 million times longer}$$

- 101. Chlorophyll** The green color of leaves is due to the presence of chlorophyll,  $C_{55}H_{72}O_5N_4Mg$ . A fresh leaf was found to have  $1.5 \times 10^{-5}$  mol of chlorophyll per  $\text{cm}^2$ . How many chlorophyll molecules can be found in  $1 \text{ cm}^2$ ?

$$\frac{1.5 \times 10^{-5} \text{ mol chlorophyll}}{1 \text{ cm}^2}$$

$$\times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol chlorophyll}}$$

$$= 9.03 \times 10^{18} \text{ chlorophyll molecules/cm}^2$$

## Section 10.2

### Mastering Concepts

- 102.** Explain the difference between atomic mass (amu) and molar mass (g).  
**Atomic mass (amu) is the mass of an individual particle (atom, molecule). Molar mass (g) is the mass of a mole of particles.**
- 103.** Which contains more atoms, a mole of silver atoms or a mole of gold atoms? Explain your answer.  
**They both contain the same number of atoms because a mole of anything contains  $6.02 \times 10^{23}$  representative particles.**
- 104.** Which has more mass, a mole of potassium or a mole of sodium? Explain your answer.  
**The molar mass of K is 39.098 g/mol; the molar mass of Na is 22.990 g/mol. Thus, a mole of K has a greater mass.**

- 105.** Explain how you would convert from number of atoms of a specific element to its mass.

**Convert the number of atoms to moles, then multiply moles by the molar mass of the element.**

- 106.** Discuss the relationships that exist between the mole, molar mass, and Avogadro's number.

**Molar mass is the mass in grams of one mole of any pure substance. Avogadro's number is the number of representative particles in one mole. The mass of  $6.02 \times 10^{23}$  representative particles of a substance is the molar mass of the substance.**

- 107. Barbed Wire** Barbed wire is often made of steel, which is primarily iron, and coated with zinc. Compare the number of particles and the mass of 1 mol of each.

**The mole of iron and the mole of zinc both contain  $6.02 \times 10^{23}$  particles but a mole of iron has a mass of 55.85 g/mol and a mole of zinc has a mass of 65.39 g/mol.**

### Mastering Problems

- 108.** Calculate the mass of each element.

- a. 5.22 mol of He

$$5.22 \text{ mol He} \times \frac{4.00 \text{ g He}}{1 \text{ mol He}} = 20.9 \text{ g He}$$

- b. 0.0455 mol of Ni

$$0.0455 \text{ mol Ni} \times \frac{58.69 \text{ g Ni}}{1 \text{ mol Ni}} = 2.67 \text{ g Ni}$$

- c. 2.22 mol of Ti

$$2.22 \text{ mol Ti} \times \frac{47.87 \text{ g Ti}}{1 \text{ mol Ti}} = 106 \text{ g Ti}$$

- d. 0.00566 mol of Ge

$$0.00566 \text{ mol Ge} \times \frac{72.61 \text{ g Ge}}{1 \text{ mol Ge}} = 0.411 \text{ g Ge}$$

- 109.** Perform the following conversions.

- a. 3.50 mol of Li to g of Li

$$3.50 \text{ mol Li} \times \frac{6.94 \text{ g Li}}{1 \text{ mol Li}} = 24.3 \text{ g Li}$$

- b. 7.65 g of Co to mol of Co

$$7.65 \text{ g Co} \times \frac{1 \text{ mol Co}}{58.93 \text{ g Co}} = 0.130 \text{ mol Co}$$

- c. 5.62 g of Kr to mol of Kr

$$5.62 \text{ g Kr} \times \frac{1 \text{ mol Kr}}{83.80 \text{ g Kr}} = 0.0671 \text{ mol Kr}$$

- d. 0.0550 mol of As to g of As

$$0.0550 \text{ mol As} \times \frac{74.92 \text{ g As}}{1 \text{ mol As}} = 4.12 \text{ g As}$$

110. Determine the mass in grams of each element.

- a.  $1.33 \times 10^{22}$  mol of Sb

$$1.33 \times 10^{22} \text{ mol Sb} \times \frac{121.76 \text{ g Sb}}{1 \text{ mol Sb}} = 1.62 \times 10^{24} \text{ g Sb}$$

- b.  $4.75 \times 10^{14}$  mol of Pt

$$4.75 \times 10^{14} \text{ mol Pt} \times \frac{195.08 \text{ g Pt}}{1 \text{ mol Pt}} = 9.27 \times 10^{16} \text{ g Pt}$$

- c.  $1.22 \times 10^{23}$  mol of Ag

$$1.22 \times 10^{23} \text{ mol Ag} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} = 1.32 \times 10^{25} \text{ g Ag}$$

- d.  $9.85 \times 10^{24}$  mol of Cr

$$9.85 \times 10^{24} \text{ mol Cr} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}} = 5.12 \times 10^{26} \text{ g Cr}$$

111. Complete Table 10.3.

| Mass, Mole, and Particle Data |               |                                 |
|-------------------------------|---------------|---------------------------------|
| Mass                          | Moles         | Particles                       |
| 88.7 g Mg                     | 3.65 mol Mg   | $2.20 \times 10^{24}$ atoms Mg  |
| 29.54 g Cr                    | 0.5681 mol Cr | $3.420 \times 10^{23}$ atoms Cr |
| 1820 g P                      | 58.8 mol P    | $3.54 \times 10^{25}$ atoms P   |
| 42.6 g As                     | 0.568 mol As  | $3.42 \times 10^{23}$ atoms As  |

Solutions:

$$3.65 \text{ mol Mg} \times (24.31 \text{ g/mol}) = 88.7 \text{ g Mg}$$

$$3.65 \text{ mol Mg} \times (6.02 \times 10^{23} \text{ atoms/mol}) = 2.20 \times 10^{24} \text{ atoms Mg}$$

$$29.54 \text{ g Cr} \times (1 \text{ mol}/52.00 \text{ g}) = 0.5681 \text{ mol Cr}$$

$$0.5681 \text{ mol Cr} \times (6.02 \times 10^{23} \text{ atoms/mol}) = 3.420 \times 10^{23} \text{ atoms Cr}$$

$$3.54 \times 10^{25} \text{ atoms P} \times (1 \text{ mol}/6.02 \times 10^{23} \text{ atoms}) = 58.8 \text{ mol P}$$

$$58.8 \text{ mol P} \times (30.97 \text{ g/mol}) = 1820 \text{ g P}$$

$$0.568 \text{ mol As} \times (74.92 \text{ g/mol}) = 42.6 \text{ g As}$$

$$0.568 \text{ mol As} \times (6.02 \times 10^{23} \text{ atoms/mol}) = 3.42 \times 10^{23} \text{ atoms As}$$

112. Convert each to mass in grams.

- a.  $4.22 \times 10^{15}$  atoms U

$$4.22 \times 10^{15} \text{ atoms U} \times \frac{1 \text{ mol U}}{6.02 \times 10^{23} \text{ atoms U}} \times \frac{238.03 \text{ g U}}{1 \text{ mol U}} = 1.67 \times 10^{-6} \text{ g U}$$

- b.  $8.65 \times 10^{25}$  atoms H

$$8.65 \times 10^{25} \text{ atoms H} \times \frac{1 \text{ mol H}}{6.02 \times 10^{23} \text{ atoms H}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 145 \text{ g H}$$

- c.  $1.25 \times 10^{22}$  atoms O

$$1.25 \times 10^{22} \text{ atoms O} \times \frac{1 \text{ mol O}}{6.02 \times 10^{23} \text{ atoms O}} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 0.332 \text{ g O}$$

- d.  $4.44 \times 10^{23}$  atoms Pb

$$4.44 \times 10^{23} \text{ atoms Pb} \times \frac{1 \text{ mol Pb}}{6.02 \times 10^{23} \text{ atoms Pb}} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 153 \text{ g Pb}$$

113. Calculate the number of atoms of each element.

- a. 25.8 g of Hg

$$25.8 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} \times \frac{6.02 \times 10^{23} \text{ atoms Hg}}{1 \text{ mol Hg}} = 7.74 \times 10^{22} \text{ atoms Hg}$$

b. 0.0340 g of Zn

$$\begin{aligned} 0.0340 \text{ g Zn} &\times \frac{1 \text{ mol Zn}}{63.59 \text{ g Zn}} \\ &\times \frac{6.02 \times 10^{23} \text{ atoms Zn}}{1 \text{ mol Zn}} \\ &= 3.13 \times 10^{20} \text{ atoms Zn} \end{aligned}$$

c. 150 g of Ar

$$\begin{aligned} 150 \text{ g Ar} &\times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}} \times \frac{6.02 \times 10^{23} \text{ atoms Ar}}{1 \text{ mol Ar}} \\ &= 2.3 \times 10^{24} \text{ atoms Ar} \end{aligned}$$

d. 0.124 g of Mg

$$\begin{aligned} 0.124 \text{ g Mg} &\times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \\ &\times \frac{6.02 \times 10^{23} \text{ atoms Mg}}{1 \text{ mol Mg}} \\ &= 3.07 \times 10^{21} \text{ atoms Mg} \end{aligned}$$

**114.** Arrange from least to most in moles:  $3.00 \times 10^{24}$  atoms Ne, 4.25 mole Ar,  $2.69 \times 10^{24}$  atoms Xe, 65.96 g Kr.

Convert atoms Ne, atoms Xe, and g Kr to moles:

$$\begin{aligned} \text{Ne: } &3.00 \times 10^{24} \text{ atoms} \\ \text{Ne} &\times \frac{1 \text{ mol Ne}}{6.02 \times 10^{23} \text{ atoms Ne}} = 4.98 \text{ mol Ne} \end{aligned}$$

$$\begin{aligned} \text{Xe: } &2.69 \times 10^{24} \text{ atoms Xe} \\ &\times \frac{1 \text{ mol Xe}}{6.02 \times 10^{23} \text{ atoms Xe}} = 4.47 \text{ mol Xe} \end{aligned}$$

$$\text{Kr: } 65.96 \text{ g Kr} \times \frac{1 \text{ mol Kr}}{83.80 \text{ g Kr}} = 0.7871 \text{ mol Kr}$$

$0.7871 \text{ mol Kr} < 4.25 \text{ mol Ar} < 4.47 \text{ mol Xe} < 4.98 \text{ mol Ne}$

**115. Balance Precision** A sensitive electronic balance can detect masses of  $1 \times 10^{-8}$  g. How many atoms of silver would be in a sample having this mass?

$$\begin{aligned} 1 \times 10^{-8} \text{ g Ag} &\times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} \\ &\times \frac{6.02 \times 10^{23} \text{ atoms Ag}}{1 \text{ mol Ag}} \\ &= 5 \times 10^{13} \text{ atoms Ag} \end{aligned}$$

**116.** A sample of a compound contains 3.86 g of sulfur and 4.08 g of vanadium. How many atoms of sulfur and vanadium does the compound contain?

$$\begin{aligned} 3.86 \text{ g S} &\times \frac{1 \text{ mol}}{32.07 \text{ g S}} \times \frac{6.02 \times 10^{23} \text{ atoms S}}{1 \text{ mol S}} \\ &= 7.25 \times 10^{22} \text{ atoms S} \end{aligned}$$

$$\begin{aligned} 4.08 \text{ g V} &\times \frac{1 \text{ mol}}{50.94 \text{ g V}} \times \frac{6.02 \times 10^{23} \text{ atoms V}}{1 \text{ mol V}} \\ &= 4.82 \times 10^{22} \text{ atoms V} \end{aligned}$$

**117.** Which has more atoms, 10.0 g of C or 10.0 g of Ca? How many atoms does each have?

$$\begin{aligned} 10.0 \text{ g C} &\times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{6.02 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} \\ &= 5.01 \times 10^{23} \text{ atoms C} \end{aligned}$$

$$\begin{aligned} 10.0 \text{ g Ca} &\times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \times \frac{6.02 \times 10^{23} \text{ atoms Ca}}{1 \text{ mol Ca}} \\ &= 1.50 \times 10^{23} \text{ atoms Ca} \end{aligned}$$

10.0 g C contains more atoms.

**118.** Which has more atoms, 10.0 mol of C or 10.0 mol of Ca? How many atoms does each have?

One mole of any substance contains  $6.02 \times 10^{23}$  representative particles. Thus, 10.0 moles of carbon and 10.0 moles of calcium contain the same number of atoms.

$$\begin{aligned} 10.0 \text{ mol} &\times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \\ &= 6.02 \times 10^{24} \text{ atoms} \end{aligned}$$

**119.** A mixture contains 0.250 mol of Fe and 1.20 g of C. What is the total number of atoms in the mixture?

$$\begin{aligned} 0.250 \text{ mol Fe} &\times \frac{6.02 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} \\ &= 1.51 \times 10^{23} \text{ atoms Fe} \end{aligned}$$

$$\begin{aligned} 1.20 \text{ g C} &\times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{6.02 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} \\ &= 0.601 \times 10^{23} \text{ atoms C} \end{aligned}$$

$$\begin{aligned} 1.51 \times 10^{23} \text{ atoms Fe} &+ 0.601 \times 10^{23} \text{ atoms C} \\ &= 2.11 \times 10^{23} \text{ total atoms} \end{aligned}$$



- 120. Respiration** Air contains several gases. When resting, every breath you take contains approximately 0.600 g of air. If argon makes up 0.934% of the air, calculate the number of argon atoms inhaled with each breath.

$$\frac{0.600 \text{ g air}}{1 \text{ breath}} \times \frac{0.934 \text{ g Ar}}{100 \text{ g air}} \times \frac{1 \text{ mol}}{39.95 \text{ g}} \\ \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = \frac{8.44 \times 10^{19} \text{ atoms Ar}}{1 \text{ breath}}$$

### Section 10.3

#### Mastering Concepts

- 121.** What information is provided by the formula for potassium chromate ( $\text{K}_2\text{CrO}_4$ )?

One mole of  $\text{K}_2\text{CrO}_4$  contains two moles of  $\text{K}^+$  ions and one mole of  $\text{CrO}_4^{2-}$  ions.

- 122.** In the formula for sodium phosphate ( $\text{Na}_3\text{PO}_4$ ), how many moles of sodium are represented? How many moles of phosphorus? How many moles of oxygen?

3 mol Na, 1 mol P, 4 mol O

- 123.** Explain how you determine the molar mass of a compound.

Molar mass is determined by multiplying the molar mass of each element present in the compound by its subscript and then adding these values.

- 124. Insect Repellent** Many insect repellents use DEET as the active ingredient. DEET was patented in 1946 and is effective against many biting insects. What must you know to determine the molar mass of DEET?

You must know the chemical formula of the substance to determine the molar mass.

- 125.** Why can molar mass be used as a conversion factor?

Molar mass is the mass of one mole of a compound. It can be used to convert moles of the compound to mass or mass of the compound to moles.

- 126.** List three conversion factors used in molar conversions.

$$\frac{1 \text{ mol}}{\text{number of grams}}, \\ \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mol}}, \\ \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ representative particles}}, \\ \frac{1 \text{ mol}}{\text{number of grams}}$$

- 127.** Which of these contains the most moles of carbon atoms per mole of the compound: ascorbic acid ( $\text{C}_6\text{H}_8\text{O}_6$ ), glycerin ( $\text{C}_3\text{H}_8\text{O}_3$ ), or vanillin ( $\text{C}_8\text{H}_8\text{O}_3$ )? Explain.

The formula for vanillin ( $\text{C}_8\text{H}_8\text{O}_3$ ) shows there are eight carbon atoms per molecule, more than ascorbic acid or glycerin.

#### Mastering Problems

- 128.** How many moles of oxygen atoms are contained in each compound?

- a. 2.50 mol of  $\text{KMnO}_4$

$$2.50 \text{ mol KMnO}_4 \times \frac{4 \text{ mol O}}{1 \text{ mol KMnO}_4} \\ = 10.0 \text{ mol O}$$

- b. 45.9 mol of  $\text{CO}_2$

$$45.9 \text{ mol CO}_2 \times \frac{2 \text{ mol O}}{1 \text{ mol CO}_2} = 91.8 \text{ mol O}$$

- c.  $1.25 \times 10^{-2}$  mol of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

$$1.25 \times 10^{-2} \text{ mol CuSO}_4 \cdot 5\text{H}_2\text{O} \\ \times \frac{9 \text{ mol O}}{1 \text{ mol CuSO}_4 \cdot 5\text{H}_2\text{O}} = 0.113 \text{ mol O}$$

- 129.** How many carbon tetrachloride ( $\text{CCl}_4$ ) molecules are in 3.00 mol of  $\text{CCl}_4$ ? How many carbon atoms? How many chlorine atoms? How many total atoms?

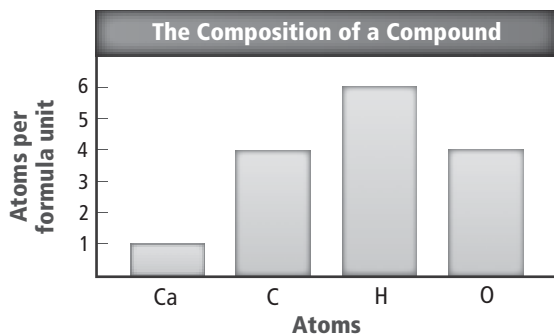
$$3.00 \text{ mol CCl}_4 \times \frac{6.02 \times 10^{23} \text{ molecules CCl}_4}{1 \text{ mol CCl}_4} \\ = 1.81 \times 10^{24} \text{ molecules CCl}_4$$

$$1.81 \times 10^{24} \text{ molecules CCl}_4 \times \frac{1 \text{ mol atoms C}}{1 \text{ molecule CCl}_4} \\ = 1.81 \times 10^{24} \text{ atoms C}$$

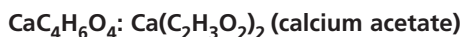
$$1.81 \times 10^{24} \text{ molecules CCl}_4 \times \frac{4 \text{ mol atoms Cl}}{1 \text{ molecule CCl}_4}$$

$$= 7.24 \times 10^{24} \text{ atoms Cl}$$

$$1.81 \times 10^{24} \text{ atoms C} + 7.24 \times 10^{24} \text{ atoms Cl} = 9.05 \times 10^{24} \text{ total atoms}$$



- 130.** The graph in **Figure 10.20** shows the numbers of atoms of each element in a compound. What is the compound's formula? What is its molar mass?



$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g Ca}$$

$$4 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 48.04 \text{ g C}$$

$$6 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 6.048 \text{ g H}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass} = 158.18 \text{ g/mol}$$

- 131.** Determine the molar mass of each compound.

a. nitric acid ( $\text{HNO}_3$ )

$$1 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.008 \text{ g H}$$

$$1 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 14.01 \text{ g N}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass} = 63.02 \text{ g/mol HNO}_3$$

b. ammonium nitrate ( $\text{NH}_4\text{NO}_3$ )

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 28.02 \text{ g N}$$

$$4 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 4.032 \text{ g H}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass} = 80.05 \text{ g/mol NH}_4\text{NO}_3$$

c. zinc oxide ( $\text{ZnO}$ )

$$1 \text{ mol Zn} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 65.39 \text{ g Zn}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$\text{molar mass} = 81.39 \text{ g/mol ZnO}$$

d. cobalt (II) chloride ( $\text{CoCl}_2$ )

$$1 \text{ mol Co} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} = 58.93 \text{ g Co}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass} = 129.83 \text{ g/mol CoCl}_2$$

- 132. Garlic** Determine the molar mass of allyl sulfide, the compound responsible for the smell of garlic. The chemical formula of allyl sulfide is  $(\text{C}_3\text{H}_5)_2\text{S}$ .

$$6 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 72.06 \text{ g C}$$

$$10 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 10.080 \text{ g H}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$\text{molar mass} = 114.21 \text{ g/mol } (\text{C}_3\text{H}_5)_2\text{S}$$

- 133.** How many moles are in 100.0 g of each compound?

a. dinitrogen oxide ( $\text{N}_2\text{O}$ )

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 28.02 \text{ g N}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$\text{molar mass} = 44.02 \text{ g/mol}$$

$$100.0 \text{ g N}_2\text{O} \times \frac{1 \text{ mol N}_2\text{O}}{44.02 \text{ g N}_2\text{O}} = 2.27 \text{ mol N}_2\text{O}$$

b. methanol (CH<sub>3</sub>OH)

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$4 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 4.032 \text{ g H}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{16.00 \text{ g O}}$$

$$\text{molar mass} = 32.04 \text{ g/mol}$$

$$100.0 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \\ = 3.12 \text{ mol CH}_3\text{OH}$$

134. What is the mass of each compound?

a.  $4.50 \times 10^{-2}$  mol of CuCl<sub>2</sub>

$$1 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 63.55 \text{ g Cu}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = \underline{70.90 \text{ g Cl}}$$

$$\text{molar mass} = 134.45 \text{ g/mol}$$

$$4.50 \times 10^{-2} \text{ mol CuCl}_2 \times \frac{134.45 \text{ g CuCl}_2}{1 \text{ mol CuCl}_2} \\ = 6.05 \text{ g CuCl}_2$$

b.  $1.25 \times 10^2$  mol of Ca(OH)<sub>2</sub>

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g Ca}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = \underline{32.00 \text{ g O}}$$

$$\text{molar mass} = 74.10 \text{ g/mol}$$

$$1.25 \times 10^2 \text{ mol Ca(OH)}_2 \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} \\ = 9.26 \times 10^2 \text{ g Ca(OH)}_2$$

135. **Acne** Benzoyl peroxide (C<sub>14</sub>H<sub>10</sub>O<sub>4</sub>) is a substance used as an acne medicine. What is the mass in grams of  $3.50 \times 10^{-2}$  mol C<sub>14</sub>H<sub>10</sub>O<sub>4</sub>?

$$14 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 168.14 \text{ g C}$$

$$10 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 10.080 \text{ g H}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass} = 168.14 + 10.080 + 64.00 \\ = 242.22 \text{ g/mol}$$

$$3.50 \times 10^{-2} \text{ mol} \times \frac{242.22 \text{ g}}{1 \text{ mol}} = 8.48 \text{ g benzoyl peroxide}$$

136. **Glass Etching** Hydrofluoric acid is a substance used to etch glass. Determine the mass of  $4.95 \times 10^{25}$  HF molecules.

$$1 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.008 \text{ g H}$$

$$1 \text{ mol F} \times \frac{19.00 \text{ g F}}{1 \text{ mol F}} = 19.00 \text{ g F}$$

$$\text{molar mass} = 1.008 + 19.00 = 20.01 \text{ g/mol}$$

$$4.95 \times 10^{25} \text{ molecules HF} \\ \times \frac{1 \text{ mol HF}}{6.02 \times 10^{23} \text{ molecules HF}} \times \frac{20.01 \text{ g HF}}{1 \text{ mol HF}} \\ = 1650 \text{ g HF}$$

137. What is the mass of a mole of electrons if one electron has a mass of  $9.11 \times 10^{-28}$  g?

$$6.02 \times 10^{23} \text{ electrons} \times \frac{9.11 \times 10^{-28} \text{ g}}{1 \text{ electron}} \\ = 5.48 \times 10^{-4} \text{ g}$$

138. How many moles of ions are in each compound?

a. 0.0200 g of AgNO<sub>3</sub>

$$1 \text{ mol Ag} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} = 107.87 \text{ g Ag}$$

$$1 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 14.01 \text{ g N}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass} = 107.87 + 14.01 + 48.00 \\ = 169.88 \text{ g/mol}$$

$$0.0200 \text{ g AgNO}_3 \times \frac{1 \text{ mol AgNO}_3}{169.88 \text{ g AgNO}_3} \\ \times \frac{2 \text{ mol ions}}{1 \text{ mol AgNO}_3} = 2.35 \times 10^{-4} \text{ mol ions}$$

b. 0.100 mol of K<sub>2</sub>CrO<sub>4</sub>

$$0.100 \text{ mol K}_2\text{CrO}_4 \times \frac{3 \text{ mol ions}}{1 \text{ mol K}_2\text{CrO}_4} = 0.300 \text{ mol ions}$$

c. 0.500 g of Ba(OH)<sub>2</sub>

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 137.33 \text{ g Ba}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass} = 137.33 + 32.00 + 2.016 \\ = 171.35 \text{ g/mol}$$

$$0.500 \text{ g Ba(OH)}_2 \times \frac{1 \text{ mol Ba(OH)}_2}{171.35 \text{ g Ba(OH)}_2} \\ \times \frac{3 \text{ mol ions}}{1 \text{ mol Ba(OH)}_2} = 8.75 \times 10^{-3} \text{ mol ions}$$

d.  $1.00 \times 10^{-9}$  mol of Na<sub>2</sub>CO<sub>3</sub>

$$1.00 \times 10^{-9} \text{ mol Na}_2\text{CO}_3 + \frac{3 \text{ mol ions}}{1 \text{ mol Na}_2\text{CO}_3} \\ = 3.00 \times 10^{-9} \text{ mol ions}$$

139. How many formula units are present in 500.0 g of lead(II) chloride?

$$1 \text{ mol Pb} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 207.2 \text{ g Pb}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass} = 207.2 + 70.90 = 278.1 \text{ g/mol}$$

$$500.0 \text{ g PbCl}_2 \times \frac{1 \text{ mol PbCl}_2}{278.1 \text{ g PbCl}_2} = 1.798 \text{ mol PbCl}_2$$

$$1.798 \text{ mol PbCl}_2 \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \text{ mol PbCl}_2} \\ = 1.082 \times 10^{24} \text{ formula units PbCl}_2$$

140. Determine the number of atoms in 3.50 g of gold.

$$3.50 \text{ g Au} \times \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \text{ mol Au}} \\ = 1.07 \times 10^{22} \text{ atoms Au}$$

141. Calculate the mass of  $3.62 \times 10^{24}$  molecules of glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>).

From a previous solution, the molar mass of glucose equals 180.16 g/mol.

$$3.62 \times 10^{24} \text{ molecules glucose} \\ \times \frac{1 \text{ mol glucose}}{6.02 \times 10^{23} \text{ molecules glucose}}$$

$$\times \frac{180.16 \text{ g glucose}}{1 \text{ mol glucose}} = 1080 \text{ g glucose}$$

142. Determine the number of molecules of ethanol (C<sub>2</sub>H<sub>5</sub>OH) in 47.0 g.

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g C}$$

$$6 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 6.048 \text{ g H}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$\text{molar mass} = 24.02 + 6.048 + 16.00 = 46.07 \text{ g/mol}$$

$$47.0 \text{ g ethanol} \times \frac{1 \text{ mol ethanol}}{46.07 \text{ g ethanol}}$$

$$\times \frac{6.02 \times 10^{23} \text{ molecules ethanol}}{1 \text{ mol ethanol}} \\ = 6.14 \times 10^{23} \text{ molecules ethanol}$$

143. What mass of iron(III) chloride contains  $2.35 \times 10^{23}$  chloride ions?

$$1 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 55.85 \text{ g Fe}$$

$$3 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 106.35 \text{ g Cl}$$

$$\text{molar mass} = 55.85 + 106.35 = 162.20 \text{ g/mol}$$

$$2.35 \times 10^{23} \text{ chloride ions}$$

$$\times \frac{1 \text{ mol chloride ions}}{6.02 \times 10^{23} \text{ chloride ions}}$$

$$\times \frac{1 \text{ mol FeCl}_3}{3 \text{ mol chloride ions}} \times \frac{162.20 \text{ g FeCl}_3}{1 \text{ mol FeCl}_3} \\ = 21.1 \text{ g FeCl}_3$$

144. How many moles of iron can be recovered from 100.0 kg of Fe<sub>3</sub>O<sub>4</sub>?

$$3 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 167.55 \text{ g Fe}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass} = 167.55 + 64.00 = 231.55 \text{ g/mol}$$

$$100.0 \text{ kg Fe}_3\text{O}_4 \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Fe}_3\text{O}_4}{231.55 \text{ g Fe}_3\text{O}_4} \times$$

$$\frac{3 \text{ mol Fe}}{1 \text{ mol Fe}_3\text{O}_4} = 1296 \text{ mol Fe}$$

- 145. Cooking** A common cooking vinegar is 5.0% acetic acid ( $\text{CH}_3\text{COOH}$ ). How many molecules of acetic acid are present in 25.0 g vinegar?

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g C}$$

$$4 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 4.032 \text{ g H}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass} = 24.02 + 4.032 + 32.00 = 60.05 \text{ g/mol}$$

$$0.050 \times 25.0 \text{ g vinegar} = 1.25 \text{ g CH}_3\text{COOH}$$

$$1.25 \text{ g CH}_3\text{COOH} \times \frac{1 \text{ mol CH}_3\text{COOH}}{60.05 \text{ g CH}_3\text{COOH}} \\ \times \frac{6.02 \times 10^{23} \text{ molecules CH}_3\text{COOH}}{1 \text{ mol CH}_3\text{COOH}} \\ = 1.3 \times 10^{22} \text{ molecules CH}_3\text{COOH}$$

- 146.** Calculate the moles of aluminum ions present in 250.0 g of aluminum oxide ( $\text{Al}_2\text{O}_3$ ).

$$250.0 \text{ g Al}_2\text{O}_3 \times \frac{1 \text{ mol Al}_2\text{O}_3}{101.96 \text{ g Al}_2\text{O}_3} \times \frac{2 \text{ mol Al}^{3+}}{1 \text{ mol Al}_2\text{O}_3} \\ = 4.90 \text{ mol Al}^{3+}$$

- 147.** Determine the number of chloride ions in 10.75 g of magnesium chloride.

$$1 \text{ mol Mg} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 24.31 \text{ g Mg}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass} = 24.31 + 70.90 = 94.31 \text{ g/mol}$$

$$10.75 \text{ g MgCl}_2 \times \frac{1 \text{ mol MgCl}_2}{94.31 \text{ g MgCl}_2} \times \frac{2 \text{ mol Cl}}{1 \text{ mol MgCl}_2} \\ \times \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol Cl}} = 1.37 \times 10^{23} \text{ ions Cl}^-$$

- 148. Pain Relief** Acetaminophen, a common aspirin substitute, has the formula  $\text{C}_8\text{H}_9\text{NO}_2$ . Determine the number of molecules of acetaminophen in a 500-mg tablet.

$$8 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 96.08 \text{ g C}$$

$$9 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 9.072 \text{ g H}$$

$$1 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 14.01 \text{ g N}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass} = 96.08 + 9.072 + 14.01 + 32.00 \\ = 151.16 \text{ g/mol}$$

$$500 \text{ mg C}_8\text{H}_9\text{NO}_2 \times \frac{1 \text{ g}}{10^3 \text{ mg}} \times \frac{1 \text{ mol C}_8\text{H}_9\text{NO}_2}{151.16 \text{ g C}_8\text{H}_9\text{NO}_2} \\ = 0.003 \text{ mol C}_8\text{H}_9\text{NO}_2 \\ \times \frac{6.02 \times 10^{23} \text{ molecules C}_8\text{H}_9\text{NO}_2}{1 \text{ mol C}_8\text{H}_9\text{NO}_2} \\ = 2 \times 10^{21} \text{ molecules C}_8\text{H}_9\text{NO}_2$$

- 149.** Calculate the number of sodium ions present in 25.0 g of sodium chloride.

$$25.0 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{1 \text{ mol Na}^+}{1 \text{ mol NaCl}} \times \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol Na}^+} \\ = 2.58 \times 10^{23} \text{ Na}^+ \text{ ions}$$

- 150.** Determine the number of oxygen atoms present in 25.0 g of carbon dioxide.

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass} = 12.01 + 32.00 = 44.01 \text{ g/mol}$$

$$25.0 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \\ = 0.568 \text{ mol CO}_2 \times \frac{2 \text{ mol O}}{1 \text{ mol CO}_2} \\ \times \frac{6.02 \times 10^{23} \text{ atoms O}}{1 \text{ mol O}} = 6.84 \times 10^{23} \text{ atoms O}$$

- 151. Espresso** There is  $1.00 \times 10^2$  mg of caffeine in a shot of espresso. The chemical formula of caffeine is  $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$ . Determine the moles of each element present in one shot of espresso.

$$\text{molar mass} = 194.19 \text{ g/mol}$$

$$0.100 \text{ g caffeine} \times \frac{1 \text{ mol caffeine}}{194.19 \text{ g caffeine}} \\ = 5.15 \times 10^{-4} \text{ mol caffeine}$$

$$\text{C} = 8(5.15 \times 10^{-4} \text{ mol}) = 4.12 \times 10^{-3} \text{ mol C}$$

$$\text{H} = 10(5.15 \times 10^{-4} \text{ mol}) = 5.15 \times 10^{-3} \text{ mol H}$$

$$\text{N} = 4(5.15 \times 10^{-4} \text{ mol}) = 2.06 \times 10^{-3} \text{ mol N}$$

$$\text{O} = 2(5.15 \times 10^{-4} \text{ mol}) = 1.03 \times 10^{-3} \text{ mol O}$$

152. The density of lead (Pb) is  $11.3 \text{ g/cm}^3$ . Calculate the volume of 1 mol of Pb.

$$1 \text{ mol Pb} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} \times \frac{1 \text{ cm}^3}{11.3 \text{ g Pb}} = 18.3 \text{ cm}^3$$

## Section 10.4

### Mastering Concepts

153. Explain what is meant by percent composition.

Percent composition is the percent by mass of each element in a compound.

154. What information must a chemist obtain in order to determine the empirical formula of an unknown compound?

the percent composition of the compound

155. What information must a chemist have to determine the molecular formula for a compound?

the percent composition of the compound and the molar mass

156. What is the difference between an empirical formula and a molecular formula? Provide an example.

An empirical formula is the smallest whole-number ratio of elements that make up a compound (CH). A molecular formula specifies the actual number of atoms of each element in one molecule or formula unit of the substance ( $\text{C}_6\text{H}_6$ ).

157. When can the empirical formula be the same as the molecular formula?

They are the same when the subscripts for each element present are the same.  $\text{Na}_2\text{O}$  is both the empirical and molecular formula for sodium oxide.

158. **Antibacterial Soap** Triclosan is an antibacterial agent included in detergents, dish soaps, laundry soaps, deodorants, cosmetics, lotions, creams, toothpastes, and mouthwashes. The chemical formula for triclosan is  $\text{C}_{12}\text{H}_7\text{Cl}_3\text{O}_2$ . What information did the chemist need to determine this formula?

The chemist must have either an analysis of the elements that compose the molecule or the percent composition of the compound to determine the chemical formula.

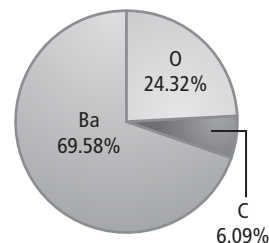
159. Which of the following formulas—NO,  $\text{N}_2\text{O}$ ,  $\text{NO}_2$ ,  $\text{N}_2\text{O}_4$ , and  $\text{N}_2\text{O}_5$ —represent the empirical and molecular formulas of the same compound? Explain your answer.

$\text{NO}_2$  is the empirical formula and  $\text{N}_2\text{O}_4$  is the molecular formula of the same compound.  $(\text{NO}_2)_2 = \text{N}_2\text{O}_4$

160. Do all pure samples of a given compound have the same percent composition? Explain.

Yes, for every pure substance, the percent by mass of each element is the same regardless of the size of the sample.

### Mastering Problems



161. The circle graph in Figure 10.21 shows the percent composition of a compound containing barium, carbon, and oxygen. What is the empirical formula of this compound?

$$\frac{69.78 \text{ g Ba}}{137.3 \text{ g/mol}} = \frac{0.5068 \text{ mol Ba}}{0.5068} = 1 \text{ mol Ba}$$

$$\frac{6.09 \text{ g C}}{12.01 \text{ g/mol}} = \frac{0.507 \text{ mol C}}{0.5068} = 1 \text{ mol C}$$

$$\frac{24.34 \text{ g O}}{16.00 \text{ g/mol}} = \frac{1.52 \text{ mol O}}{0.5068} = 3 \text{ mol O}$$

1 mol Ba:1 mol C:3 mol O; empirical formula =  $\text{BaCO}_3$

- 162. Iron** Three naturally occurring iron compounds are pyrite ( $\text{FeS}_2$ ), hematite ( $\text{Fe}_2\text{O}_3$ ), and siderite ( $\text{FeCO}_3$ ). Which contains the greatest percentage of iron?

$$1 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 55.85 \text{ g Fe}$$

$$2 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 64.14 \text{ g S}$$

$$\text{molar mass FeS}_2 = 55.85 + 64.14 = 119.99 \text{ g/mol}$$

$$1 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 111.70 \text{ g Fe}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass Fe}_2\text{O}_3 = 111.70 + 48.00 = 159.70 \text{ g/mol}$$

$$1 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 55.85 \text{ g Fe}$$

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 96.00 \text{ g O}$$

$$\text{molar mass FeCO}_3 = 55.85 + 12.01 + 96.00 \\ = 152.95 \text{ g/mol}$$

$$\text{percent mass} = \frac{55.85 \text{ g}}{119.99 \text{ g}} \times 100 = 46.55\% \text{ in FeS}_2$$

$$\text{percent mass} = \frac{111.70 \text{ g}}{159.70 \text{ g}} \times 100 = 69.944\% \text{ in Fe}_2\text{O}_3$$

$$\text{percent mass} = \frac{55.85 \text{ g}}{152.95 \text{ g}} \times 100 = 36.51\% \text{ in FeCO}_3$$

Hematite, with 69.944% Fe, has the greatest percentage of iron.

- 163.** Express the composition of each compound as the mass percent of its elements (percent composition).

a. sucrose ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ )

$$12 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 144.12 \text{ g C}$$

$$22 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 22.18 \text{ g H}$$

$$11 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 176.00 \text{ g O}$$

$$\text{molar mass} = 144.12 + 22.18 + 176.00 = \\ 342.30 \text{ g/mol}$$

$$\text{percent by mass} = \frac{144.12 \text{ g}}{342.30 \text{ g}} \times 100 = 42.10\% \text{ C}$$

$$\text{percent by mass} = \frac{22.18 \text{ g}}{342.30 \text{ g}} \times 100 = 6.480\% \text{ H}$$

$$\text{percent by mass} = \frac{176.00 \text{ g}}{342.30 \text{ g}} \times 100 = 51.42\% \text{ O}$$

b. aluminum sulfate ( $\text{Al}_2(\text{SO}_4)_3$ )

$$2 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 53.96 \text{ g Al}$$

$$3 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 96.21 \text{ g S}$$

$$12 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 192.00 \text{ g O}$$

$$\text{molar mass} = 53.96 + 96.21 + 192.00 = \\ 342.17 \text{ g/mol}$$

$$\text{percent by mass} = \frac{53.96 \text{ g}}{342.17 \text{ g}} \times 100 = 15.77\% \text{ Al}$$

$$\text{percent by mass} = \frac{96.21 \text{ g}}{342.17 \text{ g}} \times 100 = 28.12\% \text{ S}$$

$$\text{percent by mass} = \frac{192.00 \text{ g}}{342.17 \text{ g}} \times 100 = 56.11\% \text{ O}$$

c. magnetite ( $\text{Fe}_3\text{O}_4$ )

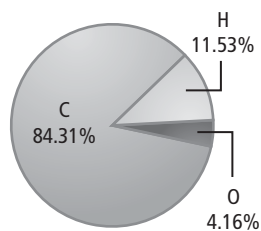
$$3 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 167.55 \text{ g Fe}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass} = 167.55 + 64.00 = 231.55 \text{ g/mol}$$

$$\text{percent by mass} = \frac{167.55 \text{ g}}{231.55 \text{ g}} \times 100 \\ = 72.360\% \text{ Fe}$$

$$\text{percent by mass} = \frac{64.00 \text{ g}}{231.55 \text{ g}} \times 100 = 27.64\% \text{ O}$$



Molar mass = 384 g/mol

- 164. Vitamin D<sub>3</sub>** Your body's ability to absorb calcium is aided by vitamin D<sub>3</sub>. Chemical analysis of vitamin D<sub>3</sub> yields the data shown in **Figure 10.22**. What are the empirical and molecular formulas for vitamin D<sub>3</sub>?

$$\frac{84.31 \text{ g C}}{12.011 \text{ g/mol}} = \frac{7.019 \text{ mol C}}{0.260} = 27 \text{ mol C}$$

$$\frac{11.53 \text{ g H}}{1.008 \text{ g/mol}} = \frac{11.44 \text{ mol H}}{0.260} = 44 \text{ mol H}$$

$$\frac{4.16 \text{ g O}}{16.00 \text{ g/mol}} = \frac{0.260 \text{ mol O}}{0.260} = 1 \text{ mol O}$$

27 mol C:44 mol H:1 mol O; empirical formula = C<sub>27</sub>H<sub>44</sub>O; molar mass is equal to 384 g/mol; 384/384 = 1; the empirical formula is equal to the molecular formula, C<sub>27</sub>H<sub>44</sub>O

- 165.** When a 30.98-g sample of phosphorus reacts with oxygen, a 71.00-g sample of phosphorus oxide is formed. What is the percent composition of the compound? What is the empirical formula for this compound?

$$30.98 \text{ g P} / 71.0 \text{ g} = 43.64 \% \text{ Phosphorus}$$

$$40.02 \text{ g O} / 71.0 \text{ g} = 56.36 \% \text{ Oxygen}$$

$$\frac{30.98 \text{ g P}}{30.97 \text{ g/mol}} = \frac{1 \text{ mol P}}{1.000} = 1 \text{ mol P}$$

$$\frac{40.02 \text{ g O}}{16.00 \text{ g/mol}} = \frac{2.501 \text{ mol O}}{1.000} = 2.5 \text{ mol O}$$

$$2(1 \text{ mol P}:2.5 \text{ mol O}) = 2 \text{ mol P}:5 \text{ mol O} = \text{P}_2\text{O}_5$$

- 166. Cholesterol** Heart disease is linked to high blood cholesterol levels. What is the percent composition of the elements in a molecule of cholesterol (C<sub>27</sub>H<sub>45</sub>OH)?

$$(27 \text{ mol C})(12.01 \text{ g/mol}) = 324.3 \text{ g C}$$

$$(46 \text{ mol H})(1.008 \text{ g/mol}) = 46.37 \text{ g H}$$

$$(1 \text{ mol O})(16.00 \text{ g/mol}) = 16.00 \text{ g O}$$

$$\text{total mass} = 386.67 \text{ g}$$

$$324.3 \text{ g C} / 386.67 \text{ g/mol} = 83.9\% \text{ C}$$

$$46.37 \text{ g H} / 386.67 \text{ g/mol} = 12.0\% \text{ H}$$

$$16.00 \text{ g O} / 386.67 \text{ g/mol} = 4.1\% \text{ O}$$

- 167.** Determine the empirical formula for each compound.

a. ethylene (C<sub>2</sub>H<sub>4</sub>)

Both subscripts can be divided by two. The empirical formula is CH<sub>2</sub>.

b. ascorbic acid (C<sub>6</sub>H<sub>8</sub>O<sub>6</sub>)

All subscripts can be divided by two. The empirical formula is C<sub>3</sub>H<sub>4</sub>O<sub>3</sub>.

c. naphthalene (C<sub>10</sub>H<sub>8</sub>)

Both subscripts can be divided by two. The empirical formula is C<sub>5</sub>H<sub>4</sub>.

- 168. Caffeine** The stimulant effect of coffee is due to caffeine, C<sub>8</sub>H<sub>10</sub>N<sub>4</sub>O<sub>2</sub>. Calculate the molar mass of caffeine. Determine its percent composition.

$$8 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 96.08 \text{ g C}$$

$$10 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 10.080 \text{ g H}$$

$$4 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 56.04 \text{ g N}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass} = 96.08 + 10.080 + 56.04 + 32.00 = 194.20 \text{ g/mol}$$

$$\text{percent by mass} = \frac{96.08 \text{ g}}{194.20 \text{ g}} \times 100 = 49.47\% \text{ C}$$

$$\text{percent by mass} = \frac{10.08 \text{ g}}{194.20 \text{ g}} \times 100 = 5.191\% \text{ H}$$

$$\text{percent by mass} = \frac{56.04 \text{ g}}{194.20 \text{ g}} \times 100 = 28.86\% \text{ N}$$

$$\text{percent by mass} = \frac{32.00 \text{ g}}{194.20 \text{ g}} \times 100 = 16.48\% \text{ O}$$



- 169.** Which of the titanium-containing minerals, rutile ( $\text{TiO}_2$ ) or ilmenite ( $\text{FeTiO}_3$ ), has the larger percentage of titanium?

$$1 \text{ mol Ti} \times \frac{47.87 \text{ g Ti}}{1 \text{ mol Ti}} = 47.87 \text{ g Ti}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass TiO}_2 = 47.87 + 32.00 = 79.87 \text{ g/mol}$$

$$1 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 55.85 \text{ g Fe}$$

$$1 \text{ mol Ti} \times \frac{47.87 \text{ g Ti}}{1 \text{ mol Ti}} = 47.87 \text{ g Ti}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass FeTiO}_3 = 55.85 + 47.87 + 48.00 \\ = 151.72 \text{ g/mol}$$

$$\text{percent by mass} = \frac{47.87 \text{ g}}{79.87 \text{ g}} \times 100 \\ = 59.93\% \text{ Ti in TiO}_2$$

$$\text{percent by mass} = \frac{47.87 \text{ g}}{151.72 \text{ g}} \times 100 \\ = 31.55\% \text{ Ti in FeTiO}_3$$

$\text{TiO}_2$  has the greater percent by mass of titanium.

- 170. Vitamin E** Many plants contain vitamin E ( $\text{C}_{29}\text{H}_{50}\text{O}_2$ ), a substance that some think slows the aging process in humans. What is the percent composition of vitamin E?

$$29 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 348.29 \text{ g C}$$

$$50 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 50.400 \text{ g H}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass} = 348.29 + 50.400 + 32.00 = 430.69 \\ \text{g/mol}$$

$$\text{percent by mass} = \frac{348.29 \text{ g}}{430.69 \text{ g}} \times 100 = 80.87\% \text{ C}$$

$$\text{percent by mass} = \frac{50.400 \text{ g}}{430.69 \text{ g}} \times 100 = 11.70\% \text{ H}$$

$$\text{percent by mass} = \frac{32.00 \text{ g}}{430.69 \text{ g}} \times 100 = 7.430\% \text{ O}$$

- 171. Artificial Sweetener** Determine the percent composition of aspartame ( $\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5$ ), an artificial sweetener.

$$14 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 168.14 \text{ g C}$$

$$18 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 18.144 \text{ g H}$$

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 28.02 \text{ g N}$$

$$5 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 80.00 \text{ g O}$$

$$\text{molar mass} = 168.14 + 18.144 + 28.02 + 80.00 \\ = 294.30 \text{ g/mol}$$

$$\text{percent by mass} = \frac{168.14 \text{ g}}{294.30 \text{ g}} \times 100 = 57.13\% \text{ C}$$

$$\text{percent by mass} = \frac{18.144 \text{ g}}{294.30 \text{ g}} \times 100 = 6.165\% \text{ H}$$

$$\text{percent by mass} = \frac{28.02 \text{ g}}{294.30 \text{ g}} \times 100 = 9.521\% \text{ N}$$

$$\text{percent by mass} = \frac{80.00 \text{ g}}{294.30 \text{ g}} \times 100 = 27.18\% \text{ O}$$

- 172. MSG** Monosodium glutamate, known as MSG, is sometimes added to food to enhance flavor. Analysis determined this compound to be 35.5% C, 4.77% H, 8.29% N, 13.6% Na, and 37.9% O. What is its empirical formula?

$$35.5 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.96 \text{ mol C}$$

$$4.77 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.73 \text{ mol H}$$

$$8.29 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.592 \text{ mol N}$$

$$13.6 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.592 \text{ mol Na}$$

$$37.9 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.37 \text{ mol O}$$

$$\frac{2.96 \text{ mol C}}{0.0592} = 5.00 \text{ mol C}$$

$$\frac{4.73 \text{ mol H}}{0.0592} = 8.00 \text{ mol H}$$

$$\frac{.0592 \text{ mol N}}{0.0592} = 1.00 \text{ mol N}$$

$$\frac{0.0592 \text{ mol Na}}{0.0592} = 1.00 \text{ mol Na}$$

$$\frac{2.37 \text{ mol O}}{0.0592} = 4.00 \text{ mol O}$$

The simplest whole-number ratio is 5.00 mol C:  
8.00 mol H: 1.00 mol N: 1.00 mol Na: 4.00 mol O.  
The empirical formula is  $\text{C}_5\text{H}_8\text{NO}_4\text{Na}$ .

- 173.** What is the empirical formula of a compound that contains 10.52 g Ni, 4.38 g C, and 5.10 g N?

$$10.52 \text{ g Ni} \times \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} = 0.1792 \text{ mol Ni}$$

$$4.38 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.3640 \text{ mol C}$$

$$5.10 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.3640 \text{ mol N}$$

$$\frac{0.1792 \text{ mol Ni}}{0.1792} = 1.000 \text{ mol Ni}$$

$$\frac{0.3640 \text{ mol C}}{0.1792} = 1.936 \text{ mol C}$$

$$\frac{0.3640 \text{ mol N}}{0.1792} = 2.031 \text{ mol N}$$

The simplest ratio is 1.00 mol Ni: 1.936 mol C:  
2.031 mol N. The empirical formula is  $\text{Ni}(\text{CN})_2$ .

- 174. Patina** The Statue of Liberty has turned green because of the formation of a patina. Two copper compounds,  $\text{Cu}_3(\text{OH})_4\text{SO}_4$  and  $\text{Cu}_4(\text{OH})_6\text{SO}_4$ , form this patina. Find the mass percentage of copper in each compound.



$$3 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 190.65 \text{ g Cu}$$

$$8 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 128.00 \text{ g O}$$

$$4 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 4.032 \text{ g H}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$\text{molar mass } \text{Cu}_3(\text{OH})_4\text{SO}_4 = 190.65 + 128.00 + 4.032 + 32.07 = 354.75 \text{ g/mol}$$

$$\text{percent by mass} = \frac{190.65 \text{ g}}{354.75} + 100 = 53.74\%$$



$$4 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 254.20 \text{ g Cu}$$

$$10 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 160.00 \text{ g O}$$

$$6 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 6.048 \text{ g H}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$\text{molar mass } \text{Cu}_4(\text{OH})_6\text{SO}_4 = 254.20 + 160.00 + 6.048 + 32.07 = 452.32 \text{ g/mol}$$

$$\text{percent by mass} = \frac{254.20 \text{ g}}{452.32 \text{ g}} \times 100 = 56.20\%$$

## Section 10.5

### Mastering Concepts

- 175.** What is a hydrated compound? Use an example to illustrate your answer.

A hydrated compound is a compound that has a specific number of water molecules associated with its atoms, for example,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  and  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .

- 176.** Explain how hydrates are named.

First, name the compound. Then, add a prefix (mono- di- tri-) that indicates how many water molecules are associated with one mole of compound.

- 177. Desiccants** Why are certain electronic devices transported with desiccants?

Desiccants are anhydrous forms of a hydrate that absorb water from the air and keep it off of electronic devices.

- 178.** In a laboratory setting, how would you determine if a compound was a hydrate?

Mass a sample of the compound before heating. Heat the compound, allow it to cool, and remass the sample. A change in mass might indicate that the compound is a hydrate.

- 179.** Write the formula for the following hydrates.

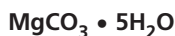
- a. nickel(II) chloride hexahydrate



b. cobalt(II) chloride hexahydrate



c. magnesium carbonate pentahydrate



d. sodium sulfate decahydrate



### Mastering Problems

**180.** Determine the mass percent of anhydrous sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) and water in sodium carbonate decahydrate ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ).

$$2 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 45.98 \text{ g Na}$$

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$13 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 208.00 \text{ g O}$$

$$20 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 20.16 \text{ g H}$$

$$\text{molar mass Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = 45.98 + 12.01 + 208.00 + 20.16 = 286.15 \text{ g/mol}$$

$$2 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 45.98 \text{ g Na}$$

$$1 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.01 \text{ g C}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass Na}_2\text{CO}_3 = 45.98 + 12.01 + 48.00 = 105.99 \text{ g/mol}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass H}_2\text{O} = 16.00 + 2.016 = 18.02 \text{ g/mol}$$

$$\frac{105.99 \text{ g}}{286.15 \text{ g}} \times 100 = 37.03\% \text{ Na}_2\text{CO}_3$$

$$\frac{10(18.02 \text{ g})}{286.15 \text{ g}} \times 100 = 62.97\% \text{ H}_2\text{O}$$

**181.** Table 10.4 shows data from an experiment to determine the formulas of hydrated barium chloride. Determine the formula for the hydrate and its name.

| Data for $\text{BaCl}_2 \cdot x\text{H}_2\text{O}$ |          |
|--|----------|
| Mass of empty crucible                             | 21.30 g  |
| Mass of hydrate + crucible                         | 31.35 mL |
| Initial mass of hydrate                            |          |
| Mass after heating 5 min                           | 29.87 g  |
| Mass of anhydrous solid                            |          |

$$\text{initial mass of hydrate} = (\text{mass of hydrate} + \text{crucible}) - (\text{mass of empty crucible}) = 31.35 \text{ g} - 21.30 \text{ g} = 10.05 \text{ g}$$

$$\text{mass of anhydrous solid} = (\text{mass after heating 5 min}) - (\text{mass of empty crucible}) = 29.87 \text{ g} - 21.30 \text{ g} = 8.57 \text{ g}$$

$$\text{mass of water} = (\text{initial mass of hydrate}) - (\text{mass of anhydrous solid}) = 10.05 \text{ g} - 8.57 \text{ g} = 1.48 \text{ g}$$

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 137.33 \text{ g Ba}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass BaCl}_2 = 137.33 + 70.90 = 208.23 \text{ g/mol}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass H}_2\text{O} = 16.00 + 2.016 = 18.02 \text{ g/mol}$$

$$8.57 \text{ g BaCl}_2 \times \frac{1 \text{ mol BaCl}_2}{208.23 \text{ g BaCl}_2} = 0.0412 \text{ mol BaCl}_2$$

$$1.48 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.0821 \text{ mol H}_2\text{O}$$

$$n = \frac{0.0821 \text{ mol}}{0.0412 \text{ mol}} = 2.00$$

$\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ , barium chloride dihydrate

- 182.** Chromium(III) nitrate forms a hydrate that is 40.50% water by mass. What is its chemical formula?

$$100.0\% - 40.5\% = 59.5\% \text{ Cr(NO}_3)_3$$

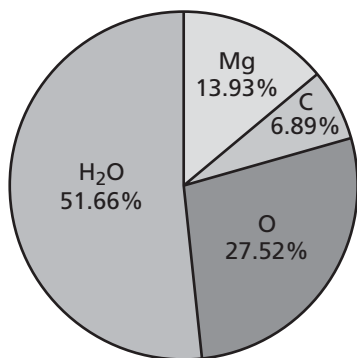
$$\frac{59.5 \text{ g Cr(NO}_3)_3}{238 \text{ g/mol}} = \frac{0.250 \text{ mol Cr(NO}_3)_3}{0.250 \text{ mol}}$$

$$= 1 \text{ mol Cr(NO}_3)_3$$

$$\frac{40.80 \text{ g H}_2\text{O}}{18.02 \text{ g/mol}} = \frac{2.26 \text{ mol H}_2\text{O}}{0.250 \text{ mol}} = 9 \text{ mol H}_2\text{O}$$

$$\text{formula} = \text{Cr(NO}_3)_3 \cdot 9 \text{ H}_2\text{O}$$

- 183.** Determine the percent composition of  $\text{MgCO}_3 \cdot 5\text{H}_2\text{O}$  and draw a pie graph to represent the hydrate.



Circle graph should show the following percentages (if addition of molar mass takes place to one-hundredths spot): 13.93% Mg, 6.89% C, 27.52% O, 51.66% H<sub>2</sub>O

$$\text{Molar mass} = 174.41 \text{ g/mol}$$

$$1 \text{ mol Mg} \times \frac{24.30 \text{ g/mol}}{174.41 \text{ g/mol}} \times 100 = 13.93\% \text{ Mg}$$

$$1 \text{ mol C} \times \frac{12.01 \text{ g/mol}}{174.41 \text{ g/mol}} \times 100 = 6.89\% \text{ C}$$

$$3 \text{ mol O} \times \frac{16.00 \text{ g/mol}}{174.41 \text{ g/mol}} \times 100 = 27.52\% \text{ O}$$

$$5 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g/mol}}{174.41 \text{ g/mol}} \times 100 = 51.66\% \text{ H}_2\text{O}$$

- 184.** What is the formula and name of a hydrate that is 85.3% barium chloride and 14.7% water?

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 137.33 \text{ g Ba}$$

$$2 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 70.90 \text{ g Cl}$$

$$\text{molar mass BaCl}_2 = 137.33 + 70.90 = 208.23 \text{ g/mol}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass H}_2\text{O} = 16.00 + 2.016 = 18.02 \text{ g/mol}$$

$$85.3 \text{ g BaCl}_2 \times \frac{1 \text{ mol BaCl}_2}{208.23 \text{ g BaCl}_2} = 0.410 \text{ mol BaCl}_2$$

$$14.7 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.816 \text{ mol H}_2\text{O}$$

$$n = \frac{0.816 \text{ mol}}{0.410 \text{ mol}} = 2.00$$

$$\text{BaCl}_2 \cdot 2\text{H}_2\text{O}, \text{ barium chloride dihydrate}$$

- 185.** Gypsum is hydrated calcium sulfate. A 4.89-g sample of this hydrate was heated. After the water was removed, 3.87 g anhydrous calcium sulfate remained. Determine the formula for this hydrate and name the compound.

$$\text{mass of hydrated CaSO}_4 = \text{mass of CaSO}_4 + \text{mass of H}_2\text{O}$$

$$\text{mass of H}_2\text{O} = \text{mass of hydrated CaSO}_4 - \text{mass of H}_2\text{O} = 4.89 \text{ g} - 3.87 \text{ g} = 1.02 \text{ g}$$

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g Ca}$$

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$4 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g O}$$

$$\text{molar mass CaSO}_4 = 40.08 + 32.07 + 64.00 = 136.15 \text{ g/mol}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass H}_2\text{O} = 16.00 + 2.016 = 18.02 \text{ g/mol}$$

$$3.87 \text{ g CaSO}_4 \times \frac{1 \text{ mol CaSO}_4}{134.15 \text{ g CaSO}_4} = 0.0284 \text{ mol CaSO}_4$$

$$1.02 \text{ g H}_2\text{O} + \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.0566 \text{ mol H}_2\text{O}$$

$$n = \frac{0.0566 \text{ mol}}{0.0284 \text{ mol}} = 2.00$$

$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ , calcium sulfate dihydrate

- 186.** A 1.628-g sample of a hydrate of magnesium iodide is heated until its mass is reduced to 1.072 g and all water has been removed. What is the formula of the hydrate?

mass of water = mass of hydrate – mass of anhydrous solid = 1.628 g – 1.072 g = 0.556 g

$$1 \text{ mol Mg} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 24.31 \text{ g Mg}$$

$$2 \text{ mol I} \times \frac{126.90 \text{ g I}}{1 \text{ mol I}} = 253.80 \text{ g I}$$

$$\text{molar mass MgI}_2 = 24.31 + 253.80 = 278.11 \text{ g/mol}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass H}_2\text{O} = 16.00 + 2.016 = 18.02 \text{ g/mol}$$

$$1.072 \text{ g MgI}_2 \times \frac{1 \text{ mol MgI}_2}{278.11 \text{ g MgI}_2} = 0.003855 \text{ mol MgI}_2$$

$$0.556 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.0309 \text{ mol H}_2\text{O}$$

$$n = \frac{0.0309 \text{ mol}}{0.003855 \text{ mol}} = 8.02$$

$\text{MgI}_2 \cdot 8\text{H}_2\text{O}$

- 187. Borax** Hydrated sodium tetraborate ( $\text{Na}_2\text{B}_4\text{O}_7 \cdot x\text{H}_2\text{O}$ ) is commonly called borax. Chemical analysis indicates that this hydrate is 52.8% sodium tetraborate and 47.2% water. Determine the formula and name the hydrate.

$$2 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 45.98 \text{ g Na}$$

$$4 \text{ mol B} \times \frac{12.01 \text{ g B}}{1 \text{ mol B}} = 48.04 \text{ g B}$$

$$7 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 112.00 \text{ g O}$$

$$\text{molar mass Na}_2\text{B}_4\text{O}_7 = 45.98 + 48.04 + 112.00 = 206.02 \text{ g/mol}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 16.00 \text{ g O}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g H}$$

$$\text{molar mass H}_2\text{O} = 16.00 + 2.016 = 18.02 \text{ g/mol}$$

$$52.8 \text{ g Na}_2\text{B}_4\text{O}_7 \times \frac{1 \text{ mol Na}_2\text{B}_4\text{O}_7}{206.02 \text{ g Na}_2\text{B}_4\text{O}_7} = 0.256 \text{ mol Na}_2\text{B}_4\text{O}_7$$

$$47.2 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 2.61 \text{ mol H}_2\text{O}$$

$$n = \frac{2.61 \text{ mol}}{0.256 \text{ mol}} = 10.2$$

$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$ , sodium tetraborate decahydrate

### Mixed Review

- 188.** Rank samples A–D from least number of atoms to greatest number of atoms. A: 1.0 mol of  $\text{H}_2$ , B: 0.75 mol of  $\text{H}_2\text{O}$ , C: 1.5 mol of  $\text{NaCl}$ , D: 0.50 mol of  $\text{Ag}_2\text{S}$

A:  $\text{H}_2$  has two atoms per mole, thus  $2(1.0 \text{ mol}) = 2 \text{ mol atoms}$

$$2.0 \text{ mol} \times \frac{6.02 + 10^{23} \text{ atoms}}{1 \text{ mol H}} = 1.20 + 10^{24} \text{ atoms}$$

B:  $\text{H}_2\text{O}$  has three atoms per mole, thus  $3(0.75 \text{ mol}) = 2.25 \text{ mol atoms}$

$$2.25 \text{ mol} \times \frac{6.02 + 10^{23} \text{ atoms}}{1 \text{ mol}} = 1.35 + 10^{24} \text{ atoms}$$

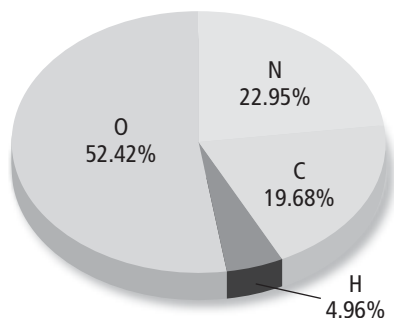
C:  $\text{NaCl}$  has two atoms per mole, thus  $(1.5 \text{ mol}) = 3.0 \text{ mol atoms}$

$$3.0 \text{ mol} \times \frac{6.02 + 10^{23} \text{ atoms}}{1 \text{ mol}} = 1.81 + 10^{24} \text{ atoms}$$

D:  $\text{Ag}_2\text{S}$  has three atoms per mole, thus  $3(0.50 \text{ mol}) = 1.5 \text{ mol atoms}$

$$1.5 \text{ mol} \times \frac{6.02 + 10^{23} \text{ atoms}}{1 \text{ mol}} = 9.03 + 10^{23} \text{ atoms}$$

$$D < A < B < C$$



- 189.** The graph in **Figure 10.23** shows the percent composition of a compound containing carbon, hydrogen, oxygen, and nitrogen. How many grams of each element are present in 100 g of the compound?

Because the percentages are 100-based values, and the masses per 100-g are desired, simply change each percent to grams.

52.42 g O, 22.95 g N, 19.68 g C, 4.96 g H

- 190.** How many grams of  $\text{CoCl}_2 \cdot 6 \text{H}_2\text{O}$  must you measure out in a container to have exactly Avogadro's number of particles?

Avogadro's number is the number of particles in one mole. You must measure 237.8 g of  $\text{CoCl}_2 \cdot 6 \text{H}_2\text{O}$  to have one mole of particles of this compound.

- 191.** One atom of an unknown element has a mass of  $6.66 \times 10^{-23}$  g. What is the identity of this element?

$$\frac{6.66 + 10^{-23} \text{ g}}{1 \text{ atom}} \times \frac{6.02 + 10^{23} \text{ atoms}}{1 \text{ mol}} = 40.09 \text{ g/mol}$$

This is closest to the molar mass of calcium.

- 192. Skunks** Analysis of skunk spray yields a molecule with 44.77% C, 7.46% H and 47.76% S. What is the chemical formula for this molecule found in the spray from skunks that scientists think is partly responsible for the foul odor?

$$\frac{44.77 \text{ g C}}{12.01 \text{ g/mol}} = \frac{3.73 \text{ mol C}}{1.49 \text{ mol}} = 2.5 \text{ mol C}$$

$$\frac{7.46 \text{ g H}}{1.008 \text{ g/mol}} = \frac{7.40 \text{ mol H}}{1.49 \text{ mol}} = 5 \text{ mol H}$$

$$\frac{47.76 \text{ g S}}{32.07 \text{ g/mol}} = \frac{1.49 \text{ mol S}}{1.49 \text{ mol}} = 1 \text{ mol S}$$

$$\frac{3.73 \text{ mol C}}{1.49 \text{ mol}} : \frac{7.40 \text{ mol H}}{1.49 \text{ mol}} : \frac{1.49 \text{ mol S}}{1.49 \text{ mol}}$$

$$(2.5 \text{ mol C}:5.0 \text{ mol H}:1 \text{ mol S})_2 = 5.0 \text{ mol C}:10.0 \text{ mol H}:2.0 \text{ mol S} = \text{C}_5\text{H}_{10}\text{S}_2$$

- 193.** How many moles are present in 1.00 g of each compound?

- a. L-tryptophan ( $\text{C}_{11}\text{H}_{12}\text{N}_2\text{O}_2$ ), an essential amino acid

$$\text{C}_{11}\text{H}_{12}\text{N}_2\text{O}_2: \text{molar mass} = 204 \text{ g/mol}$$

$$1.00 \text{ g}/204 \text{ g/mol} = 0.00490 \text{ mol C}_{11}\text{H}_{12}\text{N}_2\text{O}_2$$

- b. magnesium sulfate heptahydrate, also known as Epsom salts

$$\text{MgSO}_4 \cdot 7 \text{H}_2\text{O}: \text{molar mass} = 246.4 \text{ g/mol}$$

$$1.00 \text{ g}/246.4 \text{ g/mol} =$$

$$0.00406 \text{ mol MgSO}_4 \cdot 7 \text{H}_2\text{O}$$

- c. propane ( $\text{C}_5\text{H}_8$ ), a fuel

$$\text{C}_5\text{H}_8: \text{molar mass} = 68.1 \text{ g/mol}$$

$$1.00 \text{ g}/68.1 \text{ g/mol} = 0.0147 \text{ mol C}_5\text{H}_8$$

- 194.** A compound contains 6.0 g of carbon and 1.0 g of hydrogen, and has a molar mass of 42.0 g/mol. What are the compound's percent composition, empirical formula, and molecular formula?

$$6.0 \text{ g C}/7.0 \text{ g} = 85.7 \% \text{ C}$$

$$1.0 \text{ g H}/7.0 \text{ g} = 14.3 \% \text{ H}$$

$$(6.0 \text{ g C})/(12.01 \text{ g/mol}) = 0.50 \text{ mol C}$$

$$(1.0 \text{ g H})/(1.008 \text{ g/mol}) = 1.0 \text{ mol H}$$

$$0.50 \text{ mol C}/0.50 \text{ mol} = 1 \text{ mol C}$$

$$1.0 \text{ mol H}/0.50 \text{ mol} = 2 \text{ mol H}$$

$$\text{empirical formula: } 1 \text{ mol C}:2 \text{ mol H} = \text{CH}_2$$

$$\text{empirical mass} = 14.0 \text{ g/mol}$$

$$\text{empirical mass/molar mass} = (42.0 \text{ g/mol})/(14.0 \text{ g/mol}) = 3$$

$$\text{molecular formula} = 3(\text{CH}_2) = \text{C}_3\text{H}_6$$

195. Which of these compounds has the greatest percent of oxygen by mass:  $\text{TiO}_2$ ,  $\text{Fe}_2\text{O}_3$ , or  $\text{Al}_2\text{O}_3$ ?

$$1 \text{ mol Ti} + \frac{47.87 \text{ g Ti}}{1 \text{ mol Ti}} = 47.87 \text{ g Ti}$$

$$2 \text{ mol O} + \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

$$\text{molar mass TiO}_2 = 47.87 + 32.00 = 79.87 \text{ g/mol}$$

$$2 \text{ mol Fe} + \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 111.70 \text{ g Fe}$$

$$3 \text{ mol O} + \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass Fe}_2\text{O}_3 = 111.70 + 48.00 = 159.70 \text{ g/mol}$$

$$2 \text{ mol Al} + \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 53.96 \text{ g Al}$$

$$3 \text{ mol O} + \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 48.00 \text{ g O}$$

$$\text{molar mass Al}_2\text{O}_3 = 53.96 + 48.00 = 54.44 \text{ g/mol}$$

$$\text{percent by mass} = \frac{32.00 \text{ g}}{79.87 \text{ g}} + 100 = 40.07\% \text{ O in TiO}_2$$

$$\text{percent by mass} = \frac{48.00 \text{ g}}{159.70 \text{ g}} + 100 = 30.06\% \text{ O in Fe}_2\text{O}_3$$

$$\text{percent by mass} = \frac{48.00 \text{ g}}{54.44 \text{ g}} + 100 = 87.99\% \text{ O in Al}_2\text{O}_3$$

$\text{Al}_2\text{O}_3$  has the greatest percent of oxygen by mass.

196. **Mothballs** Naphthalene, commonly found in mothballs, is composed of 93.7% carbon and 6.3% hydrogen. The molar mass of naphthalene is 128 g/mol. Determine the empirical and molecular formulas for naphthalene.

$$93.7 \text{ g C} + \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.80 \text{ mol C}$$

$$6.3 \text{ g H} + \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.2 \text{ mol H}$$

$$\frac{7.80 \text{ mol C}}{6.2} = 1.25 \text{ mol C}$$

$$\frac{6.2 \text{ mol H}}{6.2} = 1.0 \text{ mol H}$$

The simplest ratio is 1.25 mol C : 1.0 mol H. Multiply both subscripts by four to obtain the simplest whole-number ratio. The empirical formula is  $\text{C}_5\text{H}_4$ .

$$5 \text{ mol C} + \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 60.05 \text{ g C}$$

$$4 \text{ mol H} + \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 4.032 \text{ g H}$$

$$\text{molar mass} = 60.05 + 4.032 = 64.09 \text{ g/mol}$$

$$n = \frac{128 \text{ g/mol}}{64.09 \text{ g/mol}} = 2.00$$

Molecular formula is  $\text{C}_{10}\text{H}_8$ .

197. Which of these molecular formulas are also empirical formulas: ethyl ether ( $\text{C}_4\text{H}_{10}\text{O}$ ), aspirin ( $\text{C}_9\text{H}_8\text{O}_4$ ), butyl dichloride ( $\text{C}_4\text{H}_8\text{Cl}_2$ ), glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ )?

$\text{C}_4\text{H}_{10}\text{O}$  and  $\text{C}_9\text{H}_8\text{O}_4$ ; The subscripts in  $\text{C}_4\text{H}_{10}\text{O}$  and  $\text{C}_9\text{H}_8\text{O}_4$  represent simplest whole-number ratios.

## Think Critically

198. **Apply Concepts** A mining company has two possible sources of copper: chalcopyrite ( $\text{CuFeS}_2$ ) and chalcocite ( $\text{Cu}_2\text{S}$ ). If the mining conditions and the extraction of copper from the ore were identical for each of the ores, which ore would yield the greater quantity of copper? Explain your answer.



$$1 \text{ Cu} + 63.55 \text{ g/mol} = 63.55 \text{ g}$$

$$1 \text{ Fe} + 55.85 \text{ g/mol} = 55.85$$

$$2 \text{ S} + 32.07 \text{ g/mol} = 64.14 \text{ g}$$

$$\text{molar mass} = 183.54 \text{ g/mol}$$

$$\% \text{Cu} = (63.55 \text{ g}/183.54 \text{ g/mol}) + 100 = 34.62\% \text{ Cu}$$



$$2 \text{ Cu} + 63.55 \text{ g/mol} = 127.10 \text{ g}$$

$$1 \text{ S} + 32.07 \text{ g/mol} = 32.07 \text{ g}$$

$$\text{molar mass} = 159.17 \text{ g/mol}$$

$$\% \text{Cu} = (127.10 \text{ g}/159.17 \text{ g/mol}) + 100 = 79.9\% \text{ Cu}$$

Chalcopyrite ( $\text{CuFeS}_2$ ) is 34.6% copper by mass (determined from percent composition) and chalcocite ( $\text{Cu}_2\text{S}$ ) is 79.9% copper by mass. Chalcocite would yield the greater quantity of copper because the ore has the greater percentage copper by mass.

- 199. Analyze and Conclude** On a field trip, students collected rock samples. Analysis of the rocks revealed that two of the rock samples contained lead and sulfur. **Table 10.5** shows the percent lead and sulfur in each of the rocks. Determine the molecular formula of each rock. What can the students conclude about the rock samples?

Lead and Sulfur Content

| Rock Sample | % Lead | % Sulfur |
|-------------|--------|----------|
| 1           | 86.6 % | 13.4%    |
| 2           | 76.4%  | 23.6%    |

**Rock Sample 1**

$$86.6\% \text{ g Pb} = 86.6 \text{ g Pb}; 86.6 \text{ g}/207.2 \text{ g/mol} = 0.418 \text{ mol Pb}$$

$$13.4\% \text{ g S} = 13.4 \text{ g}; 13.4 \text{ g}/32.07 \text{ g/mol} = 0.418 \text{ mol S}$$

Dividing each quantity by 0.418 mol yields:  
 $1 \text{ mol Pb} + 1 \text{ mol S} = \text{PbS}$

**Rock Sample 2**

$$76.4\% \text{ g Pb} = 76.4 \text{ g Pb}; 76.4 \text{ g}/207.2 \text{ g/mol} = 0.369 \text{ mol Pb}$$

$$23.6\% \text{ g S} = 23.6 \text{ g S}; 23.6 \text{ g}/32.07 \text{ g/mol} = 0.736 \text{ mol S}$$

Dividing each quantity by 0.369 mol yields:



The rocks contain two different compounds of lead and sulfur. Further analysis will show that Sample 1 has the formula  $\text{PbS}$  and that Sample 2 has the formula  $\text{PbS}_2$ .

- 200. Graph** A YAG, or yttrium aluminum garnet ( $\text{Y}_3\text{Al}_5\text{O}_{12}$ ), is a synthetic gemstone which has no counterpart in nature. Design a bar graph to indicate the moles of each element present in a 5.67 carat yttrium aluminum garnet. (1 carat = 0.20 g)

$$(5.67 \text{ carat } \text{Y}_3\text{Al}_5\text{O}_{12})(0.20 \text{ g/carat}) = 1.13 \text{ g } \text{Y}_3\text{Al}_5\text{O}_{12}$$

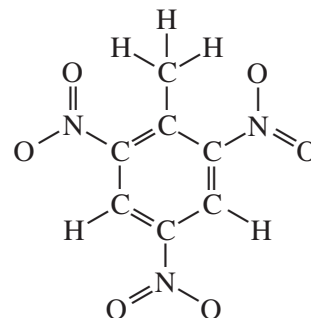
$$(1.13 \text{ g } \text{Y}_3\text{Al}_5\text{O}_{12})(1 \text{ mol}/436.75 \text{ g}) = 0.00260 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12}$$

$$(0.00260 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12})(3 \text{ mol Y}/1 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12}) = 0.00779 \text{ mol Y}$$

$$(0.00260 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12})(5 \text{ mol Al}/1 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12}) = 0.0130 \text{ mol Al}$$

$$(0.00260 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12})(12 \text{ mol O}/1 \text{ mol } \text{Y}_3\text{Al}_5\text{O}_{12}) = 0.0312 \text{ mol O}$$

The bar graph should show 0.00779 mol Y, 0.0130 mol Al, and 0.0312 mol O.



- 201. Assess** The structure of the TNT molecule is shown in **Figure 10.24**. Critique the statement “Trinitrotoluene, TNT, contains 21 atoms per mole.” What is correct about the statement and what is incorrect? Rewrite the statement.

A molecule of TNT contains 21 atoms but a mole of TNT contains  $6.02 \times 10^{23}$  molecules. The statement is false as stated. “Trinitrotoluene, TNT, has the formula  $\text{C}_7\text{H}_5\text{N}_3\text{O}_6$  which contains 21 atoms but a mole of TNT contains  $6.02 \times 10^{23}$  molecules.”



- 202. Design an Experiment** Design an experiment that can be used to determine the amount of water in alum ( $\text{KAl}(\text{SO}_4)_2 \cdot x\text{H}_2\text{O}$ ).

Determine and record the mass of an empty evaporating dish. Add about 2 g of the hydrate. Measure and record the mass. Heat the evaporating dish gently for 5 minutes, and strongly for another 5 minutes to evaporate all the water. Allow the dish to cool, and measure and record the mass. Determine the masses of the anhydrous solid and the water lost. Calculate the number of moles of anhydrous compound and water. Determine the ratio of moles of water to moles of anhydrous compound. Use the whole-number ratio of the moles as the coefficient of  $\text{H}_2\text{O}$  in the formula.

- 203.** Design a concept map that illustrates the mole concept. Include the terms *moles*, *Avogadro's number*, *molar mass*, *number of particles*, *percent composition*, *empirical formula*, and *molecular formula*.

Concept maps will vary but should show logical connection between Avogadro's number of particles, a mole, and molar mass, and also show how the molecular formula is a whole-number multiple of the empirical formula.

### Challenge Problem

- 204.** Two different compounds are composed of Elements X and Y. The formulas of the compounds are  $\text{X}_2\text{Y}_3$  and  $\text{XY}$ . A 0.25 mol sample of  $\text{XY}$  has a mass of 17.96 g, and a 0.25 mol sample of  $\text{X}_2\text{Y}_3$  has a mass of 39.92 g.

- What are the atomic masses of elements X and Y?
- What are the formulas for the compounds?

$$\text{Compound XY: } \frac{17.96 \text{ g}}{0.25 \text{ mol}} = 71.84 \text{ g/mol}$$

$$71.84 \text{ g/mol} = \text{X} + \text{Y}, \text{Y} = 71.84 \text{ g/mol} - \text{X}$$

$$\text{Compound } \text{X}_2\text{Y}_3: 39.92 \text{ g}/0.25 \text{ mol} = 159.68 \text{ g/mol}$$

$$159.68 \text{ g/mol} = 2\text{X} + 3\text{Y}$$

$$159.68 \text{ g/mol} = 2\text{X} + 3(71.84 \text{ g/mol} - \text{X})$$

[substitute for Y and solve for X]

$$159.68 \text{ g/mol} = 2\text{X} + 215.52 \text{ g/mol} - 3\text{X}$$

$$-55.85 \text{ g/mol} = -\text{X}$$

$$\text{X} = 55.84 \text{ g/mol}$$

$$\text{X} + \text{Y} = 71.84 \text{ g/mol}$$

$$55.84 \text{ g/mol} + \text{Y} = 71.84 \text{ g/mol}$$

$$\text{Y} = 16 \text{ g/mol}$$

X is the element iron (Fe); Y is the element oxygen (O)

Formulas are FeO and  $\text{Fe}_2\text{O}_3$ .

### Cumulative Review

- 205.** Express each answer with the correct number of significant figures. (*Chapter 2*)

a.  $18.23 - 456.7$

$$-438.5$$

b.  $4.233 \div 0.0131$

$$323$$

c.  $(82.44 + 4.92) + 0.125$

$$406$$

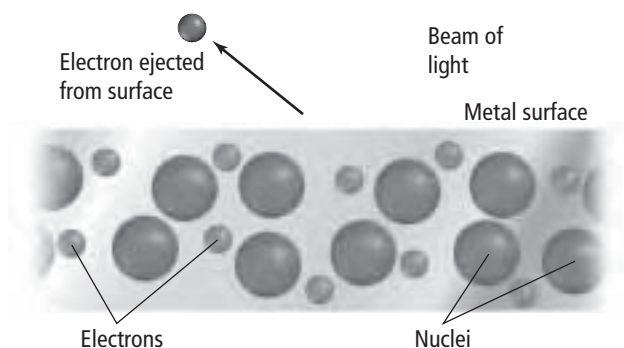
- 206. Making Candy** A recipe for pralines calls for the candy mixture to be heated until it reaches the "soft ball" stage, about  $236^\circ\text{F}$ . Can a Celsius thermometer with a range of  $-10$  to  $110^\circ\text{C}$  be used to determine when the "soft ball" stage is reached? (*Chapter 2*)

$$^\circ\text{C} = \frac{^\circ\text{F} - 32}{1.8} = \frac{204}{1.8} = 113^\circ\text{C}.$$

No, the Celsius thermometer could not be used because  $113^\circ\text{C}$  is beyond the range of the thermometer.

- 207.** Contrast atomic number and mass number. Compare these numbers for isotopes of an element. (*Chapter 4*)

Atomic number equals the number of protons. Mass number equals the number of protons plus the number of neutrons. Two isotopes of an element will have the same atomic number but different mass numbers.



- 208.** Describe the phenomenon in Figure 10.25. Explain why the electrons are not bound to the nuclei. (*Chapter 5*)

The illustration shows the photoelectric effect; the emission of electrons from a metal surface when light of sufficient energy strikes the metal. The electrons in metals are not bound to nuclei; they are free to move.

- 209.** Given the elements Ar, Cs, Br, and Ra, identify those that form positive ions. Explain your answer. (*Chapter 7*)

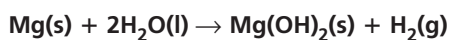
Cs and Ra can form positive ions. The configuration of Cs is  $[\text{Xe}]6s^1$ ; Cs loses 1 valence electron to form a  $1+$  ion. The configuration of Ra is  $[\text{Rn}]7s^2$ ; Ra loses 2 electrons to form a  $2+$  ion. Argon is a noble gas and does not lose or gain electrons. Br gains one electron to form a  $1-$  ion.

- 210.** Write the formula and name the compound formed when each pair of elements combine. (*Chapter 7*)

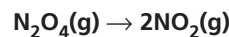
- barium and chlorine  
 $\text{BaCl}_2$ ; barium chloride
- aluminum and selenium  
 $\text{Al}_2\text{Se}_3$ ; aluminum selenide
- calcium and phosphorus  
 $\text{Ca}_3\text{P}_2$ ; calcium phosphide

- 211.** Write balanced equations for each reaction. (*Chapter 9*)

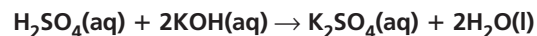
- Magnesium metal and water combine to form solid magnesium hydroxide and hydrogen gas.



- Dinitrogen tetroxide gas decomposes into nitrogen dioxide gas.



- Aqueous solutions of sulfuric acid and potassium hydroxide undergo a double-replacement reaction.



## Additional Assessment

### Writing in Chemistry

- 212. Natural Gas** Natural gas hydrates are chemical compounds known as clathrate hydrates. Research natural gas hydrates and prepare an educational pamphlet for consumers. The pamphlet should discuss the composition and structure of the compounds, the location of the hydrates, their importance to consumers, and the environmental impact of using the hydrates.

Natural gas hydrates are crystalline solids resembling snow that consist of a water-ice lattice within which light hydrocarbon molecules such as methane, ethane, and propane are trapped in the space between water molecules. They naturally form at moderately high pressures and temperatures near the freezing point of water. These conditions exist in the Arctic permafrost regions of North America and Eurasia and at depth along the world's continental slopes and rises on the sea floor. Natural gas hydrate can be thought of as a "methane concentrator". Gas hydrates may be a new clean energy source. Huge amounts of natural gas are tied up in natural gas hydrates globally. If exploited as a energy source, removal of gas hydrates may destabilize the sea floor, triggering submarine landslides and huge releases of methane. Methane is a very effective greenhouse gas, and large methane releases may explain sudden episodes of climatic warming in the geologic past.

**213. Avogadro** Research and report on the life of the Italian chemist Amedeo Avogadro (1776–1856) and how his work led scientists to determine the number of particles in a mole.

Students should mention Avogadro's hypothesis. Avogadro formulated his hypothesis as an explanation for earlier works by Gay-Lussac and Ritter. His ideas were rejected by chemists of his day but were revived later by the Italian chemist Stanislao Cannizzaro. Avogadro died before seeing his ideas accepted.

**214. Luminol** Crime-scene investigators use luminol to visualize blood residue. Research luminol and determine its chemical formula and percent composition.

Luminol( $C_8H_7O_2N_3$ ) reacts with compounds in blood and releases energy in the form of light. When crime scene investigators see the light, they know blood is at the scene. The percent composition of  $C_8H_7O_2N_3$  is 54.23% C, 3.98% H, 18.06% O, and 23.72% N.

$$8 \text{ mol C} \times 12.01 \text{ g/mol} = 96.08 \text{ g C}$$

$$7 \text{ mol H} \times 1.008 \text{ g/mol} = 7.056 \text{ g H}$$

$$2 \text{ mol O} \times 16.00 \text{ g/mol} = 32.00 \text{ g O}$$

$$3 \text{ mol N} \times 14.01 \text{ g/mol} = 42.03 \text{ g N}$$

$$\text{molar mass} = 177.17 \text{ g/mol}$$

$$\% \text{ C} = (96.08 \text{ g C} / 177.17 \text{ g/mol}) \times 100 = 54.23\% \text{ C}$$

$$\% \text{ H} = (7.056 \text{ g H} / 177.17 \text{ g/mol}) \times 100 = 3.98\% \text{ H}$$

$$\% \text{ O} = (32.00 \text{ g O} / 177.17 \text{ g/mol}) \times 100 = 18.06\% \text{ O}$$

$$\% \text{ N} = (42.03 \text{ g} / 177.17 \text{ g/mol}) \times 100 = 23.72\% \text{ N}$$

## Document-Based Questions

**Space Shuttle Propellants** At liftoff, the orbiter and an external fuel tank carry 3,164,445 L of the liquid propellants hydrogen, oxygen, hydrazine, monomethylhydrazine and dinitrogen tetroxide. Their total mass is 727,233 kg. Data for the propellants carried at liftoff are given in Table 10.6.

Data obtained from: "Space Shuttle Use of Propellants and Fluids." September 2001. *NASA Fact Sheet*

| Space Shuttle Liquid Propellants |                                   |                        |                        |                         |
|----------------------------------|-----------------------------------|------------------------|------------------------|-------------------------|
| Propellants                      | Molecular Formula                 | Mass (kgs)             | Mole                   | Molecules               |
| Hydrogen                         | H <sub>2</sub>                    | 1.04 + 10 <sup>5</sup> | 5.14 + 10 <sup>7</sup> | 3.09 + 10 <sup>31</sup> |
| Oxygen                           | O <sub>2</sub>                    | 6.18 + 10 <sup>5</sup> | 1.93 + 10 <sup>7</sup> | 1.16 + 10 <sup>31</sup> |
| Hydrazine                        | N <sub>2</sub> H <sub>4</sub>     | 493                    | 1.54 + 10 <sup>4</sup> | 9.27 + 10 <sup>27</sup> |
| Monomethylhydrazine              | CH <sub>3</sub> NHNH <sub>2</sub> | 4,909                  | 1.07 + 10 <sup>5</sup> | 6.44 + 10 <sup>28</sup> |
| Dinitrogen tetroxide             | N <sub>2</sub> O <sub>4</sub>     | 7950                   | 8.64 + 10 <sup>4</sup> | 5.20 + 10 <sup>28</sup> |

**215.** Hydrazine contains 87.45% nitrogen and 12.55% hydrogen, and has a molar mass of 32.04 g/mol. Determine hydrazine's molecular formula. Record the molecular formula in Table 10.6.

$$87.45\% \text{ g N} = 87.45 \text{ g}; 87.45 \text{ g} / 14.01 \text{ g/mol} = 6.24 \text{ mol N}$$

$$12.55\% \text{ g H} = 12.55 \text{ g H}; 12.55 \text{ g} / 1.008 \text{ g/mol} = 12.45 \text{ mol H}$$

Dividing each quantity by 6.24 mol yields: 1 mol N : 2 mol H

$$\text{Empirical formula} = \text{NH}_2$$

$$\text{Empirical mass} = 14.007 \text{ g} + 2(1.008 \text{ g}) = 16.023 \text{ g/mol}$$

$$(32.04 \text{ g/mol}) / (16.023 \text{ g/mol}) = 2$$

$$\text{molecular formula} = 2(\text{NH}_2) = \text{N}_2\text{H}_4$$

216. Complete Table 10.6 by calculating the number of moles, mass in kilograms, or molecules for each propellant. Give all answers to three significant figures.

| Space Shuttle Liquid Propellants |                                   |                        |                        |                         |
|----------------------------------|-----------------------------------|------------------------|------------------------|-------------------------|
| Propellants                      | Molecular Formula                 | Mass (kgs)             | Mole                   | Molecules               |
| Hydrogen                         | H <sub>2</sub>                    | 1.04 + 10 <sup>5</sup> | 5.14 + 10 <sup>7</sup> | 3.09 + 10 <sup>31</sup> |
| Oxygen                           | O <sub>2</sub>                    | 6.18 + 10 <sup>5</sup> | 1.93 + 10 <sup>7</sup> | 1.16 + 10 <sup>31</sup> |
| Hydrazine                        | N <sub>2</sub> H <sub>4</sub>     | 493                    | 1.54 + 10 <sup>4</sup> | 9.27 + 10 <sup>27</sup> |
| Monomethylhydrazine              | CH <sub>3</sub> NHNH <sub>2</sub> | 4,909                  | 1.07 + 10 <sup>5</sup> | 6.44 + 10 <sup>28</sup> |
| Dinitrogen tetroxide             | N <sub>2</sub> O <sub>4</sub>     | 7950                   | 8.64 + 10 <sup>4</sup> | 5.20 + 10 <sup>28</sup> |

$$5.14 \times 10^7 \text{ mol} \times 2.016 \text{ g/mol} \times 1\text{kg}/1000\text{g} = 1.04 \times 10^5 \text{ kg}$$

$$5.14 \times 10^7 \text{ mol} \times 6.02 \times 10^{23} \text{ particles/mol} = 3.09 \times 10^{31}$$

$$1.93 \times 10^7 \text{ mol} \times 32.00\text{g/mol} \times 1\text{kg}/1000 \text{ g} = 6.18 \times 10^5$$

$$1.16 \times 10^{31} \text{ particles} \times 1\text{mol}/6.02 \times 10^{23} \text{ particles} = 1.93 \times 10^7 \text{ mol}$$

$$493 \text{ kg} \times 1000\text{g/kg} \times 1 \text{ mol}/32.05 \text{ g} = 1.54 \times 10^4$$

$$1.54 \times 10^4 \text{ mol} \times 6.02 \times 10^{23} \text{ particles/mol} = 9.27 \times 10^{27} \text{ particles}$$

$$4,909 \text{ kg} \times 1000 \text{ g/kg} \times 1 \text{ mol}/46.08 \text{ g} = 1.07 \times 10^5$$

$$1.07 \times 10^5 \text{ mol} \times 6.02 \times 10^{23} \text{ particles/mol} = 6.41 \times 10^{28} \text{ particles}$$

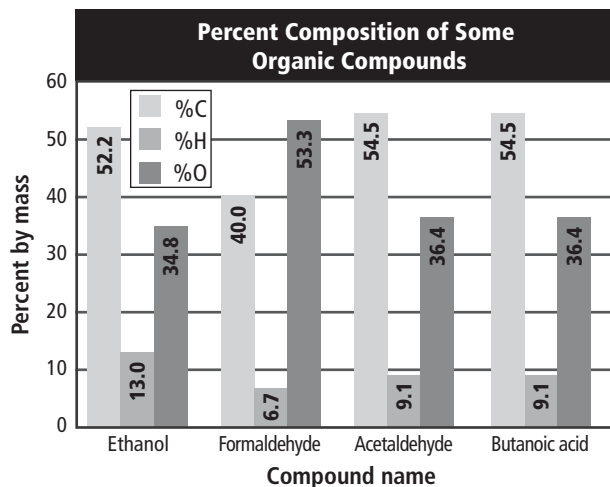
$$8.64 \times 10^4 \text{ mol} \times 92.02 \text{ g/mol} \times 1 \text{ kg}/1000 \text{ g} = 7950 \text{ kg}$$

$$8.64 \times 10^4 \text{ mol} \times 6.02 \times 10^{23} \text{ particles/mol} = 5.20 \times 10^{28} \text{ particles}$$

## Standardized Test Practice

pages 364–365

Use the graph below to answer Questions 1 to 4.



- Acetaldehyde and butanoic acid must have the same
  - molecular formula.
  - empirical formula
  - molar mass.
  - chemical properties.
- If the molar mass of butanoic acid is 88.1 g/mol, what is its molecular formula?
  - C<sub>3</sub>H<sub>4</sub>O<sub>3</sub>
  - C<sub>2</sub>H<sub>4</sub>O
  - C<sub>5</sub>H<sub>12</sub>O<sub>1</sub>
  - C<sub>4</sub>H<sub>8</sub>O<sub>2</sub>

Solution:

Determine the empirical formula. Assume a 100.0 g sample.

$$54.5 \text{ g-C} \times \frac{1 \text{ mol C}}{12.01 \text{ g-C}} = 4.54 \text{ mol C}$$

$$9.1 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g-H}} = 9.03 \text{ mol H}$$

$$36.4 \text{ g-O} \times \frac{1 \text{ mol O}}{16.00 \text{ g-O}} = 2.28 \text{ mol O}$$

Determine the mole ratios.

$$\frac{4.54 \text{ mol C}}{2.28} = 1.99 = 2 \text{ mol C}$$

$$\frac{9.03 \text{ mol H}}{2.28} = 3.96 = 4 \text{ mol H}$$

$$\frac{2.28 \text{ mol O}}{2.28} = 1.00 = 1 \text{ mol O}$$

The empirical formula for butanoic acid is  $\text{C}_2\text{H}_4\text{O}$ .  
The mass of the empirical formula is 44.06 g/mol.

$$\frac{88.1 \text{ g/mol}}{44.06 \text{ g/mol}} = 2.00$$

The molecular formula for butanoic acid =  
 $(\text{C}_2\text{H}_4\text{O})_2 = \text{C}_4\text{H}_8\text{O}_2$ .

3. What is the empirical formula of ethanol?

- a.  $\text{C}_4\text{HO}_3$
- b.  $\text{C}_2\text{H}_6\text{O}$
- c.  $\text{C}_2\text{H}_6\text{O}$
- d.  $\text{C}_4\text{H}_{13}\text{O}_2$

C

Solution:

$$52.2 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.35 \text{ mol C}$$

$$13.0 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 12.9 \text{ mol H}$$

$$34.8 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.18 \text{ mol O}$$

Determine the mole ratios.

$$\frac{4.35 \text{ mol C}}{2.18} = 2.00 \text{ mol C} = 2 \text{ mol C}$$

$$\frac{12.9 \text{ mol H}}{2.18} = 5.92 \text{ mol H} = 6 \text{ mol H}$$

$$\frac{2.18 \text{ mol O}}{2.18} = 1 \text{ mol O} = 1 \text{ mol O}$$

The empirical formula for ethanol is  $\text{C}_2\text{H}_6\text{O}$ .

4. The empirical formula of formaldehyde is the same as its molecular formula. How many grams are in 2.000 mol of formaldehyde?

- a. 30.00 g
  - b. 60.06 g
  - c. 182.0 g
  - d. 200.0 g
- b

Determine the empirical formula for formaldehyde.

$$40.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.33 \text{ mol C}$$

$$6.7 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.65 \text{ mol H}$$

$$53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.33 \text{ mol O}$$

$$\frac{3.33 \text{ mol C}}{3.33} = 1.00 \text{ mol C} = 1 \text{ mol C}$$

$$\frac{6.65 \text{ mol H}}{3.33} = 2.00 \text{ mol H} = 2 \text{ mol H}$$

$$\frac{3.33 \text{ mol O}}{3.33} = 1.00 \text{ mol O} = 1 \text{ mol O}$$

The empirical formula for formaldehyde is  $\text{CH}_2\text{O}$ .

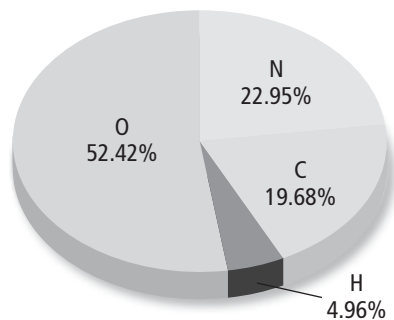
Molar mass of  $\text{CH}_2\text{O} = 30.03 \text{ g/mol}$

$$\frac{30.03 \text{ g}}{1 \text{ mol}} \times 2.00 \text{ mol} = 60.06 \text{ g formaldehyde}$$

5. Which does NOT describe a mole?

- a. a unit used to count particles directly
  - b. Avogadro's number of molecules of a compound
  - c. the number of atoms in exactly 12 g of pure C-12
  - d. the SI unit for the amount of a substance
- a

Use the graph below to answer Question 6.



6. What is the empirical formula for this compound?

- $C_6H_{12}N_6O_3$
- $C_4HN_5O_{10}$
- $CH_3NO_2$
- $CH_5NO_3$

c

Assume a 100.00 g sample.

$$19.68 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.639 \text{ mol C}$$

$$4.96 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.921 \text{ mol H}$$

$$22.95 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 1.638 \text{ mol N}$$

$$52.42 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 3.276 \text{ mol O}$$

Determine the mole ratios.

$$\frac{1.639 \text{ mol C}}{1.638} = 1.001 \text{ mol C} = 1 \text{ mol C}$$

$$\frac{4.921 \text{ mol H}}{1.638 \text{ mol}} = 3.004 \text{ mol H} = 3 \text{ mol H}$$

$$\frac{1.638 \text{ mol N}}{1.638} = 1 \text{ mol N} = 1 \text{ mol N}$$

$$\frac{3.276 \text{ mol O}}{1.638} = 2 \text{ mol O} = 2 \text{ mol O}$$

The empirical formula for this compound is  $CH_3NO_2$ .

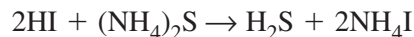
7. Which is NOT true of molecular compounds?

- Triple bonds are stronger than single bonds.
- Electrons are shared in covalent bonds.

- All atoms have eight valence electrons when they are chemically stable.
- Lewis structures show the arrangement of electrons in covalent molecules.

c

8. Which type of reaction is shown below?



- synthesis
- decomposition
- single replacement
- double replacement

d

9. How many atoms are in 0.625 moles of GE (atomic mass = 72.59 amu)?

- $2.73 \times 10^{25}$
- $6.99 \times 10^{25}$
- $03.76 \times 10^{23}$
- $9.63 \times 10^{23}$

c

$$0.625 \text{ mol Ge} \times \frac{6.02 \times 10^{23}}{1 \text{ mol Ge}} = 3.76 \times 10^{23} \text{ atoms Ge}$$

10. What is the mass of one molecule of barium hexafluorosilicate ( $\text{BaSiF}_6$ )?

- $1.68 \times 10^{26} \text{ g}$
- $2.16 \times 10^{21} \text{ g}$
- $4.64 \times 10^{-22}$
- $6.02 \times 10^{-23}$

c

Determine the molar mass of  $\text{BaSiF}_6$ .

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{1 \text{ mol Ba}} = 137.33 \text{ g Ba}$$

$$1 \text{ mol Si} \times \frac{28.09 \text{ g Si}}{1 \text{ mol Si}} = 28.09 \text{ g Si}$$

$$6 \text{ mol F} \times \frac{19.00 \text{ g F}}{1 \text{ mol F}} = 114.00 \text{ g F}$$

$$\text{Molar mass} = 279.42 \text{ g/mol BaSiF}_6$$

$$\frac{279.42 \text{ g BaSiF}_6}{1 \text{ mol BaSiF}_6} \times \frac{1 \text{ mol BaSiF}_6}{6.02 \times 10^{23} \text{ g/molecule BaSiF}_6} = 4.64 \times 10^{-22} \text{ g/molecule BaSiF}_6$$

Use the table below to answer question 11.

| Charges of Some Ions |               |
|----------------------|---------------|
| Ion                  | Formula       |
| Sulfide              | $S^{-2}$      |
| Sulfite              | $SO_3^{-2}$   |
| Sulfate              | $SO_4^{-2}$   |
| Thiosulfate          | $S_2O_3^{-2}$ |
| Copper(I)            | $Cu^+$        |
| Copper(II)           | $Cu^{2+}$     |

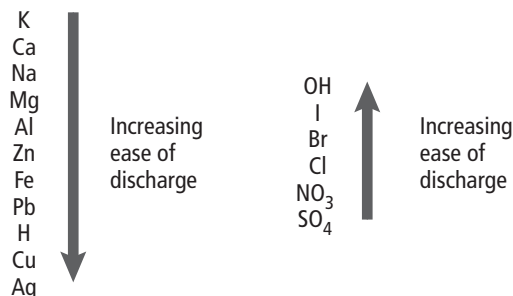
11. How many possible compounds can be made that contain copper, sulfur, and oxygen? Write their names and formulas.

Six compounds are possible.

|                        |              |
|------------------------|--------------|
| Copper(I) thiosulfate  | $Cu_2S_2O_3$ |
| Copper(II) thiosulfate | $CuS_2O_3$   |
| Copper(I) sulfate      | $Cu_2SO_4$   |
| Copper(II) sulfate     | $CuSO_4$     |
| Copper(I) sulfite      | $Cu_2SO_3$   |
| Copper(II) sulfite     | $CuSO_3$     |

## Extended Response

Use the figure below to answer Question 12.



12. You have been asked to identify a sample of a metal. It is known to be either zinc, lead, or lithium. You have aqueous solutions of KCl,  $AlCl_3$ ,  $FeCl_3$ , and  $CuCl_2$  available. Explain how you would use these solutions to identify what metal your sample is made of.

You would need to test each sample of metal for a reaction with each of the solutions in turn. If the sample is lithium, it is more reactive than any of the four aqueous chloride solutions, and will display a reaction when placed into any of them. Lead is less reactive than any of the aqueous solutions except for the copper; therefore, if the metal reacts with the copper solution but none of the others, it is likely to be lead. Zinc is more reactive than iron and copper but less reactive than potassium and aluminum, so if the unidentified metal is zinc it will react with the iron and copper solutions but not with the potassium or aluminum solutions.

13. It takes 2 iron atoms and 6 chlorine atoms to make 2 iron (III) chloride particles. How many chlorine atoms are required to make 18 iron (III) chloride particles?
- 9
  - 18
  - 27
  - 54
  - 72
- d

$$\frac{6 \text{ chlorine atoms}}{2 \text{ iron (III) chloride}} = \frac{x \text{ chlorine atoms}}{18 \text{ iron (III) chloride}}$$

$$x = 54 \text{ chlorine atoms}$$

14. What is the molar mass of fluorapatite ( $\text{Ca}_5(\text{PO}_4)_3\text{F}$ )?

- 314 g/mol
- 344 g/mol
- 442 g/mol
- 504 g/mol
- 534 g/mol

$$5 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 200.40 \text{ g Ca}$$

$$3 \text{ mol P} \times \frac{30.97 \text{ g P}}{1 \text{ mol P}} = 92.91 \text{ g P}$$

$$12 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol P}} = 192.00 \text{ g O}$$

$$1 \text{ mol F} \times \frac{19.00 \text{ g F}}{1 \text{ mol F}} = 19.00 \text{ g F}$$

$$\text{molar mass} = 504.31 \text{ g/mol} = 504 \text{ g/mol}$$

15. Which is not a correct formula for an ionic compound?

- $\text{CaC}_{12}$
- $\text{Na}_2\text{SO}_4$
- $\text{Al}_3\text{S}_2$
- $\text{Mg}(\text{NO}_3)_2$
- $\text{NaCl}$

Use the table below to answer Question 16.

| Percent composition of selected hydrocarbons |       |       |       |
|--|-------|-------|-------|
| Compound                                     | % C   | % H   | % O   |
| $\text{C}_2\text{H}_{10}\text{O}$            | 64.81 | 13.60 | 21.59 |
| $\text{C}_6\text{H}_{12}\text{O}_4$          | 48.64 | 8.108 | 43.24 |
| $\text{C}_7\text{H}_{16}\text{O}_3$          | 56.76 | 10.81 | 32.43 |
| $\text{C}_5\text{H}_8\text{O}_5$             | 40.54 | 5.405 | 54.05 |

16. A 25.0-g sample of an unknown hydrocarbon is composed of 12.16 g carbon, 2.027 g hydrogen, and 10.81 g oxygen. If its molecular weight is 148 g/mol, what is the molecular formula for this compound?

- $\text{C}_2\text{H}_{10}\text{O}$
- $\text{C}_6\text{H}_{12}\text{O}_4$
- $\text{C}_7\text{H}_{16}\text{O}_3$
- $\text{C}_5\text{H}_8\text{O}_5$
- $\text{C}_8\text{H}_5\text{O}_5$

Determine the number of moles.

$$12.16 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.012 \text{ mol C}$$

$$2.027 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 2.007 \text{ mol H}$$

$$10.81 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.6756 \text{ mol O}$$

Determine the empirical formula.

$$\frac{1.012 \text{ mol C}}{0.6756} = 1.498 = 1.5 \times 2 = 3$$

$$\frac{2.007 \text{ mol H}}{0.6756} = 2.971 = 3 \times 2 = 6$$

$$\frac{6.756 \text{ mol O}}{0.6756} = 1.000 = 1 \times 2 = 2$$

Empirical formula is  $\text{C}_3\text{H}_6\text{O}_2$ .

Determine the mass of the empirical formula.

$$3 \times 12.01 \text{ g/mol} = 36.03 \text{ g/mol}$$

$$6 \times 1.01 \text{ g/mol} = 6.06 \text{ g/mol}$$

$$2 \times 16.00 \text{ g/mol} = 32.00 \text{ g/mol}$$

The mass of the empirical formula is 74.09 g/mol.

$$\frac{148 \text{ g/mol}}{74.09 \text{ g/mol}} = 2$$

The molecular formula is  $\text{C}_6\text{H}_{12}\text{O}_4$ .