Lesson 9.1 Chemical Reactions and Equations

Chemical Reactions

An example of a **chemical reaction** is the burning of hydrogen gas (H_2) to produce steam (hot H_2O gas). In chemistry, to **burn** something is to react it with oxygen gas (O_2) to form one or more new substances.

Chemical equations are the language used to describe chemical reactions. In a chemical equation written using *molecular* formulas, the above reaction would be represented as

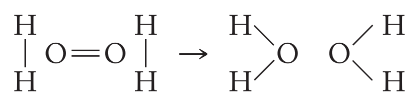
$$2\ H_2 + O_2 \rightarrow 2\ H_2O \qquad (9.1)$$

This equation can be read as either "two H two plus one O two react to form two H two O," or as "two molecules of hydrogen plus one molecule of oxygen react to form two molecules of water."

The substances on the left side of a reaction equation are termed the **reactants**. The substances on the right side of the arrow are the **products**.

> In chemical reactions, reactants are *used up* and products *form*.

Most chemical reactions are represented by equations using molecular formulas, as in equation 9.1 above. However, more information is supplied if the equation is written using structural formulas. An example is



(9.2)

TRY IT

(See "How to Use These Lessons," point 1, at the front of the book.)

Q. Compare equation 9.1 to equation 9.2. Are they the same reaction?

STOP

Answer:

Yes.

However, by writing the structural formulas it is easier to see that in many respects, after the reaction, not much has changed. We began with four hydrogen atoms and two oxygen atoms; we end with the same.

> In chemical reactions, bonds break, and new bonds form, but the number and kinds of atoms stay the same.

The fact that chemical reactions can neither create nor destroy atoms is called the **law of conservation of atoms** or the **law of conservation of matter**. In this usage, *conservation* means that what you start with is conserved at the end.

Before, during, and after a reaction, there is also **conservation of mass**: the total mass of the reactants and products also does not change. Total mass is determined by the number and kind of atoms, which a chemical reaction does not change.

What does change? Because of the new positions of the bonds, after the reaction the products will have characteristics and behavior that are different from those of the reactants.

In the above reaction, the molecules on the left are explosive when ignited, but the water molecules on the right are quite stable. The oxygen molecules on the left cause many materials to burn. To stop burning, we often use the water on the right.

The position of the bonds can also change the *economic value* of atoms. The historic importance of chemistry to society has included the discovery of reactions that change

- Brittle rock into metals that can be molded and shaped
- Willow bark into aspirin; and fungus into antibiotics
- Animal waste into explosives; and sand into computer chips for electronic devices

Another outcome of chemical reactions is quite often the storage or release of *energy*. In burning hydrogen to form water, large amounts of stored energy are released. Including the energy term, the burning of hydrogen is represented as

$$2\ H_2 + O_2 \rightarrow 2\ H_2O + \text{energy}$$

It was the release of energy from this simple reaction that led to the explosive destruction of the airship *Hindenburg* in 1937. Currently, researchers are seeking ways to harness the energy released by burning hydrogen as an alternative to burning gasoline in cars.

Nearly all reactions either absorb or release energy; however, if energy is not the focus of a particular problem, the energy term is often omitted when writing a reaction equation.

Reaction Equation Terminology

1. In the substance formula $\mathbf{H_2O}$, the 2 is a *subscript*. The omission of a subscript, such as after the O in H_2O, means the subscript is understood to be 1.

2. In a reaction equation, if $\mathbf{5\ H_2O}$ is a term, the **5** is a called a **coefficient**. Coefficients are exact numbers that express the exact *particle ratios* in a reaction.

   It is important to distinguish between subscripts and coefficients.

   - Subscripts are numbers written *after* and *lower than* the atom symbols in a molecule or ion formula. Subscripts count the atoms of each type inside the particle.
   - Coefficients are numbers written in *front* of a particle formula. Coefficients represent the particle ratios in a reaction: a count of how many of one particle reacts with how many of another particle.

3. In a reaction equation involving substances, if the number and kind of atoms on each side of the arrow is the same, the equation is said to be **balanced**. The coefficients of a *balanced* equation show the exact ratios in which the particles react (are used up) and are formed.

In a balanced equation, writing a coefficient of 1 is optional. If no coefficient is written in front of an equation term, the coefficient is understood to be 1.

4. To balance equations, we will need to *count atoms* based on coefficients and subscripts.

To count each kind of atom in a term in a reaction equation, multiply the coefficient times the subscript(s) for the atom.

Example: The term **5 H_2O** represents five molecules of water. Each molecule has three atoms. In those five molecules are $(5 \times 2) =$ **10 hydrogen** atoms *and* $(5 \times 1) =$ **5 oxygen** atoms.

TRY IT

Q. Count the **H** atoms in the following:

a. **5 CH_4**

b. **3 CH_3COOH**

c. **2 $Pb(C_2H_5)_4$**

STOP

Answer:

a. Each molecule has four H atoms. Five molecules $= 5 \times 4$ H $=$ **20 H** atoms

b. Each molecule has four H atoms. Three molecules $= 3 \times 4$ H $=$ **12 H** atoms

c. The H in this case has *two* subscripts. Multiply $2 \times 4 \times 5 =$ **40 H** atoms

PRACTICE

If you are not sure that your answer is correct, check it before proceeding to the next question.

1. Label the *reactants* and *products* in this reaction equation. Circle the coefficients.

$$4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$$

2. How many *oxygen* atoms are represented in the following?

a. 7 Na_3PO_4

b. 3 $Co(OH)_2$

c. 2 $No(NO_3)_2$

d. 5 $Al_2(SO_4)_3$

3. How many *total* atoms are represented in 2a and 2d above?

4. The following equation uses structural rather than molecular formulas.

$$\mathrm{C + C + H{-}H + H{-}H + H{-}H \rightarrow} \begin{array}{ccccccc} & & \mathrm{H} & & \mathrm{H} & & \\ & & | & & | & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{H} \\ & & | & & | & & \\ & & \mathrm{H} & & \mathrm{H} & & \end{array}$$

a. Is the equation balanced?

b. Write the reaction using *molecular* formulas (the type used in problem 1).

c. In going from reactants to products, what changed? What stayed the same?

Changes:

Stays the same:

Lesson 9.2 Balancing Equations

Balancing by Trial and Error

The coefficients that balance an equation are not always supplied with the equation. However, if you are given the chemical formulas for the reactants and products of a reaction, you can determine the coefficients that balance the equation by *trial and error*.

One consequence of the law of conservation of atoms is that *only one set of ratios* will balance a chemical equation. However, because coefficients are ratios, if you multiply all of the coefficients by the same number, you continue to have a balanced equation. This means that a balanced equation may be shown with different sets of coefficients, as long as the ratios among the coefficients are the same.

Example: $2\ H_2 + O_2 \rightarrow 2\ H_2O$ *and* $H_2 + 1/2\ O_2 \rightarrow H_2O$ are the same balanced equation, because the coefficient *ratios* are the same. Note that if no coefficient is shown, a 1 is understood.

In a balanced equation, showing a coefficient that is *one* is not required, but it's not improper, either.

There are many ways to balance equations. In later lessons, we will learn methods that balance specific types of equations more quickly. However, trial-and-error balancing uses the same rules for all types of equations, and with perseverance works for all types of equations, so it's a good place to start.

Let's learn to balance with an example. Using the question below, apply the rules and steps below the question, then check your answers.

TRY IT

Q. Add the coefficients that balance this equation for the burning of *n*-propanol.

$$C_3H_7OH + \quad O_2 \rightarrow \quad CO_2 + \quad H_2O$$

Steps in balancing:

1. Most important: During balancing, you cannot change a formula or a subscript. To balance, *you* must add numbers to the equation, but the *only* numbers that you can add are the *coefficients* that go in *front* of substance formulas.

2. To start, put a coefficient of 1 in front of the *most complex* formula (the one with the most atoms or the most different kinds of atoms). If two formulas seem complex, choose either one.

   Writing coefficients that are 1 is optional, but in an equation that you need to balance, including the 1s helps in tracking which coefficients have been determined.

STOP

Answer:

In this case, the first formula is the most complex, so start with

$$\mathbf{1}\ C_3H_7OH + \quad O_2 \rightarrow \quad CO_2 + \quad H_2O$$

3. Now add coefficients to the other side that must be true if atoms are balanced.

STOP

   In chemical reactions, atoms cannot be created or destroyed. This means that in chemical equations, each side of the arrow must have the same kinds of atoms, and the same number of each kind of atom.

> Each **term** in a chemical equation is a coefficient followed by a substance formula.
>
> For a given atom, to count the number of atoms represented by a term, multiply the coefficient by the subscript(s) for that atom.
>
> To count each type of atom on a *side*, add the atoms in each term on that side.

   Above, the left side has three carbon atoms. Because only CO_2 on the right has C, the only way to have three carbon atoms on the right is to have the CO_2 coefficient be 3. The left has eight hydrogen atoms total; because only H_2O on the right has H, the only way to have eight H atoms on the right is to have the H_2O coefficient be 4. So far, this gives us

$$1\ C_3H_7OH + \underline{\quad}\ O_2 \rightarrow \mathbf{3}\ CO_2 + \mathbf{4}\ H_2O$$

   The right side is now finished, because each particle has a coefficient. Only the O_2 on the left side lacks a coefficient.

4. Add the coefficient that must be true for the oxygen to balance.

STOP

   We count the oxygens on the right side and get 10. On the left, we see one oxygen in propanol, which means we must have a total of **nine** oxygens from O_2. We can write

$$1\ C_3H_7OH + \mathbf{9/2}\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

or we can multiply *all* of the coefficients by 2 to avoid using fractions.

$$\mathbf{2}\,C_3H_7OH + \mathbf{9}\,O_2 \rightarrow \mathbf{6}\,CO_2 + \mathbf{8}\,H_2O$$

Because coefficients are ratios, and both sets of ratios above are the same, both answers above are *equally correct*. We can multiply all the coefficients by the same number and still have the same ratios *and* a balanced equation.

Balancing Using Fractions as Coefficients

Initially, our primary use for coefficients will be as *ratios* in calculating amounts of substances involved in chemical reactions. The arithmetic in these calculations will be easier if the coefficients are converted to *whole numbers* at the end of balancing. Any whole numbers with proper ratios would be correct, but *lowest* whole-number ratios are preferred.

Fractions are permitted when adding coefficients to balance equations, and in some types of problems, including some energy calculations, the use of fractions to balance equations is required. In other situations, fractions may be inappropriate (you cannot have half a molecule). We will address these differences as we encounter them.

In a few problems below, we will include balancing with fractions to be familiar with their use. However, in *most* upcoming cases, if we encounter a fraction when balancing, it will simplify the arithmetic if at that point all of the coefficients are multiplied by the fraction's denominator. This will change coefficients that are fractions to lowest whole numbers.

Our rule will be

> Unless other coefficients are specified or required in a problem, convert to lowest-whole-number coefficients at the final step in balancing equations.

PRACTICE

Read each numbered step below, then do every other or every third lettered problem. As you go, check your answers. If you need more practice at a step, do a few more letters for that step. Save a few for your next practice session.

1. A balanced equation must have the same number of each kind of atom on both sides.

   To check for a balanced equation, count the total number of one kind of atom on one side, then count the number of that kind of atom on the other side. The left- and right-side counts must be equal.

   Repeat those counts for each kind of atom in the equation.

   Using this counting method, label each equation below as *balanced* or *unbalanced*.

   a. $2\,Cs + Cl_2 \rightarrow 2\,CsCl$

   b. $4\,HI + O_2 \rightarrow 2\,H_2O + I_2$

   c. $Pb(NO_3)_2 + 2\,LiBr \rightarrow PbBr_2 + LiNO_3$

   d. $BaCO_3 + 2\,NaCl \rightarrow Na_2CO_3 + BaCl_2$

(continued)

2. In the equations below, one coefficient has been supplied. Use that coefficient to decide one or more coefficients on the *other* side. Then use your added coefficient(s) to go back and forth, from side to side, filling the remaining blanks on both sides to balance the equation.

   In these, some coefficients may be fractions. Fractions are sometimes needed when balancing.

   Remember that balancing is trial and error. Do what works. If you need help, check your answer after each letter.

   Tip: It helps to balance *last* an atom that is used in two or more different formulas on the same side. Oxygen is the atom most frequently encountered in compounds, so "saving O until last" usually helps in balancing.

   a. $4\ Al + \_\_\_\_\ O_2 \rightarrow \_\_\_\_\ Al_2O_3$

   b. $3\ Ca + \_\_\_\_\ N_2 \rightarrow \_\_\_\_\ Ca_3N_2$

   c. $\_\_\_\_\ P_4 + \_\_\_\_\ O_2 \rightarrow 2\ P_4O_6$

   d. $\_\_\_\_\ C_6H_{14} + \_\_\_\_\ O_2 \rightarrow 18\ CO_2 + \_\_\_\_\ H_2O$

   e. $\_\_\_\_\ MgH_2 + \_\_\_\_\ H_2O \rightarrow 1\ Mg(OH)_2 + \_\_\_\_\ H_2$

   f. $\_\_\_\_\ C_2H_6 + 7\ O_2 \rightarrow \_\_\_\_\ CO_2 + \_\_\_\_\ H_2O$

3. Balance these equations. Start by placing a coefficient of **1** in front of the underlined substance.

   a. $K + \underline{F_2} \rightarrow KF$

   b. $Cs + \underline{O_2} \rightarrow Cs_2O$

   c. $PCl_3 \rightarrow \underline{P_4} + Cl_2$

   d. $\underline{C_2H_5OH} + O_2 \rightarrow CO_2 + H_2O$

   e. $FeS + O_2 \rightarrow \underline{Fe_2O_3} + SO_2$

4. When balancing equations without suggested ways to start,

   > *Begin* by putting a **1** in front of the most complex formula on *either* the left or right side of the equation (the one with the most atoms, or the most different kinds of atoms). If two formulas are complex, use either one.

   Then add as many coefficients as you are sure of to the side *opposite* the side where you put the **1**.

   On this problem, if you get a fraction as you balance, multiply *all* of the existing coefficients by the denominator of the fraction. Repeat this step if you get additional fractions while balancing. Fractions are permitted, but it will be easier in most calculations if you have whole-number coefficients.

Expect to need to start over on occasion, because balancing is trial and error. Be persistent! All of the equations below can be balanced.

a. $Mg + O_2 \rightarrow MgO$

b. $N_2 + O_2 \rightarrow NO$

c. $C_6H_6 + O_2 \rightarrow CO_2 + H_2O$

d. $P_4 + O_2 \rightarrow P_4O_6$

e. $Al + HBr \rightarrow AlBr_3 + H_2$

5. *Lowest*-whole-number coefficients are not required, but they are preferred when writing most balanced equations.

   Example: If all of your coefficients are *even* numbers at the end of balancing, it is preferred to divide all the coefficients by 2.

   This converts $4\ H_2 + 2\ O_2 \rightarrow 4\ H_2O$ to $2\ H_2 + O_2 \rightarrow 2\ H_2O$

   The latter equation has *lowest-whole-number* coefficients.

   On these, balance, then convert your answers to lowest-whole-number ratios.

a. $Al_2O_3 + HCl \rightarrow AlCl_3 + H_2O$

b. $Fe_3O_4 + H_2 \rightarrow Fe + H_2O$

c. $C + SiO_2 \rightarrow CO + SiC$

d. $N_2 + O_2 + H_2O \rightarrow HNO_3$

e. $Rb_2O + H_2O \rightarrow RbOH$

f. $Mg(NO_3)_2 + Na_3PO_4 \rightarrow Mg_3(PO_4)_2 + NaNO_3$

g. $Pb(C_2H_5)_4 + O_2 \rightarrow PbO + CO_2 + H_2O$

h. $Cd + HNO_3 \rightarrow Cd(NO_3)_2 + H_2O + NO$

6. Working in your notebook, write the formulas, then balance these. (Need formula help? See Lessons 7.2 and 7.3.)

a. Dinitrogen tetroxide → nitrogen dioxide

b. Barium carbonate + cesium chloride → cesium carbonate + barium chloride

c. Silver nitrate + calcium iodide → silver iodide + calcium nitrate