Chapter

4

Quantities in Chemical Formulas

In this chapter, you will be able to

- explain the law of definite proportions and the significance of different proportions of elements in compounds;
- explain the relationship between isotopic abundance and relative atomic mass;
- describe and explain Avogadro's constant, the mole concept, and the relationship between the mole and molar mass;
- determine empirical and molecular formulas using percentage composition obtained from given data and through experimentation;
- develop technological skills for quantitative analysis;
- solve problems involving quantities in moles, numbers of particles, and mass numbers.

When you read the list of ingredients on a package of cereal, do you also notice how much of each ingredient is contained in a serving? We can compare the quantities of sugar or fat or the percentage of daily requirements of vitamins and minerals in different brands (**Figure 1**). This quantitative information helps us decide which product to select to suit our needs.

Quantities in chemical formulas offer similar important information about the composition and properties of compounds. For example, water ($H_2O_{(l)}$) and hydrogen peroxide ($H_2O_{2(l)}$) both contain the same types of atoms. The only difference is in the number of oxygen atoms. This difference, which appears small, actually results in significant differences in the properties of the two compounds.

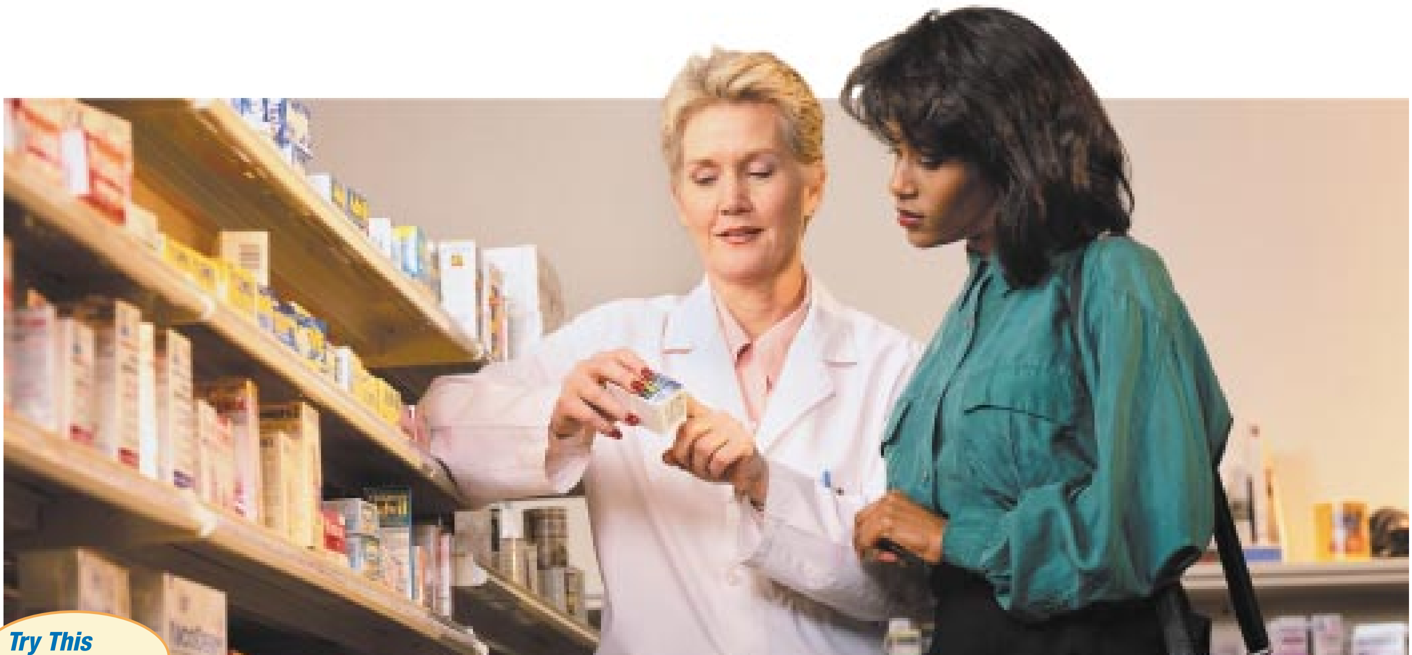
Water is very stable, can be stored safely for indefinite periods, and can be used for drinking and washing. It is an important ingredient of living cells and is essential for all life on Earth. Hydrogen peroxide, on the other hand, is so unstable that it must be stored in darkened containers to slow its decomposition. Concentrated solutions of hydrogen peroxide must be used with caution, because the chemical reacts readily with other substances and will cause blistering of the skin on contact. Because hydrogen peroxide is toxic it is used to kill bacteria—in low concentrations (0.03%) it is used as an antiseptic to treat minor cuts and abrasions. At higher concentrations (6%) it is used to bleach hair, pulp and paper, and synthetic and natural fibres. At even higher concentrations, its bactericidal properties can be applied as part of the treatment of waste water. At sufficiently high concentrations it is explosive.

How can we determine the chemical composition of a substance? Identifying the substance's properties is one method. In this chapter, you will learn other methods to do this. You will also learn how to measure and communicate quantities when dealing with entities as small as atoms, ions, and molecules.

Reflect on your Learning

1. The label on a bag of jelly beans states that the bag contains 40% large jelly beans and 60% small jelly beans, by mass. Do you think this information is sufficient if we want to know how many jelly beans of each size there are in the bag? Explain your answer.
2. Given your knowledge of chemical reactions, list reasons why you think it is important to be able to communicate information about the number of atoms, ions, or molecules that are reacting or that are produced.
3. Many fossil fuels that are burned in factories contain sulfur; when sulfur reacts with oxygen in the air, sulfur oxides—known air pollutants—are produced. Technicians use different methods to predict the masses of these and other chemicals released into the air. Suggest reasons why it is important to be able to make these predictions. Speculate on how technicians can make this kind of analysis.

Throughout this chapter, note any changes in your ideas as you learn new concepts and develop your skills.



Try This Activity

Do Ions Combine in Definite Ratios?

In this activity you will mix copper(II) ions and carbonate ions in different ratios and see how they combine. The solutions of ions will each contain the same number of ions per unit volume.

Copper(II) sulfate is poisonous if ingested.
Wash hands after use.

Materials: lab apron, eye protection, 5 small test tubes of equal size, test-tube rack, eyedropper, 10 mL copper(II) sulfate solution (0.10 mol/L), 10 mL sodium carbonate solution (0.10 mol/L), distilled water

- Number the test tubes from 1 to 5, and place in the test-tube rack.
- Using the dropper, add drops of copper(II) sulfate solution to each test tube, according to **Table 1**.
- Wash the dropper thoroughly with distilled water and use the same dropper to add drops of sodium carbonate solution to each test tube, according to **Table 1**. After you have finished putting drops in each tube, the test tubes should be filled to equal depth since they contain the same number of drops (10 drops total).
- Swirl each test tube gently to mix the contents. Allow the precipitates to settle for about 5 min.
- Wash your hands thoroughly.
  (a) What is the ratio of Cu^{2+} ions to CO_3^{2-} ions in each test tube?
  (b) Which test tube contains the most precipitate ($CuCO_{3(s)}$)? What is the ratio of Cu^{2+} ions to CO_3^{2-} ions in this test tube?
  (c) From what you learned in Unit 1 about ionic bonds, does this ratio agree with a prediction of how copper(II) ions and carbonate ions would combine to form a compound?
  (d) Which test tubes contain the smallest amount of precipitate? Suggest reasons why the ratios of ions in these test tubes produced the least amounts.
  (e) What evidence is there that some copper ions remain unused in solution in some tubes?
  (f) Explain how the evidence suggests that ions combine in definite ratios.
- Dispose of the materials according to your teacher's instructions.

Figure 1
Quantities in a list of ingredients help us compare and select products to suit our needs.

Table 1: Mixing Ions

	Test tube				
	1	**2**	**3**	**4**	**5**
drops $Cu^{2+}_{(aq)}$	1	3	5	7	9
drops $CO^{2-}_{3(aq)}$	9	7	5	3	1