The statement “one plus one equals two” can be represented by the mathematical equation \(1 + 1 = 2\). Mathematicians use symbols for numbers and operations as a kind of shorthand that conveniently and clearly expresses mathematical concepts. Scientists use a different type of shorthand to represent chemical concepts. As you have seen in this unit, scientists use chemical symbols, such as H and O, to represent elements. They use chemical formulas, such as \(\text{H}_2\text{O}\), to represent compounds. They represent chemical reactions using a combination of chemical formulas and symbols. For example, the reaction of hydrogen with oxygen to form water is represented as follows:

\[
2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l)
\]

To represent chemical reactions accurately, chemical equations should be written so that they reflect theories and laws that relate to chemical reactions. One such law is the law of conservation of mass. The law was first developed by French chemist Antoine Lavoisier (1743–1794).

A painting depicting Lavoisier at work is shown in Figure 3.9. One of Lavoisier’s most successful and influential techniques as a scientist was his careful measurement of mass. He stressed the importance of measuring the mass of all the substances involved in a chemical change. These measurements were crucial for making accurate inferences about what happened to the substances. By generalizing many observations of the same results, Lavoisier wrote his version of the law of conservation of mass.

Lavoisier’s law of conservation of mass

During a chemical reaction, the total mass of the reacting substances (the reactants) is always equal to the total mass of the resulting substances (the products).

One of Lavoisier’s innovations was carrying out his reactions in closed systems. A closed system does not allow any exchange of matter between the system and its surroundings. A chemical system includes all the substances involved in a reaction and the container that holds them. The surroundings include everything that is not part of the system. A tightly capped test tube containing a solution is a closed system. Gases, liquids, and solids cannot enter or exit the system. An open system allows the exchange of substances between the system and its surroundings. An open test tube containing a solution is an example of an open system.

The law of conservation of mass is a vital part of modern chemical theory. In Inquiry Investigation 3-B on the following pages, you can carry out an investigation to illustrate the law of conservation of mass.
Comparing the Masses of Reactants and Products

Laws such as the law of conservation of mass are generalizations based on numerous experiments that have given the same results. In this investigation, you will consider why it is important for scientists to analyze the appropriateness of their procedure and the precision of their instruments when interpreting results.

Question
How do you decide whether observations of a specific chemical reaction demonstrate the law of conservation of mass?

Hypothesis
• State the law of conservation of mass and describe and explain the conditions necessary to demonstrate the law.
• Explain how you expect the law of conservation of mass will apply to this investigation.

Prediction
• Predict how the mass of your system will compare before and after the reaction. Remember, you should be able to explain your prediction using your hypothesis.
• Once you have measured the initial mass of your system (Procedure step 4), record your prediction as a specific mass.

Safety Precautions
• Sodium hydroxide solution is caustic. Avoid skin contact.
• Copper(II) sulfate is toxic.
• Always report spills of any chemicals to your teacher immediately.
• If any solutions touch your skin, rinse immediately with plenty of cold water. Report the accident to your teacher.
• Wash your hands thoroughly at the end of this investigation.
• Dispose of materials as your teacher directs.

Apparatus
- electronic balance
- Erlenmeyer flask
- stopper
- small test tube
- tongs

Materials
- copper(II) sulfate solution
- sodium hydroxide solution
Procedure

1. Pour 20 mL of sodium hydroxide solution into the Erlenmeyer flask.

2. Pour copper(II) sulfate solution into the small test tube until the test tube is about half full.

3. Tilt the Erlenmeyer flask to one side, and carefully place the test tube inside. Do not let the solutions mix. Seal the flask with a stopper.

4. Measure the mass of the flask and its contents. Ensure that you record all certain digits and one estimated digit in your reported mass. For electronic balances, the last number of the digital readout is uncertain. Record your prediction for the mass of the system after the reaction. Think carefully about how many significant digits you include in your prediction.

5. Tip the flask to allow the solutions to mix.

6. Measure the mass of the flask and its contents again. Record your measurements. Also record the appearance of the contents of the flask.

7. Remove the small test tube with tongs. Dispose of the contents of the flask as directed by your teacher. Clean up your workspace and wash your hands thoroughly.

Analyze

1. How do you know that a chemical reaction occurred when the solutions were mixed?

2. Subtract the mass of the flask and contents after the reaction from the mass of the flask and contents before the reaction.

3. Indicate the precision (degree of uncertainty) in your measurement. Refer to SkillFocus 4 for help.

4. Consider your mass measurements. Based on the precision you reported, were the initial and final masses different or the same? Can you be certain that this is the case? Explain your answer.

5. Find the percent difference between the final mass that you predicted and the result you obtained empirically. Use the following formula:

   \[ \% \text{ difference} = \left| \frac{\text{final mass of system} - \text{predicted mass}}{\text{predicted mass}} \right| \times 100\% \]

6. Compile a list of the percentage differences observed by all lab groups in your class. If any figures differ greatly from the others, decide as a class whether to include them when calculating an average. Then find the average percentage difference.

7. Identify some possible sources of experimental errors in the class results from question 6.

Conclude and Apply

8. According to your analysis, do the experimental results from your class illustrate the law of conservation of mass? Explain.

9. Evaluate the procedure for this investigation. Suggest some improvements in the apparatus or procedure that would increase your confidence in your results.

Extend Your Knowledge

10. Suppose you performed a reaction in an open system, such as a flask without a stopper.

   (a) If the products included a gas, how would you expect the mass of chemicals and glassware to compare before and after the reaction?

   (b) Why do you think it took chemists many years to realize that the mass of the reactants is always equal to the mass of the products?
Writing Balanced Chemical Equations

The simplest form of a chemical equation is a word equation. For example,

reactants on left side of arrow → hydrogen + oxygen → water → products on right side of arrow

plus sign on left side means "reacts with"
arrow means "produce"

Word equations provide only limited information about a chemical reaction. They do not tell anything about the chemical composition of the reactants and products. They also do not give the numbers of atoms, molecules, or formula units that are involved. You can start to write a more useful equation by replacing words with chemical symbols and formulas:

\[ \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} \]

This "bare-bones" representation of a word equation is called a skeleton equation. A skeleton equation is not a complete or accurate equation, however. It may not show conservation of mass. Use Figure 3.10 to help you understand why the skeleton equation above violates the law of conservation of mass.

Lavoisier's law of conservation of mass demands that the reactants have the same total mass as the products. This is why chemists write balanced chemical equations. According to the atomic theories you studied in Chapter 1, atoms of elements and compounds are not created or destroyed in chemical reactions. Therefore, there should be the same number of atoms of each element on each side of a chemical equation. A balanced chemical equation shows that atoms are conserved in a chemical reaction.

In the skeleton equation above, you can see that the number of hydrogen atoms is balanced. There are two hydrogen atoms on each side. The oxygen atoms are not balanced, though. There are two oxygen atoms on the left, and there is only one oxygen atom on the right. So the skeleton equation is unbalanced. To make it balanced, you add (where necessary) numbers called coefficients before the chemical formulas. Examine the modified chemical equation below. The coefficient 2 in front of the formula for water means that there are two molecules of water:

\[ \text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

Does adding this coefficient balance the equation? No. Count the number of hydrogen atoms on both sides, and you will see why. As shown in Figure 3.11, there are two hydrogen atoms on the left, but two molecules of water would contain four atoms of hydrogen.
To finally balance this chemical equation, you need to add one more coefficient:

$$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$$

Figure 3.12 shows why this equation is balanced. Often it is useful to know the state of the chemicals that are involved in a chemical reaction. Add this information to a chemical equation by inserting the appropriate abbreviation in parentheses after each chemical formula. Table 3.3 provides the abbreviations.

The final balanced equation for the formation of water, if the reactants and product are at room temperature, is:

$$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$$

Notice that you do not include a coefficient 1.

Chemical equations sometimes also show whether a reaction is exothermic or endothermic. If a reaction is endothermic, energy is written as a reactant. If a reaction is exothermic, energy is written as a product. For example, the decomposition of copper(II) oxide into copper and oxygen is endothermic. The reaction can be written like this:

$$\text{CuO}(s) + \text{thermal energy} \rightarrow \text{Cu}(s) + \text{O}_2(\text{g})$$

The exothermic reaction of sulfuric acid and sodium hydroxide to form water and sodium sulfate can be written as follows:

$$\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\ell) + \text{Na}_2\text{SO}_4(\text{aq}) + \text{thermal energy}$$

### Writing Balanced Chemical Equations

Balancing a chemical equation requires patience, practice, and perseverance. Here are some guidelines and examples to help.

1. **Skeleton Equation**: Write the skeleton equation. Be sure you have the correct formulas for all the compounds and elements. Each reactant and product make up a “term” in the equation.

2. **Balancing**: Identify unbalanced atoms and polyatomic ions. Write coefficients to balance them. The steps below will help for more complex equations.
   - Look for an unbalanced atom or a polyatomic ion that appears just once on each side of the equation. Add balancing coefficients to the two terms that contain the element or polyatomic ion. Leave atoms that appear as elements in the equation until later (e.g., \(\text{O}_2\), \(\text{H}_2\), or \(\text{Na}\)).
   - Pick another unbalanced element that appears once on each side. Choose coefficients to balance the atoms of this element as well. Adjust coefficients so that the terms you balanced in the previous step remain balanced. Repeat for any other elements that appear once on each side.
   - Balance any remaining unbalanced elements or ions. Use fractional coefficients if necessary.
   - Clear any fractional coefficients by multiplying the whole equation by the same factor. The resulting coefficients should be the smallest possible whole numbers.

3. **Check It**: Check that the equation is balanced. Make a table of the atoms of each element on each side of the equation.
Examples of Balancing Chemical Equations

Model Problem 1

Write a balanced chemical equation to represent the following word equation.
sodium + water → sodium hydroxide + hydrogen gas

1. Skeleton Equation

First write a skeleton equation. Be sure you have used the correct formulas.

**Na(s) + H$_2$O(l) → NaOH(aq) + H$_2$(g)**

**Remember:** Hydrogen gas is diatomic.

2. Balancing

- Identify unbalanced atoms and polyatomic ions.

  **Na(s) + H$_2$O(l) → NaOH(aq) + H$_2$(g)**

  The only atoms that do not balance on each side of the skeleton equation are hydrogen atoms.

- There are two hydrogen atoms on the reactant side, and three hydrogen atoms on the product side. On the product side, hydrogen appears as an element, H$_2$, and also as part of sodium hydroxide, NaOH.

- Since the atoms of all other elements are balanced, begin to balance the equation by changing the coefficient in front of H$_2$O(l) to 2. There are now four hydrogen atoms on the reactant side. Balance the hydrogen atoms by changing the coefficient in front of NaOH(aq) to 2. The hydrogen atoms are balanced.

  **Na(s) + 2H$_2$O(l) → 2NaOH(aq) + H$_2$(g)**

- The sodium atoms are now unbalanced. There are two sodium atoms on the product side, but only one on the reactant side. Balance the sodium atoms by placing a coefficient of 2 in front of Na(s) on the reactant side.

  **2Na(s) + 2H$_2$O(l) → 2NaOH(aq) + H$_2$(g)**

  The equation now appears to be balanced.

3. Check It

Your final step should always be to check the total number of atoms on each side of the equation. Make a chart to list the reactants and products, and record the number of atoms and ions to see if they match.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 Na atoms</td>
<td>2 Na$^+$ ions</td>
</tr>
<tr>
<td>2 O atoms</td>
<td>2 O atoms</td>
</tr>
<tr>
<td>4 H atoms</td>
<td>4 H atoms</td>
</tr>
</tbody>
</table>

The equation is balanced.

Pause & Reflect

Sometimes students try to balance a chemical equation by changing the formula of a reactant or product. For example, you might be tempted to balance the skeleton equation for water by changing H$_2$O to H$_2$O$_2$. This approach seems simple, and the numbers balance. So why is it incorrect?
Model Problem 2

Write a chemical equation for the following word equation.
copper + silver nitrate $\rightarrow$ copper(II) nitrate + silver

1. **Skeleton Equation**

\[ \text{Cu}(s) + \text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag}(s) \]

2. **Balancing**

On the left and right sides, there are compounds that contain nitrate ions, \( \text{NO}_3^- \). Balance these ions as a unit first.

\[ \text{Cu}(s) + 2\text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag}(s) \]

Then balance the silver atoms.

\[ \text{Cu}(s) + 2\text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + 2\text{Ag}(s) \]

The copper atoms are balanced without any further changes.

3. **Check It**

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 Cu atom</td>
<td>1 Cu(^+) ion</td>
</tr>
<tr>
<td>2 Ag(^+) ions</td>
<td>2 Ag atoms</td>
</tr>
<tr>
<td>2 NO(_3^-) ions</td>
<td>2 NO(_3^-) ions</td>
</tr>
</tbody>
</table>

The equation is balanced.

Figure 3.13 This photo shows the reaction in Model Problem 2. Which ions are in the solution? What is the precipitate that is forming on the wire?

Model Problem 3

Write a chemical equation for the following word equation.
calcium nitrate + sodium hydroxide $\rightarrow$ calcium hydroxide + sodium nitrate

1. **Skeleton Equation**

\[ \text{Ca(NO}_3)_2(aq) + \text{NaOH}(aq) \rightarrow \text{Ca(OH)}_2(s) + \text{NaNO}_3(aq) \]

2. **Balancing**

First balance the nitrate ions and the hydroxide ions.

\[ \text{Ca(NO}_3)_2(aq) + 2\text{NaOH}(aq) \rightarrow \text{Ca(OH)}_2(s) + 2\text{NaNO}_3(aq) \]

The equation appears to be balanced without further changes.

3. **Check It**

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 Ca(^+) ion</td>
<td>1 Ca(^+) ion</td>
</tr>
<tr>
<td>2 NO(_3^-) ions</td>
<td>2 NO(_3^-) ions</td>
</tr>
<tr>
<td>2 Na(^+) ions</td>
<td>2 Na(^+) ions</td>
</tr>
<tr>
<td>2 OH(^-) ions</td>
<td>2 OH(^-) ions</td>
</tr>
</tbody>
</table>

The equation is balanced.
In 1995 a scientist spent 15 days in a sealed chamber. His oxygen was supplied entirely by photosynthesis, courtesy of 30,000 wheat plants. In turn, the carbon dioxide needed for the photosynthesis came from the scientist's breath, as a product of his cellular respiration. Human-carrying interplanetary expeditions will likely depend on similar closed systems to supply and recycle water, oxygen, and many nutrients.

5. Balance each of the following equations.
(a) \( \text{Li}_2(\text{s}) + \text{H}_2\text{O}(\ell) \rightarrow \text{LiOH}(\text{aq}) + \text{H}_2(\text{g}) \)
(b) \( \text{Fe}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s}) \)
(c) \( \text{Zn}(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g}) \)
(d) \( (\text{NH}_4)_3\text{PO}_4(\text{aq}) + \text{Ba(OH)}_2(\text{aq}) \rightarrow \text{Ba}_3(\text{PO}_4)_2(\text{s}) + \text{NH}_4\text{OH}(\text{aq}) \)
(e) \( \text{Pb(NO}_3)_2(\text{aq}) + \text{KCl}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) + \text{KNO}_3(\text{aq}) \)
(f) \( \text{Cu(NO}_3)_2(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{NO}_2(\text{g}) + \text{O}_2(\text{g}) \)

6. Aqueous copper(II) nitrate reacts with aqueous potassium hydroxide to form aqueous potassium nitrate and solid copper(II) hydroxide. Write a balanced chemical equation to represent this reaction.

7. In section 3.1, you learned that plants use the Sun's energy to make their own food through the process of photosynthesis. Both plants and animals use food to make energy through the process of cellular respiration.
(a) The skeleton equation for photosynthesis is below. Balance the equation. **Hint:** Leave oxygen for last.
\( \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{C}_6\text{H}_12\text{O}_6(\text{aq}) + \text{O}_2(\text{g}) \)
(b) The skeleton equation for cellular respiration is below. Balance the equation.
\( \text{C}_6\text{H}_12\text{O}_6(\text{aq}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell) \)
(c) Classify each reaction as exothermic or endothermic. For each equation, add "+ energy" on the product side or reactant side.

8. Figure 3.14 below shows the contact process for producing sulfuric acid, an important and versatile industrial chemical.
(a) Read about the reactions that take place in each step. For each of the four steps, write a chemical equation to represent the reaction.
(b) Identify three dangerous substances that are used or produced in the contact process. Explain why they are dangerous.
(c) The contact process produces no harmful waste substances that must be disposed of. Explain what happens to the dangerous substances you identified in question (b).

**Figure 3.14**
Step 1 Liquid sulfur, \( \text{S}(\ell) \), is burned in air. The sulfur reacts with oxygen to produce sulfur dioxide, \( \text{SO}_2(\text{g}) \), a stable, poisonous gas. The sulfur can be obtained from underground deposits or as a byproduct of natural gas production.
Step 2 After impurities are removed, the sulfur dioxide is reacted with more oxygen to produce toxic sulfur trioxide gas, \( \text{SO}_3(\text{g}) \). To speed up the reaction, the gases are heated to a temperature of 400°C.
Step 3 After cooling, the sulfur trioxide is bubbled through highly concentrated sulfuric acid, \( \text{H}_2\text{SO}_4(\text{aq}) \). The reaction produces pyrosulfuric acid, \( \text{H}_2\text{SO}_7(\text{aq}) \).
Step 4 Water is added to the pyrosulfuric acid, producing highly concentrated sulfuric acid, \( \text{H}_2\text{SO}_4(\text{aq}) \).
Section 3.2 Summary

In this section, you learned that the law of conservation of mass states that the mass of a system remains constant during a chemical reaction. You learned how to write balanced chemical equations to reflect this law. In Section 3.3, you will continue to work on your balancing skills as you learn to classify chemical reactions.

Check Your Understanding

1. Create a chart that summarizes the key features of word equations, skeleton equations, and balanced equations, showing the advantages and disadvantages of each type of equation.

2. Explain why a balanced chemical equation is consistent with the law of conservation of mass. Use an example of a balanced chemical equation in your answer.

3. Each of the following chemical equations is balanced, but is incorrect in some other way. State what is wrong and then write the equation correctly.
   (a) \( H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g) \)
   (b) \( NH_3(g) \rightarrow N(g) + 3H(g) \)
   (c) \( 2C(s) + 2O_2(g) \rightarrow 2CO_2(g) \)

4. When you light a campfire, you are burning carbon compounds in the wood. The main compound in the wood is cellulose, a compound containing carbon, hydrogen, and oxygen. The products are carbon dioxide and water vapour. There is one other reactant besides cellulose. What is it? Write a word equation to describe the burning of cellulose.

5. Apply Given the following word and skeleton equations, write balanced chemical equations for each reaction.
   (a) copper(II) oxide \( \rightarrow \) copper + oxygen gas
      \( CuO(s) \rightarrow Cu(s) + O_2(g) \)
   (b) methane + oxygen \( \rightarrow \) carbon dioxide + water
      \( CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g) \)
   (c) ammonia gas + oxygen \( \rightarrow \) nitrogen + water
      \( NH_3(g) + O_2(g) \rightarrow N_2(g) + H_2O(l) \)
   (d) sulfur dioxide + oxygen \( \rightarrow \) sulfur trioxide
      \( SO_2(g) + O_2(g) \rightarrow SO_3(g) \)

6. Thinking Critically Suppose that you measure the mass of a chemical in an open container, and then heat it for a few minutes over a Bunsen burner flame. After the container and contents have cooled, you find that the mass is larger than before.
   (a) If you accept the law of conservation of mass, how can you explain your observation?
   (b) Is there a way to carry out the same reaction so that your results illustrate the law of conservation of mass? Explain your answer.