31

Stoichiometry

Relationships Stoichiometry

Chemical equations are symbolic representations of chemical and physical changes. Formulas for the substances undergoing the change (reactants) and substances generated by the change (products) are separated by an arrow and preceded by integer coefficients indicating their relative numbers. Balanced equations are those whose coefficients result in equal numbers of atoms for each element in the reactants and products. Chemical reactions in aqueous solution that involve ionic reactants or products may be represented more realistically by complete ionic equations and, more succinctly, by net ionic equations. A balanced chemical equation may be used to describe a reaction's stoichiometry (the relationships between amounts of reactants and products). Coefficients from the equation are used to derive stoichiometric factors that subsequently may be used for computations relating reactant and product masses, molar amounts, and other quantitative properties.

31.1 Writing and Balancing Chemical Equations

Learning Objectives

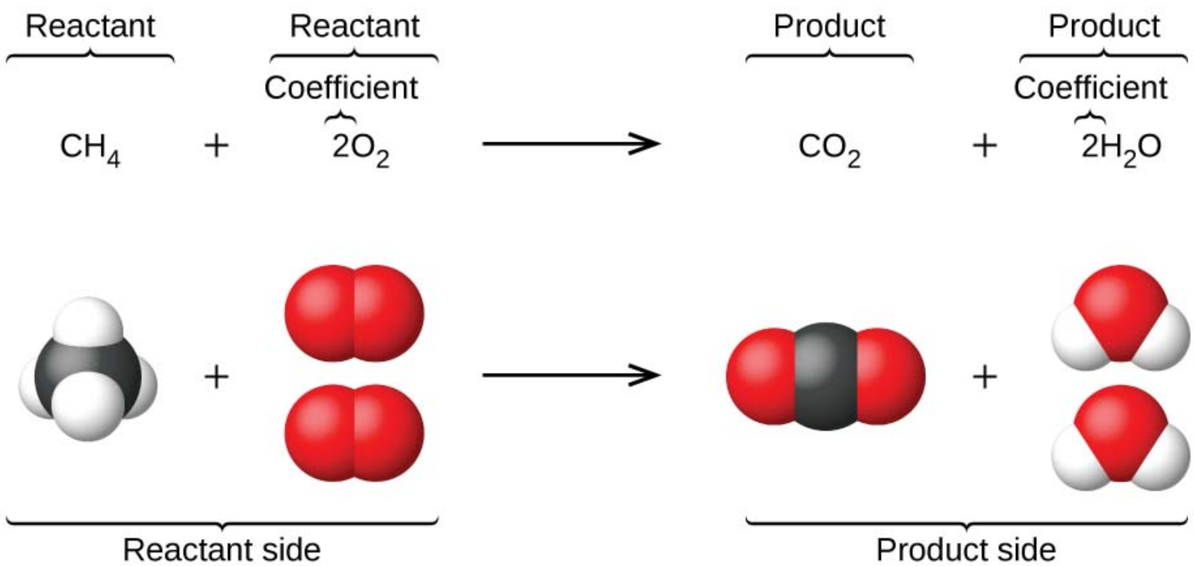
By the end of this section, you will be able to:

- Derive chemical equations from narrative descriptions of chemical reactions.
- Write and balance chemical equations in molecular, total ionic, and net ionic formats.

An earlier chapter of this text introduced the use of element symbols to represent individual atoms. When atoms gain or lose electrons to yield ions, or combine with other atoms to form molecules, their symbols are modified or combined to generate chemical formulas that appropriately represent these species. Extending this symbolism to represent both the identities and the relative quantities of substances undergoing a chemical (or physical) change involves writing and balancing a chemical equation. Consider as an example the reaction between one methane molecule (CH_4) and two diatomic oxygen molecules (O_2) to produce one carbon dioxide molecule (CO_2) and two water molecules (H_2O). The chemical equation representing this process is provided in the upper half of Figure 31.1, with space-filling molecular models shown in the lower half of the figure.

Figure 31.1

The reaction between methane and oxygen to yield carbon dioxide and water (shown at bottom) may be represented by a chemical equation using formulas (top).



This example illustrates the fundamental aspects of any chemical equation:

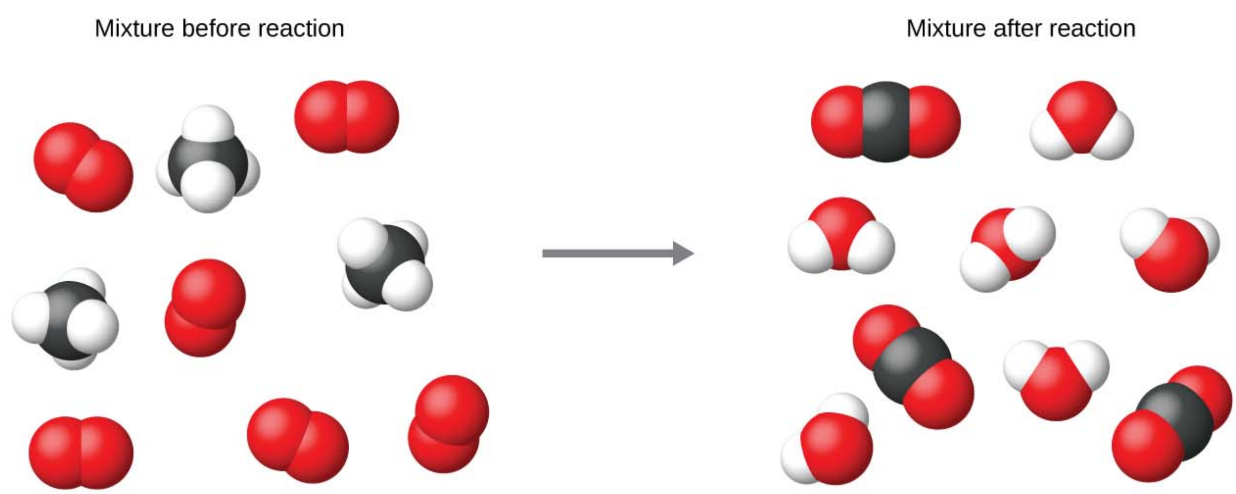
1. The substances undergoing reaction are called reactants, and their formulas are placed on the left side of the equation.
2. The substances generated by the reaction are called products, and their formulas are placed on the right side of the equation.
3. Plus signs (+) separate individual reactant and product formulas, and an arrow $(\longrightarrow)$ separates the reactant and product (left and right) sides of the equation.
4. The relative numbers of reactant and product species are represented by coefficients (numbers placed immediately to the left of each formula). A coefficient of 1 is typically omitted.

It is common practice to use the smallest possible whole-number coefficients in a chemical equation, as is done in this example. Realize, however, that these coefficients represent the *relative* numbers of reactants and products, and, therefore, they may be correctly interpreted as ratios. Methane and oxygen react to yield carbon dioxide and water in a 1:2:1:2 ratio. This ratio is satisfied if the numbers of these molecules are, respectively, 1-2-1-2, or 2-4-2-4, or 3-6-3-6, and so on (Figure 31.2). Likewise, these coefficients may be interpreted with regard to any amount (number) unit, and so this equation may be correctly read in many ways, including:

- *One* methane molecule and *two* oxygen molecules react to yield *one* carbon dioxide molecule and *two* water molecules.
- *One dozen* methane molecules and *two dozen* oxygen molecules react to yield *one dozen* carbon dioxide molecules and *two dozen* water molecules.
- *One mole* of methane molecules and *2 moles* of oxygen molecules react to yield *1 mole* of carbon dioxide molecules and *2 moles* of water molecules.

Figure 31.2

Regardless of the absolute numbers of molecules involved, the ratios between numbers of molecules of each species that react (the reactants) and molecules of each species that form (the products) are the same and are given by the chemical reaction equation.



Balancing Equations

The chemical equation described in section 4.1 is balanced, meaning that equal numbers of atoms for each element involved in the reaction are represented on the reactant and product sides. This is a requirement the equation must satisfy to be consistent with the law of conservation of matter. It may be confirmed by simply summing the numbers of atoms on either side of the arrow and comparing these sums to ensure they are equal. Note that the number of atoms for a given element is calculated by multiplying the coefficient of any formula containing that element by the element's subscript in the formula. If an element appears in more than one formula on a given side of the equation, the number of atoms represented in each must be computed and then added together. For example, both product species in the example reaction, CO_2 and H_2O, contain the element oxygen, and so the number of oxygen atoms on the product side of the equation is

$$\left(1\ \mathrm{CO_2\ molecule} \times \frac{2\ \mathrm{O\ atoms}}{\mathrm{CO_2\ molecule}}\right) + \left(2\ \mathrm{H_2O\ molecules} \times \frac{1\ \mathrm{O\ atom}}{\mathrm{H_2O\ molecule}}\right) = 4\ \mathrm{O\ atoms}$$

The equation for the reaction between methane and oxygen to yield carbon dioxide and water is confirmed to be balanced per this approach, as shown here:

$$\mathrm{CH_4} + 2\mathrm{O_2} \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O}$$

Element	**Reactants**	**Products**	**Balanced?**
C	1 × 1 = 1	1 × 1 = 1	1 = 1, yes
H	4 × 1 = 4	2 × 2 = 4	4 = 4, yes
O	2 × 2 = 4	(1 × 2) + (2 × 1) = 4	4 = 4, yes

A balanced chemical equation often may be derived from a qualitative description of some chemical reaction by a fairly simple approach known as balancing by inspection. Consider as an example the decomposition of water to yield molecular hydrogen and oxygen. This process is represented qualitatively by an *unbalanced* chemical equation:

$$\mathrm{H_2O} \longrightarrow \mathrm{H_2} + \mathrm{O_2} \qquad \text{(unbalanced)}$$

Comparing the number of H and O atoms on either side of this equation confirms its imbalance:

Element	Reactants	Products	Balanced?
H	$1 \times 2 = 2$	$1 \times 2 = 2$	2 = 2, yes
O	$1 \times 1 = 1$	$1 \times 2 = 2$	1 ≠ 2, no

The numbers of H atoms on the reactant and product sides of the equation are equal, but the numbers of O atoms are not. To achieve balance, the *coefficients* of the equation may be changed as needed. Keep in mind, of course, that the *formula subscripts* define, in part, the identity of the substance, and so these cannot be changed without altering the qualitative meaning of the equation. For example, changing the reactant formula from H_2O to H_2O_2 would yield balance in the number of atoms, but doing so also changes the reactant's identity (it's now hydrogen peroxide and not water). The O atom balance may be achieved by changing the coefficient for H_2O to 2.

$$2H_2O \longrightarrow H_2 + O_2 \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
H	$\mathbf{2} \times 2 = 4$	$1 \times 2 = 2$	4 ≠ 2, no
O	$2 \times 1 = 2$	$1 \times 2 = 2$	2 = 2, yes

The H atom balance was upset by this change, but it is easily reestablished by changing the coefficient for the H_2 product to 2.

$$2H_2O \longrightarrow 2H_2 + O_2 \qquad \text{(balanced)}$$

Element	Reactants	Products	Balanced?
H	$2 \times 2 = 4$	$\mathbf{2} \times 2 = 4$	4 = 4, yes
O	$2 \times 1 = 2$	$1 \times 2 = 2$	2 = 2, yes

These coefficients yield equal numbers of both H and O atoms on the reactant and product sides, and the balanced equation is, therefore:

$$2H_2O \longrightarrow 2H_2 + O_2$$

Example 31.1

Balancing Chemical Equations

Write a balanced equation for the reaction of molecular nitrogen (N_2) and oxygen (O_2) to form dinitrogen pentoxide.

Solution

First, write the unbalanced equation.

$$N_2 + O_2 \longrightarrow N_2O_5 \qquad \text{(unbalanced)}$$

Next, count the number of each type of atom present in the unbalanced equation.

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	1 × 2 = 2	2 = 2, yes
O	1 × 2 = 2	1 × 5 = 5	2 ≠ 5, no

Though nitrogen is balanced, changes in coefficients are needed to balance the number of oxygen atoms. To balance the number of oxygen atoms, a reasonable first attempt would be to change the coefficients for the O_2 and N_2O_5 to integers that will yield 10 O atoms (the least common multiple for the O atom subscripts in these two formulas).

$$N_2 + 5O_2 \longrightarrow 2N_2O_5 \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	**2** × 2 = 4	2 ≠ 4, no
O	**5** × 2 = 10	**2** × 5 = 10	10 = 10, yes

The N atom balance has been upset by this change; it is restored by changing the coefficient for the reactant N_2 to 2.

$$2N_2 + 5O_2 \longrightarrow 2N_2O_5$$

Element	Reactants	Products	Balanced?
N	**2** × 2 = 4	2 × 2 = 4	4 = 4, yes
O	5 × 2 = 10	2 × 5 = 10	10 = 10, yes

The numbers of N and O atoms on either side of the equation are now equal, and so the equation is balanced.

Check Your Learning

Write a balanced equation for the decomposition of ammonium nitrate to form molecular nitrogen, molecular oxygen, and water. (Hint: Balance oxygen last, since it is present in more than one molecule on the right side of the equation.)

Answer

$$2NH_4NO_3 \longrightarrow 2N_2 + O_2 + 4H_2O$$

It is sometimes convenient to use fractions instead of integers as intermediate coefficients in the process of balancing a chemical equation. When balance is achieved, all the equation's coefficients may then be multiplied by a whole number to convert the fractional coefficients to integers without upsetting the atom balance. For example, consider the reaction of ethane (C_2H_6) with oxygen to yield H_2O and CO_2, represented by the unbalanced equation:

$$C_2H_6 + O_2 \longrightarrow H_2O + CO_2 \qquad \text{(unbalanced)}$$

Following the usual inspection approach, one might first balance C and H atoms by changing the coefficients for the two product species, as shown:

$$C_2H_6 + O_2 \longrightarrow 3H_2O + 2CO_2 \qquad \text{(unbalanced)}$$

This results in seven O atoms on the product side of the equation, an odd number—no integer coefficient can be used with the O_2 reactant to yield an odd number, so a fractional coefficient,

$\frac{7}{2}$,

is used instead to yield a provisional balanced equation:

$$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 3H_2O + 2CO_2$$

A conventional balanced equation with integer-only coefficients is derived by multiplying each coefficient by 2:

$$2C_2H_6 + 7O_2 \longrightarrow 6H_2O + 4CO_2$$

Finally with regard to balanced equations, recall that convention dictates use of the *smallest whole-number coefficients*. Although the equation for the reaction between molecular nitrogen and molecular hydrogen to produce ammonia is, indeed, balanced,

$$3N_2 + 9H_2 \longrightarrow 6NH_3$$

the coefficients are not the smallest possible integers representing the relative numbers of reactant and product molecules. Dividing each coefficient by the greatest common factor, 3, gives the preferred equation:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

Link to Learning

Use this interactive tutorial for additional practice balancing equations.

Additional Information in Chemical Equations

The physical states of reactants and products in chemical equations very often are indicated with a parenthetical abbreviation following the formulas. Common abbreviations include *s* for solids, *l* for liquids, *g* for gases, and *aq* for substances dissolved in water (*aqueous solutions*, as introduced in the preceding chapter). These notations are illustrated in the example equation here:

$$2\text{Na}(s) + 2\text{H}_2\text{O}(l) \longrightarrow 2\text{NaOH}(aq) + \text{H}_2(g)$$

This equation represents the reaction that takes place when sodium metal is placed in water. The solid sodium reacts with liquid water to produce molecular hydrogen gas and the ionic compound sodium hydroxide (a solid in pure form, but readily dissolved in water).

Special conditions necessary for a reaction are sometimes designated by writing a word or symbol above or below the equation's arrow. For example, a reaction carried out by heating may be indicated by the uppercase Greek letter delta (Δ) over the arrow.

$$\text{CaCO}_3(s) \xrightarrow{\Delta} \text{CaO}(s) + \text{CO}_2(g)$$

Other examples of these special conditions will be encountered in more depth in later chapters.

Equations for Ionic Reactions

Given the abundance of water on earth, it stands to reason that a great many chemical reactions take place in aqueous media. When ions are involved in these reactions, the chemical equations may be written with various levels of detail appropriate to their intended use. To illustrate this, consider a reaction between ionic compounds taking place in an aqueous solution. When aqueous solutions of $CaCl_2$ and $AgNO_3$ are mixed, a reaction takes place producing aqueous $Ca(NO_3)_2$ and solid AgCl:

$$\mathrm{CaCl_2}\,(aq) + \mathrm{2AgNO_3}\,(aq) \longrightarrow \mathrm{Ca(NO_3)_2}\,(aq) + \mathrm{2AgCl}(s)$$

This balanced equation, derived in the usual fashion, is called a molecular equation because it doesn't explicitly represent the ionic species that are present in solution. When ionic compounds dissolve in water, they may *dissociate* into their constituent ions, which are subsequently dispersed homogenously throughout the resulting solution (a thorough discussion of this important process is provided in the chapter on solutions). Ionic compounds dissolved in water are, therefore, more realistically represented as dissociated ions, in this case:

$$\begin{aligned} \mathrm{CaCl_2}\,(aq) &\longrightarrow \mathrm{Ca^{2+}}\,(aq) + \mathrm{2Cl^-}\,(aq) \\ \mathrm{2AgNO_3}\,(aq) &\longrightarrow \mathrm{2Ag^+}\,(aq) + \mathrm{2NO_3{}^-}\,(aq) \\ \mathrm{Ca(NO_3)_2}\,(aq) &\longrightarrow \mathrm{Ca^{2+}}\,(aq) + \mathrm{2NO_3{}^-}\,(aq) \end{aligned}$$

Unlike these three ionic compounds, AgCl does not dissolve in water to a significant extent, as signified by its physical state notation, *s*.

Explicitly representing all dissolved ions results in a complete ionic equation. In this particular case, the formulas for the dissolved ionic compounds are replaced by formulas for their dissociated ions:

$$\mathrm{Ca^{2+}}\,(aq) + \mathrm{2Cl^-}\,(aq) + \mathrm{2Ag^+}\,(aq) + \mathrm{2NO_3{}^-}\,(aq) \longrightarrow \mathrm{Ca^{2+}}\,(aq) + \mathrm{2NO_3{}^-}\,(aq) + \mathrm{2AgCl}(s)$$

Examining this equation shows that two chemical species are present in identical form on both sides of the arrow, Ca^{2+} (*aq*) and

$$\mathrm{NO_3{}^-}\,(aq)\,.$$

These spectator ions—ions whose presence is required to maintain charge neutrality—are neither chemically nor physically changed by the process, and so they may be eliminated from the equation to yield a more succinct representation called a net ionic equation:

$$\begin{aligned} \cancel{\mathrm{Ca^{2+}}\,(aq)} + \mathrm{2Cl^-}\,(aq) + \mathrm{2Ag^+}\,(aq) + \cancel{\mathrm{2NO_3{}^-}\,(aq)} &\longrightarrow \cancel{\mathrm{Ca^{2+}}\,(aq)} + \cancel{\mathrm{2NO_3{}^-}\,(aq)} + \mathrm{2AgCl}(s) \\ \mathrm{2Cl^-}\,(aq) + \mathrm{2Ag^+}\,(aq) &\longrightarrow \mathrm{2AgCl}(s) \end{aligned}$$

Following the convention of using the smallest possible integers as coefficients, this equation is then written:

$$\mathrm{Cl^-}\,(aq) + \mathrm{Ag^+}\,(aq) \longrightarrow \mathrm{AgCl}(s)$$

This net ionic equation indicates that solid silver chloride may be produced from dissolved chloride and silver(I) ions, regardless of the source of these ions. These molecular and complete ionic equations provide additional information, namely, the ionic compounds used as sources of Cl^- and Ag^+.

Example 31.2

Ionic and Molecular Equations

When carbon dioxide is dissolved in an aqueous solution of sodium hydroxide, the mixture reacts to yield aqueous sodium carbonate and liquid water. Write balanced molecular, complete ionic, and net ionic equations for this process.

Solution

Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

$$\mathrm{CO_2}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{Na_2CO_3}(aq) + \mathrm{H_2O}(l) \qquad \text{(unbalanced)}$$

Balance is achieved easily in this case by changing the coefficient for NaOH to 2, resulting in the molecular equation for this reaction:

$$\mathrm{CO_2}(aq) + 2\mathrm{NaOH}(aq) \longrightarrow \mathrm{Na_2CO_3}(aq) + \mathrm{H_2O}(l)$$

The two dissolved ionic compounds, NaOH and Na_2CO_3, can be represented as dissociated ions to yield the complete ionic equation:

$$\mathrm{CO_2}(aq) + 2\mathrm{Na^+}(aq) + 2\mathrm{OH^-}(aq) \longrightarrow 2\mathrm{Na^+}(aq) + \mathrm{CO_3}^{2-}(aq) + \mathrm{H_2O}(l)$$

Finally, identify the spectator ion(s), in this case Na^+(*aq*), and remove it from each side of the equation to generate the net ionic equation:

$$\begin{aligned}\mathrm{CO_2}(aq) + 2\mathrm{Na^+}(aq) + 2\mathrm{OH^-}(aq) &\longrightarrow 2\mathrm{Na^+}(aq) + \mathrm{CO_3}^{2-}(aq) + \mathrm{H_2O}(l)\\ \mathrm{CO_2}(aq) + 2\mathrm{OH^-}(aq) &\longrightarrow \mathrm{CO_3}^{2-}(aq) + \mathrm{H_2O}(l)\end{aligned}$$

Check Your Learning

Diatomic chlorine and sodium hydroxide (lye) are commodity chemicals produced in large quantities, along with diatomic hydrogen, via the electrolysis of brine, according to the following unbalanced equation:

$$\mathrm{NaCl}(aq) + \mathrm{H_2O}(l) \xrightarrow{\text{electricity}} \mathrm{NaOH}(aq) + \mathrm{H_2}(g) + \mathrm{Cl_2}(g)$$

Write balanced molecular, complete ionic, and net ionic equations for this process.

Answer

$$
\begin{array}{ll}
2\text{NaCl}(aq) + 2\text{H}_2\text{O}(l) \longrightarrow 2\text{NaOH}(aq) + \text{H}_2(g) + \text{Cl}_2(g) & \text{(molecular)} \\
2\text{Na}^+(aq) + 2\text{Cl}^-(aq) + 2\text{H}_2\text{O}(l) \longrightarrow 2\text{Na}^+(aq) + 2\text{OH}^-(aq) + \text{H}_2(g) + \text{Cl}_2(g) & \\
2\text{Cl}^-(aq) + 2\text{H}_2\text{O}(l) \longrightarrow 2\text{OH}^-(aq) + \text{H}_2(g) + \text{Cl}_2(g) & \text{(net ionic)}
\end{array}
$$

Link to Supplemental Exercises

Supplemental exercises are available if you would like more practice with these concepts.

31.2 Reaction Stoichiometry

Learning Objectives

By the end of this section, you will be able to:

- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products
- Perform stoichiometric calculations involving mass, moles, and solution molarity

A balanced chemical equation provides a great deal of information in a very succinct format. Chemical formulas provide the identities of the reactants and products involved in the chemical change, allowing classification of the reaction. Coefficients provide the relative numbers of these chemical species, allowing a quantitative assessment of the relationships between the amounts of substances consumed and produced by the reaction. These quantitative relationships are known as the reaction's stoichiometry, a term derived from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In this module, the use of balanced chemical equations for various stoichiometric applications is explored.

The general approach to using stoichiometric relationships is similar in concept to the way people go about many common activities. Food preparation, for example, offers an appropriate comparison. A recipe for making eight pancakes calls for 1 cup pancake mix,

$\frac{3}{4}$

cup milk, and one egg. The "equation" representing the preparation of pancakes per this recipe is

$$1 \text{ cup mix} + \frac{3}{4} \text{ cup milk} + 1 \text{ egg} \longrightarrow 8 \text{ pancakes}$$

If two dozen pancakes are needed for a big family breakfast, the ingredient amounts must be increased proportionally according to the amounts given in the recipe. For example, the number of eggs required to make 24 pancakes is

$$24 \text{ pancakes} \times \frac{1 \text{ egg}}{8 \text{ pancakes}} = 3 \text{ eggs}$$

Balanced chemical equations are used in much the same fashion to determine the amount of one reactant required to react with a given amount of another reactant, or to yield a given amount of product, and so forth. The coefficients in the balanced equation are used to derive stoichiometric factors that permit computation of the desired quantity. To illustrate this idea, consider the production of ammonia by reaction of hydrogen and nitrogen:

$$\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \longrightarrow 2\mathrm{NH_3}(g)$$

This equation shows ammonia molecules are produced from hydrogen molecules in a 2:3 ratio, and stoichiometric factors may be derived using any amount (number) unit:

$$\frac{2\ \mathrm{NH_3}\ \text{molecules}}{3\ \mathrm{H_2}\ \text{molecules}} \text{ or } \frac{2\ \text{doz}\ \mathrm{NH_3}\ \text{molecules}}{3\ \text{doz}\ \mathrm{H_2}\ \text{molecules}} \text{ or } \frac{2\ \text{mol}\ \mathrm{NH_3}\ \text{molecules}}{3\ \text{mol}\ \mathrm{H_2}\ \text{molecules}}$$

These stoichiometric factors can be used to compute the number of ammonia molecules produced from a given number of hydrogen molecules, or the number of hydrogen molecules required to produce a given number of ammonia molecules. Similar factors may be derived for any pair of substances in any chemical equation.

Example 31.3

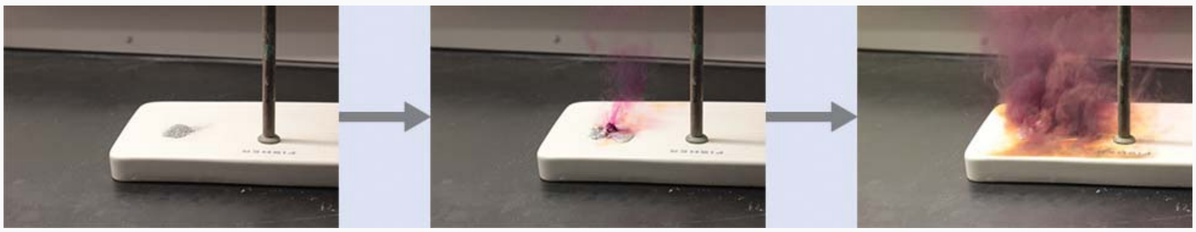
Moles of Reactant Required in a Reaction

How many moles of I_2 are required to react with 0.429 mol of Al according to the following equation (see Figure 31.3)?

$$2Al + 3I_2 \longrightarrow 2AlI_3$$

Figure 31.3

Aluminum and iodine react to produce aluminum iodide. The heat of the reaction vaporizes some of the solid iodine as a purple vapor. (credit: modification of work by Mark Ott)

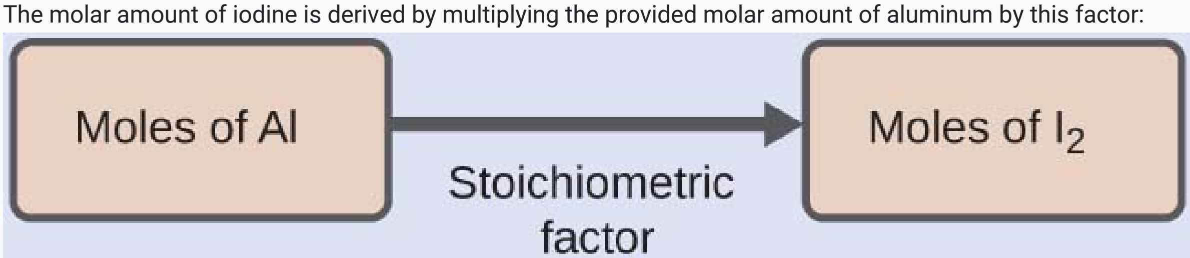


Solution

Referring to the balanced chemical equation, the stoichiometric factor relating the two substances of interest is

$$\frac{3 \text{ mol } I_2}{2 \text{ mol Al}}.$$

The molar amount of iodine is derived by multiplying the provided molar amount of aluminum by this factor:



$$\begin{aligned} \text{mol } I_2 &= 0.429 \text{ mol Al} \times \frac{3 \text{ mol } I_2}{2 \text{ mol Al}} \\ &= 0.644 \text{ mol } I_2 \end{aligned}$$

Check Your Learning

How many moles of $Ca(OH)_2$ are required to react with 1.36 mol of H_3PO_4 to produce $Ca_3(PO_4)_2$ according to the equation

$$3Ca(OH)_2 + 2H_3PO_4 \longrightarrow Ca_3(PO_4)_2 + 6H_2O?$$

Answer

2.04 mol

Example 31.4

Number of Product Molecules Generated by a Reaction

How many carbon dioxide molecules are produced when 0.75 mol of propane is combusted according to this equation?

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

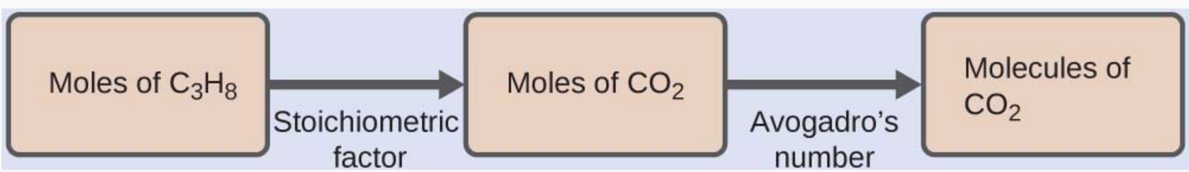
Solution

The approach here is the same as for Example 31.3, though the absolute number of molecules is requested, not the number of moles of molecules. This will simply require use of the moles-to-numbers conversion factor, Avogadro's number.

The balanced equation shows that carbon dioxide is produced from propane in a 3:1 ratio:

$$\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$$

Using this stoichiometric factor, the provided molar amount of propane, and Avogadro's number,



$$0.75 \text{ mol } C_3H_8 \times \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} \times \frac{6.022 \times 10^{23} \; CO_2 \text{ molecules}}{\text{mol } CO_2} = 1.4 \times 10^{24} \; CO_2 \text{ molecul}$$

Check Your Learning

How many NH_3 molecules are produced by the reaction of 4.0 mol of $Ca(OH)_2$ according to the following equation:

$$(NH_4)_2SO_4 + Ca(OH)_2 \longrightarrow 2NH_3 + CaSO_4 + 2H_2O$$

Answer

4.8

$\times$

10^{24} NH_3 molecules

These examples illustrate the ease with which the amounts of substances involved in a chemical reaction of known stoichiometry may be related. Directly measuring numbers of atoms and molecules is, however, not an easy task, and the practical application of stoichiometry requires that we use the more readily measured property of mass.

Example 31.5

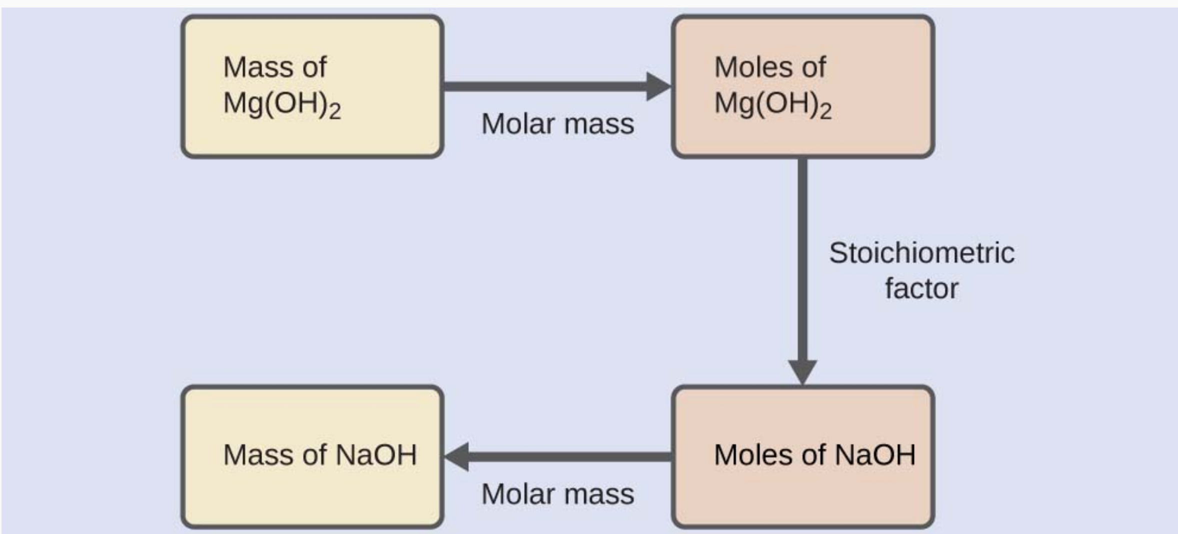
Relating Masses of Reactants and Products

What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the following reaction?

$$MgCl_2\,(aq) + 2NaOH(aq) \longrightarrow Mg(OH)_2\,(s) + 2NaCl(aq)$$

Solution

The approach used previously in Example 31.3 and in Example 31.4 is likewise used here; that is, we must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart:



$$16\text{ g Mg(OH)}_2 \times \frac{1\text{ mol Mg(OH)}_2}{58.3\text{ g Mg(OH)}_2} \times \frac{2\text{ mol NaOH}}{1\text{ mol Mg(OH)}_2} \times \frac{40.0\text{ g NaOH}}{\text{mol NaOH}} = 22\text{ g NaOH}$$

Check Your Learning

What mass of gallium oxide, Ga_2O_3, can be prepared from 29.0 g of gallium metal? The equation for the reaction is

$$4Ga + 3O_2 \longrightarrow 2Ga_2O_3\ .$$

Answer

39.0 g

Example 31.6

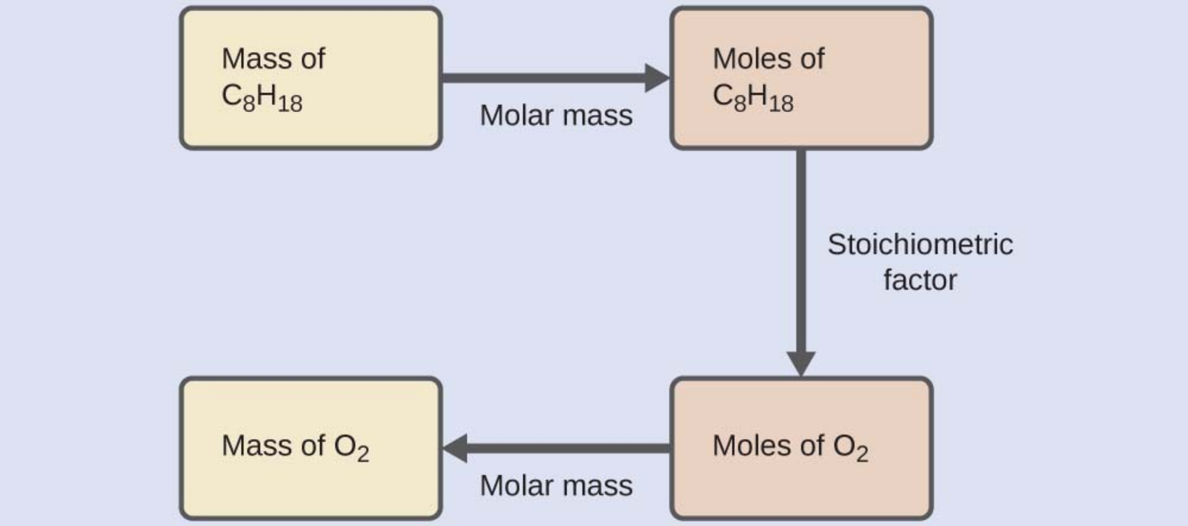
Relating Masses of Reactants

What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C_8H_{18}, one of the principal components of gasoline?

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

Solution

The approach required here is the same as for Example 31.5, differing only in that the provided and requested masses are both for reactant species.



$$702 \text{ g } C_8H_{18} \times \frac{1 \text{ mol } C_8H_{18}}{114.23 \text{ g } C_8H_{18}} \times \frac{25 \text{ mol } O_2}{2 \text{ mol } C_8H_{18}} \times \frac{32.00 \text{ g } O_2}{\text{mol } O_2} = 2.46 \times 10^3 \text{ g } O_2$$

Check Your Learning

What mass of CO is required to react with 25.13 g of Fe_2O_3 according to the equation

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2 \text{ ?}$$

Answer

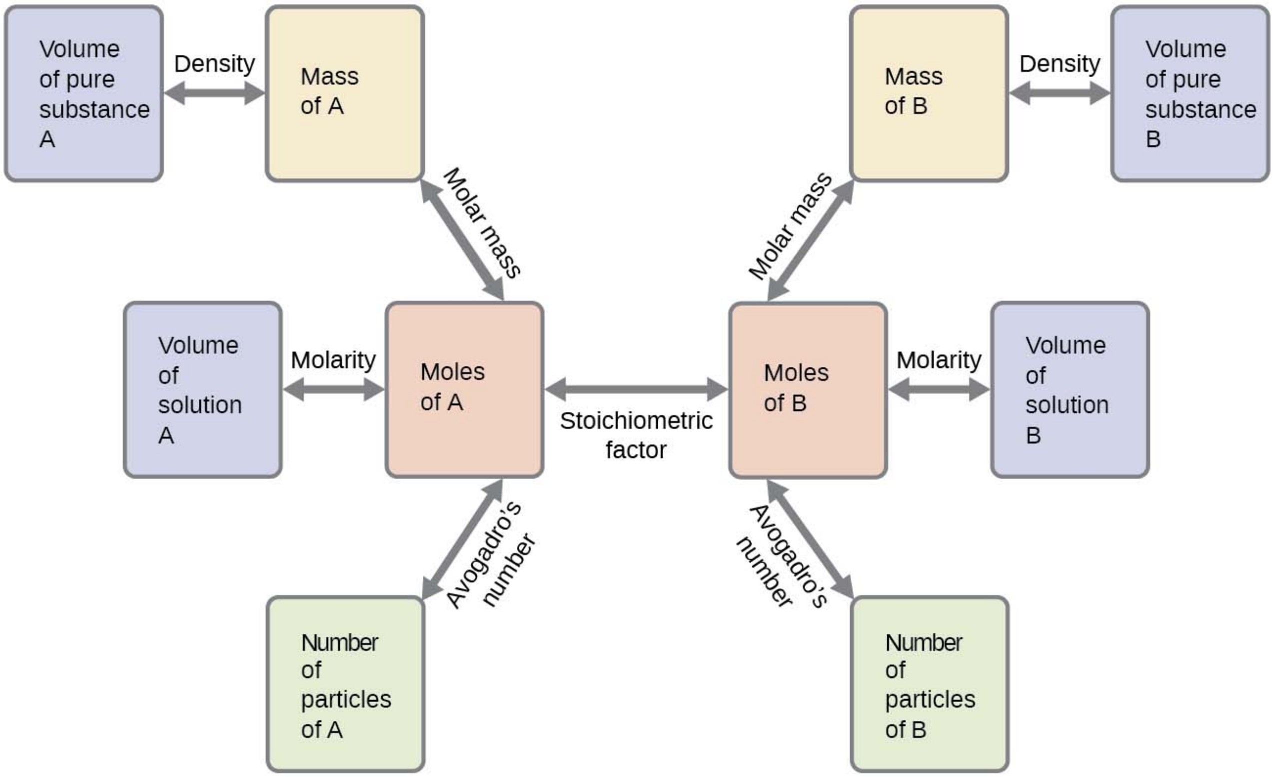
13.22 g

These examples illustrate just a few instances of reaction stoichiometry calculations. Numerous variations on the beginning and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. Figure

31.6 provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

Figure 31.6

The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations.



Chemistry in Everyday Life

Airbags

Airbags (Figure 31.7) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

$$2NaN_3(s) \longrightarrow 3N_2(g) + 2Na(s)$$

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.

Figure 31.7

Airbags deploy upon impact to minimize serious injuries to passengers. (credit: Jon Seidman)

Link to Supplemental Exercises

Supplemental exercises are available if you would like more practice with these concepts.

Files

Open in Google Drive

Previous Citation(s)

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