10.3 Percent Composition and Chemical Formulas

Connecting to Your World
If your shirt is made of 100 percent cotton or wool, or is the fabric a combination of two or more fibers? A tag sewn into the seam of the shirt usually tells you what fibers were used to make the cloth and the percent of each. It helps to know the percents of the components in the shirt because they affect how the shirt “green” the grass. In fall, you may want to use a fertilizer with a higher percent of nitrogen to help grass grow. In spring, you may use a fertilizer that has a relatively high percent of potassium to help flowers bloom and fruits to grow. The percents of the nutrients in fertilizer is important. In spring, flowers need a high amount of potassium to grow. In fall, you may want to use a fertilizer with a high percent of nitrogen to help grass grow. If you have had experience with lawn care, you know that the relative amounts of the components in a mixture or compound is often useful.

The Percent Composition of a Compound
If you have had experience with lawn care, you know that the relative amount, or the percent, of each nutrient in fertilizer is important. In spring, you may use a fertilizer that has a relatively high percent of nitrogen to “green” the grass. In fall, you may want to use a fertilizer with a higher percent of potassium to strengthen the root system. Knowing the relative amounts of the components of a mixture or compound is often useful.

The relative amounts of the elements in a compound are expressed as the percent composition or the percent by mass of each element in the compound. The percent composition of a compound consists of a percent value for each different element in the compound. As you can see in Figure 10.13, the percent composition of K2Cr2O7 is K = 40.3%, Cr = 35.4%, and O = 26.8%. These percents must total 100%. (40.3% + 35.4% + 26.8% = 100%). The percent by mass of an element in a compound is the number of grams of the element divided by the mass in grams of the compound, multiplied by 100%.

\[
\text{Percent composition} \times 100\% = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%
\]

Percent Composition from Mass Data
Imagine you are a chemist who has just finished the synthesis of a new compound. You have purified your product and stored the crystalline solid in a vial. Now you must verify the composition of your new compound and determine its molecular formula. You use analytical procedures to determine the relative masses of each element in the compound and calculate the percent composition.

Figure 10.13 Potassium chromate (K2CrO4) is composed of 40.3% potassium, 35.4% chromium, and 26.8% oxygen. Interpreting Diagrams How does this percent composition differ from the percent composition of potassium dichromate (K2Cr2O7), a compound composed of the same three elements?

Guide for Reading

Key Concepts
- How do you calculate the percent by mass of an element in a compound?
- What does the empirical formula of a compound show?
- How does the molecular formula of a compound compare with the empirical formula?

Vocabulary
percent composition
empirical formula
molecular formula

Comparing and Contrasting
When you compare and contrast things, you examine how they are alike and different. As you read, list the similarities and differences between empirical and molecular formulas.

Guide for Reading

Build Vocabulary
Paraphrase Have students read the definition of percent composition on this page. Then, have them define the term according to the definition in the dictionary. Their definitions should indicate that percent composition refers to the relative size of parts that make up 100 percent of something.

Connecting to Your World
Ask, What should be the total of the percents listed on the label? (100%) If the shirt were 25% nylon, what percent would be cotton? (75%)

The Percent Composition of a Compound

Use Visuals

Figure 10.13 Have students study the figure and read the text on percent composition. Point out that the three numbers in each circle graph add up to a total of 100%. Ask, Which compound is a better source of potassium? (K2CrO4)

Answers to...

Figure 10.13 The percent composition of K2Cr2O7 is 26.5% K, 35.4% Cr, and 38.1% O.
Section 10.3 (continued)

Sample Problem 10.9

Calculating Percent Composition from Mass Data

When a 13.60-g sample of a compound containing only magnesium and oxygen is decomposed, 5.40 g of oxygen is obtained. What is the percent composition of this compound?

Analyze

List the knowns and the unknowns.

Knowns

• mass of compound = 13.60 g
• mass of oxygen = 5.40 g O

Unknowns

• percent Mg
• percent O

The percent by mass of an element in a compound is the mass of that element divided by the mass of the compound multiplied by 100%.

Calculate

Solve for the unknown.

\[
\% \text{Mg} = \frac{\text{mass of Mg}}{\text{mass of compound}} \times 100% = \frac{8.20 \text{ g}}{13.60 \text{ g}} \times 100% = 60.3% \\
\% \text{O} = \frac{\text{mass of O}}{\text{mass of compound}} \times 100% = \frac{5.40 \text{ g}}{13.60 \text{ g}} \times 100% = 39.7%
\]

Evaluate

Does the result make sense?

The percents of the elements add up to 100%:

60.3% Mg + 39.7% O = 100%

Practice Problems

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

33. When a 14.2-g sample of mercury(II) oxide is decomposed into its elements by heating, 13.2 g Hg is obtained. What is the percent composition of the compound?

Facts and Figures

Parts per Million and Parts per Billion

Percents are used to show relative parts of mixtures as well as the composition of a compound. But when an extremely small amount of a substance is present in a large amount of another substance, it might not be practical to use percents (parts per one hundred) to show the makeup of the mixture. Instead, concentrations of extremely dilute solutions are sometimes measured in units of parts per million (ppm) or parts per billion (pppb). For example, the composition of a mixture that consists of 1 gram of a substance per 100 grams of water (or 1 milligram of substance per liter of water) can be expressed as 1 ppm.
Chemical Quantities

Sample Problem 10.10

Calculating Percent Composition from a Chemical Formula

Propane (C₃H₈), the fuel commonly used in gas grills, is one of the compounds obtained from petroleum. Calculate the percent composition of propane.

1. Analyze List the knowns and the unknowns.
   - Knowns:
     - mass of C in 1 mol C₃H₈ = 36.0 g
     - mass of H in 1 mol C₃H₈ = 8.0 g
     - molar mass of C₃H₈ = 44.0 g/mol
   - Unknowns:
     - % C
     - % H

2. Calculate Solve for the unknowns.
   - % C = \( \frac{\text{mass of } C}{\text{mass of propane}} \times 100\% = \frac{36.0 \, \text{g}}{44.0 \, \text{g}} \times 100\% = 81.8\% \)
   - % H = \( \frac{\text{mass of } H}{\text{mass of propane}} \times 100\% = \frac{8.0 \, \text{g}}{44.0 \, \text{g}} \times 100\% = 18\% \)

3. Evaluate Does the result make sense?
   - The percents of the elements add up to 100% when the answers are expressed to two significant figures.

Practice Problems

34. Calculate the percent composition of these compounds.
   - a. ethane (C₂H₆)
   - b. sodium hydrogen sulfate (NaHSO₄)

35. Calculate the percent nitrogen in these common fertilizers.
   - a. NH₃
   - b. NH₄NO₃

The percent composition of a compound is always the same, as Figure 10.14 on the preceding page indicates.

\( \text{% mass} = \frac{\text{mass of element in 1 mol compound}}{\text{molar mass of compound}} \times 100\% \)

Answers

34. a. 24.0 g C/30.0 g \times 100\% = 80.0\% C
   - b. 6.00 g H/30.0 g \times 100\% = 20.0\% H

35. a. 14.0 g N/17.0 g \times 100\% = 82.4\% N
   - b. 28.0 g N/80.0 g \times 100\% = 35.0\% N

Practice Problems Plus

1. Determine the percent composition of the following oxides:
   - a. Fe₂O₃ (69.9\% Fe, 30.1\% O)
   - b. HgO (92.6\% Hg, 7.39\% O)
   - c. Ag₂O (93.1\% Ag, 6.90\% O)
   - d. Na₂O (74.2\% Na, 25.8\% O)

2. Calculate the grams of oxygen in 90.0 g of Cl₂O. (16.6 g)

Math Handbook

For a math refresher and practice, direct students to significant figures, page R59.

Download a worksheet on Percent Composition for students to complete, and find additional teacher support from NSTA SciLinks.

Answers to...

Checkpoin... Divide the mass of the element in one mole of the compound by the molar mass and multiply by 100%.

Differentiated Instruction

Gifted and Talented

Have students research the formulas of the three different oxides of iron. Ask, Which of the oxides contains a higher percent of iron? (Of FeO (77.7\% Fe), Fe₂O₃ (69.9\% Fe), and Fe₃O₄ (72.3\% Fe), FeO has the highest percent of iron.)
Quick LAB

Percent Composition

Objective After completing this activity, students will be able to:
• determine the percent of water in a hydrate.

Skills Focus observing, calculating

Prep Time 20 minutes

Class Time 30 minutes

Safety Students should wear safety goggles and tie back loose hair. Caution students that while heating test tubes, they should not aim the opening of the tube toward anyone. Tell them to move the test tube in the flame and not to heat one spot excessively. CAUTION! Be sure that students allow the tubes to cool completely before they touch them. Hot glass looks exactly like cold glass!

Teaching Tips For best results, students should do a second heating and cooling of each sample to determine whether all of the water has been driven off.

Expected Outcome See data table at the bottom of the page.

Think About It
1–3. See data table.
4. The hydrated salt of sodium sulfate lost the greatest percent. The hydrated salt of calcium chloride lost the smallest percent.

For Enrichment

Have students design and conduct a similar experiment to determine the percent of oxygen in potassium chlorate. Tell students that when potassium chlorate is heated, potassium chlorate and oxygen are produced, $2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2$. For classroom safety, no more than 5 g of potassium chlorate should be used. Results should show that potassium chlorate is approximately 39% oxygen.

Quick LAB

Percent Composition

Purpose To measure the percent of water in a series of crystalline compounds called hydrates.

Materials
• centigram balance
• Bunsen burner
• 2 medium-sized test tubes
• test tube holder
• test tube rack
• spatula
• hydrated salts of copper(II) sulfate, calcium chloride, and sodium sulfate

Procedure
1. Label each test tube with the name of a salt. Measure and record the masses.
2. Add 2–3 g of salt (a good-sized spatula full) to the appropriately labeled test tube. Measure and record the mass of each test tube and salt.
3. Hold one of the tubes at a 45° angle and gently heat its contents over the burner, slowly passing it in and out of the flame. Note any change in the appearance of the solid salt.
4. As moisture begins to condense in the upper part of the test tube, gently heat the entire length of the tube. Continue heating until all of the moisture is driven from the tube. This may take 2–3 minutes. Repeat Steps 3 and 4 for the other two tubes.
5. Allow each tube to cool. Then measure and record the mass of each test tube and the heated salt.

Analyze and Conclude
1. Set up a data table so that you can subtract the mass of the empty tube from the mass of the salt and the test tube, both before and after heating.
2. Calculate the difference between the mass of each salt before and after heating. This difference represents the amount of water lost by the hydrate on heating.
3. Calculate the percent by mass of water lost by each compound. Which compound lost the greatest percent by mass of water? The smallest?

Percent Composition as a Conversion Factor

You can use percent composition to calculate the number of grams of any element in a specific mass of a compound. To do this, multiply the mass of the compound by the percent composition of the element in the compound. Suppose you want to know how much carbon and hydrogen are contained in 82.0 g of propane. In Sample Problem 10.10, you found that propane is 81.8% carbon and 18% hydrogen. That means that in a 100-g sample of propane, you would have 81.8 g of carbon and 18 g of hydrogen. You can use the ratio $\frac{81.8 \text{ g C}}{100 \text{ g propane}}$ to calculate the mass of carbon contained in 82.0 g of propane: $\frac{81.8 \text{ g C}}{100 \text{ g propane}} \times 82.0 \text{ g propane} = 67.1 \text{ g C}$

Using the ratio $\frac{18 \text{ g H}}{100 \text{ g propane}}$, you can calculate the mass of hydrogen: $\frac{18 \text{ g H}}{100 \text{ g propane}} \times 82.0 \text{ g propane} = 14.8 \text{ g H}$

The sum of the two masses equals 82.0 g, the sample size, to two significant figures ($\text{67.1 g C} + \text{14.8 g H} = 82.0 \text{ g}$).

Checkpoint How many grams of hydrogen are contained in a 100-g sample of propane?

Data Table with Sample Data

<table>
<thead>
<tr>
<th>Compound</th>
<th>CuSO$_4 \cdot 5\text{H}_2\text{O}$</th>
<th>CaCl$_2 \cdot 2\text{H}_2\text{O}$</th>
<th>Na$_2$SO$_4 \cdot 10\text{H}_2\text{O}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Test tube + hydrate (before heating)</td>
<td>23.88 g</td>
<td>23.60 g</td>
<td>23.92 g</td>
</tr>
<tr>
<td>Empty test tube</td>
<td>21.19 g</td>
<td>21.25 g</td>
<td>21.17 g</td>
</tr>
<tr>
<td>Mass of hydrate</td>
<td>2.69 g</td>
<td>2.35 g</td>
<td>2.75 g</td>
</tr>
<tr>
<td>Test tube + salt (after heating)</td>
<td>22.88 g</td>
<td>23.07 g</td>
<td>22.71 g</td>
</tr>
<tr>
<td>Empty test tube</td>
<td>21.19 g</td>
<td>21.25 g</td>
<td>21.17 g</td>
</tr>
<tr>
<td>Mass of anhydrous salt</td>
<td>1.69 g</td>
<td>1.82 g</td>
<td>1.54 g</td>
</tr>
<tr>
<td>Mass of water lost</td>
<td>1.00 g</td>
<td>0.33 g</td>
<td>1.21 g</td>
</tr>
<tr>
<td>Percent water (experimental)</td>
<td>37.2%</td>
<td>22.6%</td>
<td>44.0%</td>
</tr>
<tr>
<td>Percent water (theoretical)</td>
<td>36.1%</td>
<td>24.5%</td>
<td>55.9%</td>
</tr>
</tbody>
</table>
Empirical Formulas

A useful formula for cooking rice is to use one cup of rice and two cups of water. If a larger amount of rice is needed, you could double or triple the amounts, for example, two cups of rice and four cups of water. The formulas for some compounds also show a basic ratio of elements. Multiplying that ratio by any factor can produce the formulas for other compounds.

The percent composition of your newly synthesized compound is the data you need to calculate the basic ratio of the elements contained in the compound. The basic ratio, called the empirical formula, gives the lowest whole-number ratio of the atoms of the elements in a compound. For example, a compound may have the empirical formula \( \text{CO}_2 \). The empirical formula shows the kinds and lowest relative count of atoms or moles of atoms in molecules or formula units of a compound. Figure 10.15 shows that empirical formulas may be interpreted at the microscopic (atomic) or macroscopic (molar) level.

An empirical formula may or may not be the same as a molecular formula. For example, the lowest ratio of hydrogen to oxygen in hydrogen peroxide is \( 1:1 \). Thus the empirical formula of hydrogen peroxide is \( \text{HO} \). The actual molecular formula of hydrogen peroxide has twice the number of oxygen atoms; \( 2 \text{ mol O atoms} \). But notice that the ratio of hydrogen to oxygen is still the same, \( 1:1 \).

The empirical formula of a compound shows the smallest whole-number ratio of the atoms in the compound. The molecular formula tells the actual number of each kind of atom present in a molecule of the compound. For carbon dioxide, the empirical and molecular formulas are the same—\( \text{CO}_2 \). Figure 10.16 shows two compounds of carbon having the same empirical formula (\( \text{CH} \)) but different molecular formulas.

**Differentiated Instruction**

**Less Proficient Readers**

As students read about how to determine empirical and molecular formulas, have groups of students develop numbered lists of steps they would take to determine these formulas. They should have three lists: one for determining empirical formulas from percent composition; one for determining empirical formulas from mass data; and one for determining molecular formulas from the empirical formula and molar mass.

**Word Origins**

Empirical comes from the Latin word empiricus meaning a doctor relying on experience alone. An empirical formula must be obtained from experimental data. Thus, an empirical formula relies on experience. Is a molecular formula also based on experimental data?

**Empirical Formulas from Percent Composition**

**Purpose**

Students are provided with an analogy that helps clarify the concepts of percent composition and empirical formulas.

**Materials**

3 red marbles, 6 green marbles, 3 black marbles, and 12 blue marbles

**Procedure**

Provide pairs of students with sets of marbles. Have students express the number of different colored marbles as fractions and percents of the whole collection. Ask, What percent of the collection do the red marbles represent? (12.5%) Show them that the sums of the fractions and percents are equal to 1 and 100%, respectively. Ask, What is the ratio of red: green:black:blue marbles in lowest terms? (1:2:1:4) This activity can be extended if the different colored marbles are assumed to be atoms of different elements. Ask, What is the empirical formula of a hypothetical “compound” that consists of 25% red marbles and 75% green marbles? (The ratio of red marbles to green marbles in the empirical formula would be 1:3.)

**Expected Outcomes**

Students express percent composition of the marbles and determine the “empirical formula” of a marble combination.

**Word Origins**

A molecular formula is based on experimental data in two different ways. The molar mass is determined experimentally, as is the empirical formula.
Section 10.3 (continued)

Sample Problem 10.11

Answers
36. a. 94.1 g O × 1 mol O/16.0 g O = 5.88 mol O
5.9 g H × 1 mol H/1.0 g H = 5.9 mol H
5.88 mol O/5.88 = 1.00 mol O
5.9 mol H/5.88 = 1.0 mol H
Empirical formula = HO

b. 67.6 g Hg × 1 mol Hg/200.6 g Hg = 0.337 mol Hg
10.8 g S × 1 mol S/32.1 g S = 0.336 mol S
21.6 g O × 1 mol O/16.0 g O = 1.35 mol O
0.337 mol Hg/0.336 = 1.00 mol Hg
0.336 mol S/0.336 = 1.00 mol S
1.35 mol O/0.336 = 4.02 mol O
Empirical formula = HgSO₄

37. 62.1 g C × 1 mol C/12.0 g C = 5.18 mol C
13.8 g H × 1 mol H/1.00 g H = 13.8 mol H
24.1 g N × 1 mol N/14.0 g N = 1.72 mol N
5.18 mol C/1.72 = 3.01 mol C
13.8 mol H/1.72 = 8.02 mol H
1.72 mol N/1.72 = 1.00 mol N
Empirical formula = C₃H₈N

Practice Problems Plus
What is the empirical formula of each of the following compounds?
a. 36.1% Ca, 63.9% Cl (CaCl₂)
b. 40.0% C, 6.7% H, 53.3% O (CH₂O)
c. 3.7% H, 44.4% C, and 51.9% N (HCN)

Facts and Figures
Computing Formulas
In using percent composition to determine empirical formula, 100.0 g of compound is arbitrarily chosen because it is easy to use. If an element comprises 28.5% of the mass of a compound, for example, it makes up 28.5 g of a 100.0-g sample. Any other mass of compound can be used but computation will be more difficult.
Table 10.3

<table>
<thead>
<tr>
<th>Formula (name)</th>
<th>Classification of formula</th>
<th>Molar mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH</td>
<td>Empirical</td>
<td>13</td>
</tr>
<tr>
<td>C₂H₂ (ethyne)</td>
<td>Molecular</td>
<td>26 (2 × 13)</td>
</tr>
<tr>
<td>C₆H₆ (benzene)</td>
<td>Molecular</td>
<td>78 (6 × 13)</td>
</tr>
<tr>
<td>CH₂O₂ (methanal)</td>
<td>Empirical and Molecular</td>
<td>30</td>
</tr>
<tr>
<td>C₆H₁₂O₆ (glucose)</td>
<td>Molecular</td>
<td>180 (6 × 30)</td>
</tr>
</tbody>
</table>

Molecular Formulas

Look at the compounds listed in Table 10.3. Ethyne and benzene have the same empirical formula—CH. Methanal, ethanoic acid, and glucose, shown in Figure 10.17 have the same empirical formula—CH₂O. But the compounds in these two groups have different molar masses. Their molar masses are simple whole-number multiples of the molar masses of the empirical formulas, CH and CH₂O. The molecular formula of a compound is either the same as its experimentally determined empirical formula, or it is a simple whole-number multiple of its empirical formula.

Once you have determined the empirical formula of your newly synthesized compound, you can determine its molecular formula, but you must know the compound’s molar mass. A chemist often uses an instrument called a mass spectrometer to determine molar mass. The compound is broken into charged fragments (ions) that travel through a magnetic field. The magnetic field deflects the particles from their straight-line paths. The mass of the compound is determined from the amount of deflection experienced by the particles.

From the empirical formula, you can calculate the empirical formula mass (efm). This is simply the molar mass represented by the empirical formula. Then you can divide the experimentally determined molar mass by the empirical formula mass. This gives the number of empirical formula units in a molecule of the compound and is the multiplier to convert the empirical formula to the molecular formula. For example, recall that the empirical formula of hydrogen peroxide is HO. Its empirical formula mass is 17.0 g/mol. The molar mass of H₂O₂ is 34.0 g/mol.

To obtain the molecular formula of hydrogen peroxide from its empirical formula, multiply the subscripts in the empirical formula by 2: H₂O₂ → 2H₂O₂.

How does the molecular formula for a compound relate to its empirical formula?

Facts and Figures

Some Carbohydrates Share the Same Empirical Formula

Several carbohydrates have the empirical formula CH₉O₆. Examples include glucose, which is abundant in plants and animals, and fructose, which is found in fruits and honey. Both of these simple structures have the molecular formula C₆H₁₂O₆, but they differ in structure.
Sample Problem 10.12

Answers
38. molar mass/efm = 62/31 = 2

molecular formula = 2(CH$_2$O) = C$_2$H$_4$O$_2$

39. a. same empirical formula (CH$_2$O)

b. different empirical formulas

Practice Problems Plus
1. What is the molecular formula of a compound with the empirical formula CCIN and a molar mass of 184.5?

2. What is the molecular formula of a compound that is 56.6% K, 8.7% C, and 34.7% O?

I ASSESS

Evaluate Understanding
Have students list the steps they would take to calculate the molecular formula in each of the following situations:
• The empirical formula and molar mass are known.

• The percent composition and molar mass are known.

Reteach
Point out to students that when they know the percent composition and molar mass of a compound, they must first use the percent composition to calculate the empirical formula. They can then calculate the empirical formula mass and compare it to the molar mass of the molecular compound to determine the molecular formula.

CaO: 71.5%
CaCO$_3$: 40.1%
Ca(OH)$_2$: 54.1%
CaSO$_4$•2H$_2$O: 23.3%
Ca$_3$(PO$_4$)$_2$: 38.8%

Element Handbook

ASSESSMENT

10.3 Section Assessment

38. Find the molecular formula of ethylene glycol, which is used as antifreeze. The molar mass is 62 g/mol and the empirical formula is CH$_2$O.

39. Which pair of molecules has the same empirical formula?

a. C$_2$H$_4$O$_2$, C$_3$H$_6$O$_3$

b. Na$_2$O, Na$_3$CO$_3$

c. CaCO$_3$, MgCO$_3$

d. MgCl$_2$, Mg(NO$_3$)$_2$

40. What is an empirical formula? Which of the following molecular formulas are also empirical formulas?

a. glucose (C$_6$H$_12$O$_6$)

b. ethyl butyrate (C$_6$H$_12$O$_2$)

c. chlorophyll (C$_55$H$_72$N$_4$O$_4$)

d. DEET (C$_8$H$_17$NO$_3$)

Element Handbook

Calcium: Select three important compounds that contain calcium from among those discussed on page R11 of the Elements Handbook. Determine the percent of calcium in each.

ASAP Test yourself on the concepts in Section 10.3.

chem_TE_ch10_IPL.fm  Page 312  Thursday, August 5, 2004  12:06 AM
Drugs in the Blood
Current drug tests can detect even small traces of drugs in the blood or urine. For example, THC (the metabolic product of marijuana use) remains detectable in the body for at least 18 hours for an occasional user and up to 30 days for a habitual user. It can be detected in concentrations as small as 50 ng/mL of solution.