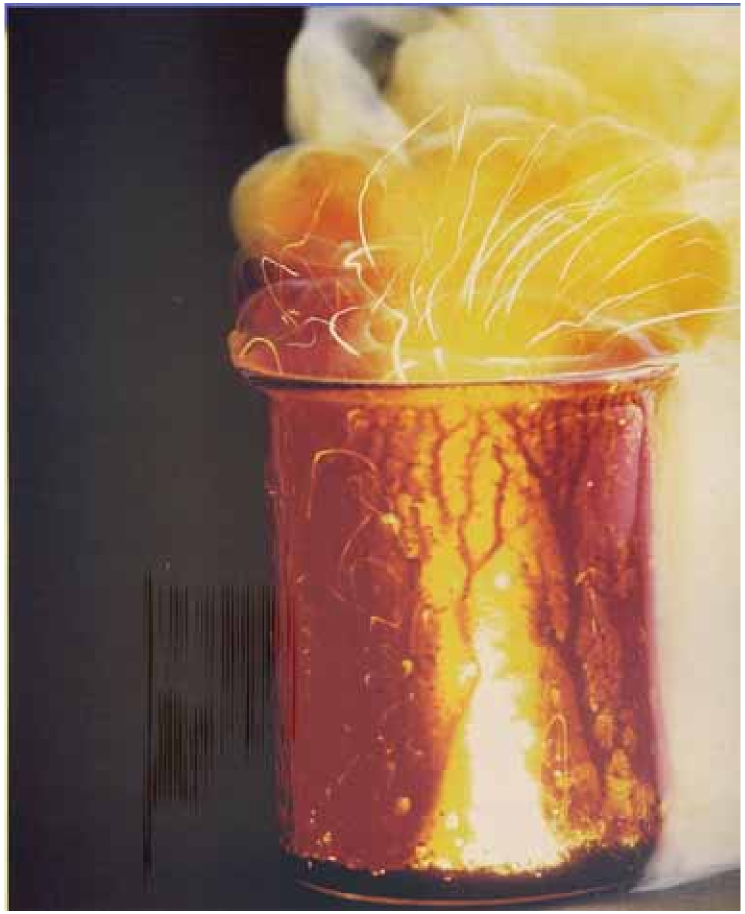
Chemistry

Zumdahl, 7th edition

歐亞書局

CH3 Stoichiometry



The violent chemical reaction of bromine and phosphorus.

Contents

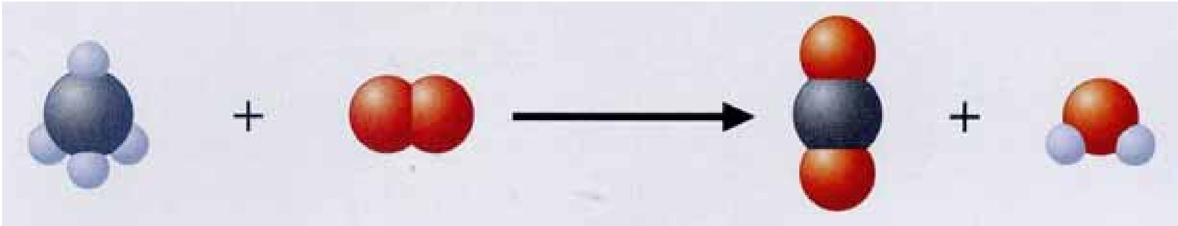
3.7 Chemical equations

Chemical Reactions

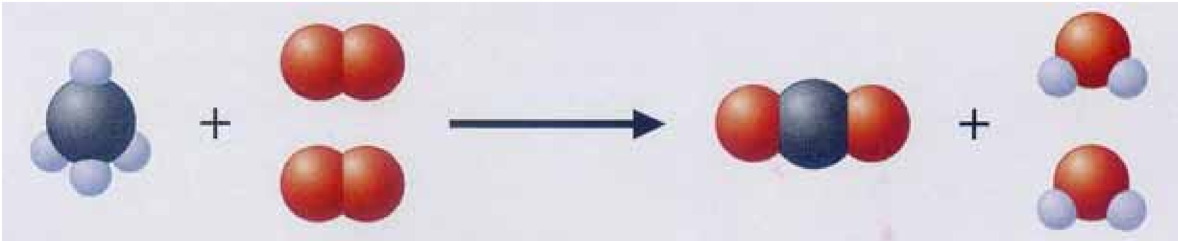
- A chemical change involves a reorganization of the atoms in one or more substances.
- This process is represented by a chemical equation with the reactants (here methane and oxygen) on the left side of an arrow and the products (carbon dioxide and water) on the right side:

$$CH_4 + O_2 \longrightarrow CO_2 + H_2O$$

- *Bonds have been broken, and new ones have been formed*. It is important to recognize that *in a chemical reaction, atoms are neither created nor destroyed.*
- *All atoms present in the reactants must be accounted for among the products*.
- In other words, there must be the same number of each type of atom on the product side and on the reactant side of the arrow.
- Making sure that this rule is obeyed is called **balancing a chemical equation** for a reaction.



- The needed numbers of molecules are



- Notice that now we have the same number of each type of atom represented among the reactants and the products.

- We can represent the preceding situation in a shorthand manner by the following chemical equation:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

Reactants: 1 C, 4H, 4 O — Products: 1 C, 2 O, 4 H, 2 O

Reactants	Products
1 C	1 C
4 H	4 H
4 O	4 O

Methane reacts with oxygen to produce the flame in a Bunsen burner.

The Meaning of a Chemical Equation

- The chemical equation for a reaction gives two important types of information: the nature of the reactants and products and the relative numbers of each.
- Besides specifying the compounds involved in the reaction, the equation often gives the physical states of the reactants and products:

State	Symbol
Solid	(*s*)
Liquid	(*l*)
Gas	(*g*)
Dissolved in water (in aqueous solution)	(*aq*)

$$HCl(aq) + NaHCO_3(s) \longrightarrow CO_2(g) + H_2O(l) + NaCl(aq)$$

- The relative numbers of reactants and products in a reaction are indicated by the coefficients in the balanced equation. (The coefficients can be determined because we know that the same number of each type of atom must occur on both sides of the equation.)

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Hydrochloric acid reacts with solid sodium hydrogen carbonate to produce gaseous carbon dioxide.

TABLE 3.2 Information Conveyed by the Balanced Equation for the Combustion of Methane

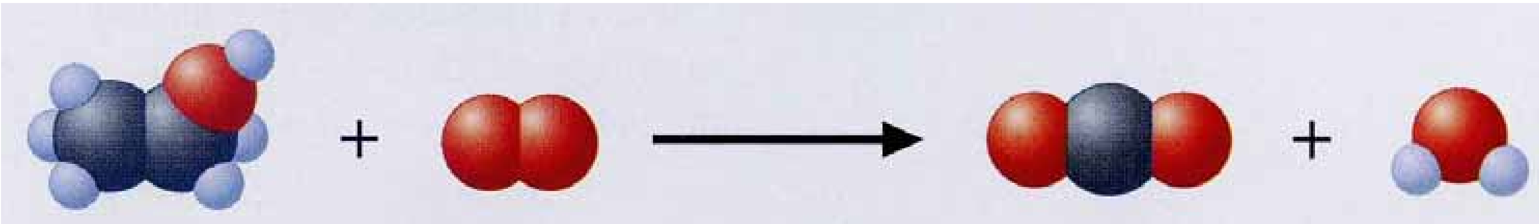
Reactants		Products
$CH_4(g) + 2O_2(g)$	$\longrightarrow$	$CO_2(g) + 2H_2O(g)$
1 molecule + 2 molecules	$\longrightarrow$	1 molecule + 2 molecules
1 mole + 2 moles	$\longrightarrow$	1 mole + 2 moles
6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)	$\longrightarrow$	6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)
16 g + 2 (32 g)		44 g + 2 (18 g)
80 g reactants	$\longrightarrow$	80 g products

3.8 Balancing Chemical Equations

- Whenever you see an equation, you should ask yourself whether it is balanced.
- The principle that lies at the heart of the balancing process is that atoms are conserved in a chemical reaction.
- It is also important to recognize that the identities of the reactants and products of a reaction are determined by experimental observation.

- When the equation for this reaction is balanced, the identities of the reactants and products must not be changed.
- *The formulas of the compounds must never be changed in balancing a chemical equation.*

$$C_2H_5OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$



$$C_2H_5OH(l) + O_2(g) \longrightarrow 2CO_2(g) + H_2O(g)$$

2 C atoms　　　　　2 C atoms

$$C_2H_5OH(l) + O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

(5＋1) H　　　　　(3 x 2) H

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

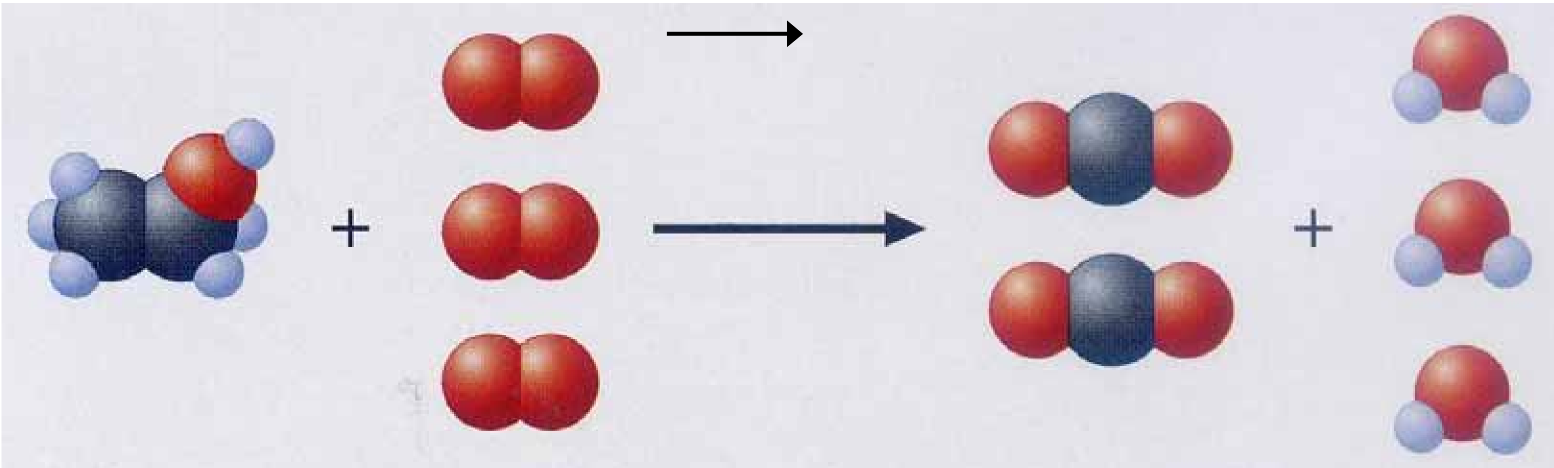
1 O　　6 O　　　　(2 x 2) O　　3 O

7 O　　　　　7 O

$$C_2H_5OH(l) + O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

2 C atoms	2 C atoms
6 H atoms	6 H atoms
7 O atoms	7 O atoms

- The balanced equation can be represented as follows:



Writing and Balancing the Equation for a Chemical Reaction

1 Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?

2 Write the unbalanced equation that summarizes the reaction described in step 1.

3 Balance the equation by inspection, starting with the most complicated mole-cule(s). Determine what coefficients are necessary so that the same number of each type of atom appears on both reactant and product sides. Do not change the identities (formulas) of any of the reactants or products.

Sample Exercise 3.14

Balancing a Chemical Equation Ⅰ

Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, $(NH_4)_2Cr_2O_7$, a vivid orange compound, is ignited, a spectacular reaction occurs, as shown in two photographs on the next page. Although the reaction is actually somewhat more complex, let's assume here that the products are solid chromium(Ⅲ) oxide, nitrogen gas (consisting of N_2 molecules), and water vapor. Balance the equation for this reaction.

Sample Exercise 3.14

Solution

1 From the description given, the reactant is solid ammonium dichromate, $(NH_4)_2Cr_2O_7(s)$, and the products are nitrogen gas, $N_2(g)$, water vapor, $H_2O(g)$, and solid chromium(III) oxide, $Cr_2O_3(s)$. The formula for chromium(III) oxide can be determined by recognizing that the Roman numeral III means that Cr^{3+} ions are present. For a neutral compound, the formula must then be Cr_2O_3, since each oxide ion is O^{2-}.

Sample Exercise 3.14

➡ 2 The unbalanced equation is

$$(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + H_2O(g)$$

➡ 3

$$(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + 4H_2O(g)$$

(4 × 2) H (4 × 2) H

Reality Check:

2 N, 8 H, 2 Cr, 7 O → 2 N, 8 H, 2 Cr, 7 O

Reactant atoms Product atoms

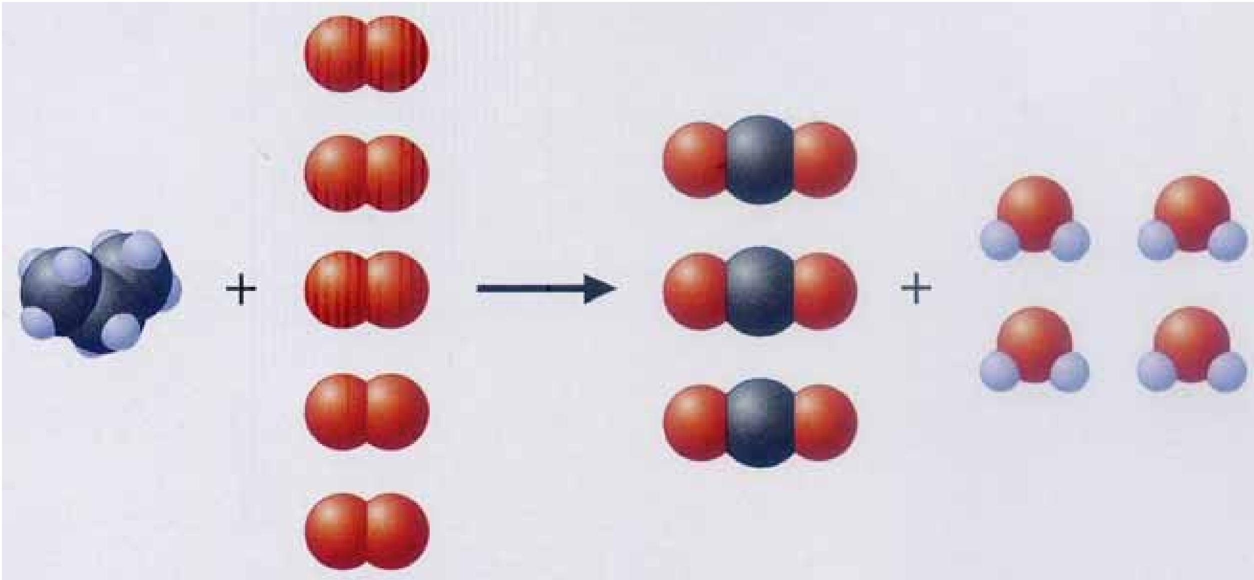
See Exercises 3.81 and 3.82

3.9 Stoichiometric Calculations: Amounts of Reactants and Products

- The coefficients in chemical equations represent numbers of molecules, not masses of molecules.
- In this section we will see how chemical equations can be used to determine the masses of reacting chemicals.
- To develop the principles for dealing with the stoichiometry of reactions, we will consider the reaction of propane with oxygen to produce carbon dioxide and water.

- We will consider the question: "*What mass of oxygen will react with 96.1 grams of propane*?"
- In doing stoichiometry, the first thing we must do is write the balanced chemical equation for the reaction. In this case the balanced equation is

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$



$$96.1\ \cancel{g\ C_3H_8} \times \frac{1\ mol\ C_3H_8}{44.1\ \cancel{g\ C_3H_8}} = 2.18\ mol\ C_3H_8$$

The best way to do this is to use the balanced equation to construct a **mole ratio**.

$$\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}$$

$$2.18 \text{ mol } C_3H_8 \times \frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8} = 10.9 \text{ mol } O_2$$

$$10.9 \text{ mol } O_2 \times \frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2} = 349 \text{ g } O_2$$

- Therefore, 349 grams of oxygen is required to burn 96.1 grams of propane.
- The example can be extended by asking: "*What mass of carbon dioxide is produced when 96.1 grams of propane is combusted with oxygen*?"

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

- We will now summarize the sequence of steps needed to carry out stoichiometric calculations.

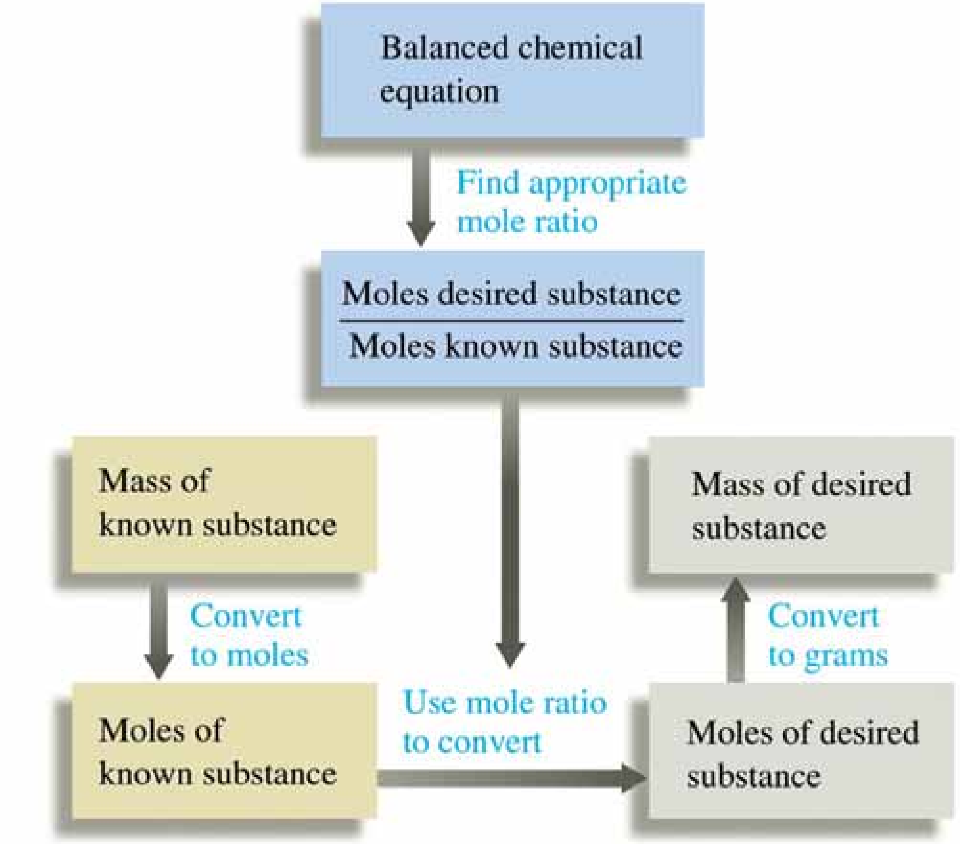
$$96.1\ \text{g}\ C_3H_8 \xrightarrow{\frac{1\ \text{mol}\ C_3H_8}{44.1\ \text{g}\ C_3H_8}} 2.18\ \text{mol}\ C_3H_8 \xrightarrow{\frac{3\ \text{mol}\ CO_2}{1\ \text{mol}\ C_3H_8}} 6.54\ \text{mol}\ CO_2$$

$$\xrightarrow{\frac{44.0\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2}} 288\ \text{g}\ CO_2$$

Calculating Masses of Reactants and Products in Chemical Reactions

1. **Balance the equation for the reaction.**
2. **Convert the known mass of the reactant or product to moles of that substance.**
3. **Use the balanced equation to set up the appropriate mole ratios.**
4. **Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product.**
5. **Convert from moles back to grams if required by the problem.**

- These steps are summarized by the following diagram:



Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00 kg of lithium hydroxide?

Solution

1 Using the description of the reaction, we can write the unbalanced equation:

$$LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

Sample Exercise 3.16

The balanced equation is

$$2\text{LiOH}(s) + \text{CO}_2(g) \longrightarrow \text{Li}_2\text{CO}_3(s) + \text{H}_2\text{O}(l)$$

➡ 2 We convert the given mass of LiOH to moles, using the molar mass of LiOH (6.941 + 16.00 + 1.008 = 23.95 g/mol):

$$1.00\ \text{kg LiOH} \times \frac{1000\ \text{g LiOH}}{1\ \text{kg LiOH}} \times \frac{1\ \text{mol LiOH}}{23.95\ \text{g LiOH}} = 41.8\ \text{mol LiOH}$$

➡ 3 Since we want to determine the amount of CO_2 that reacts with the given amount of LiOH, the appropriate mole ratio is

Sample Exercise 3.16

$$\frac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}}$$

4 We calculate the moles of CO_2 needed to react with the given mass of LiOH using this mole ratio:

$$41.8 \text{ mol LiOH} \times \frac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}} = 20.9 \text{ mol } CO_2$$

3 Next we calculate the mass of CO_2, using its molar mass (44.0 g/mol):

$$20.9 \text{ mol } CO_2 \times \frac{44.0 \text{ g } CO_2}{1 \text{ mol } CO_2} = 9.20 \times 10^2 \text{ g } CO_2$$

Sample Exercise 3.16

Thus 920. g of $CO_2(g)$ will be absorbed by 1.00 kg of LiOH(s).

Astronaut Sidney M. Gutierrez changes the lithium hydroxide cannisters on space shuttle Columbia. The lithium hydroxide is used to purge carbon dioxide from the air in the shuttle's cabin.

See Exercises 3.89 and 3.90

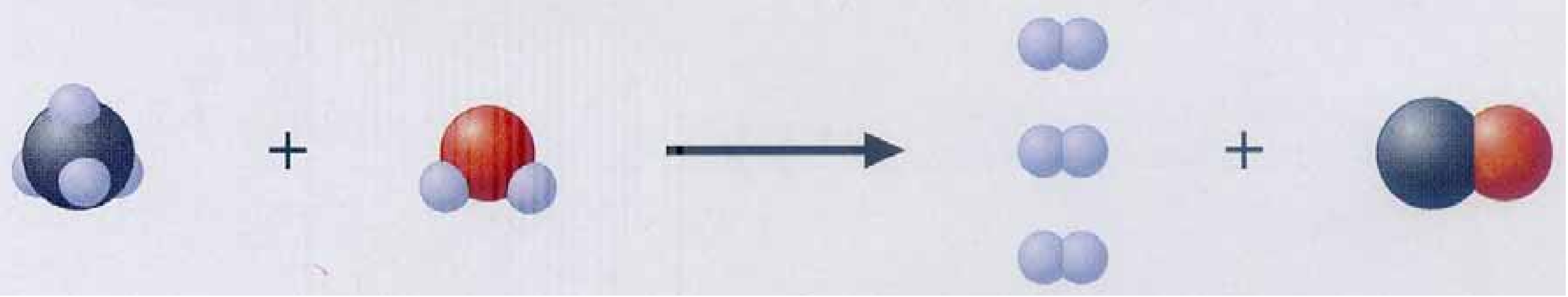
3.10 Calculations Involving a Limiting Reactant

- When chemicals are mixed together to undergo a reaction, they are often mixed in **stoichiometric quantities**.
- To clarify this concept, let's consider the production of hydrogen for use in the manufacture of ammonia by the **Haber process**.
- Ammonia is made by combining nitrogen (from the air) with hydrogen according to the equation

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

- Hydrogen can be obtained from the reaction methane with water vapor:

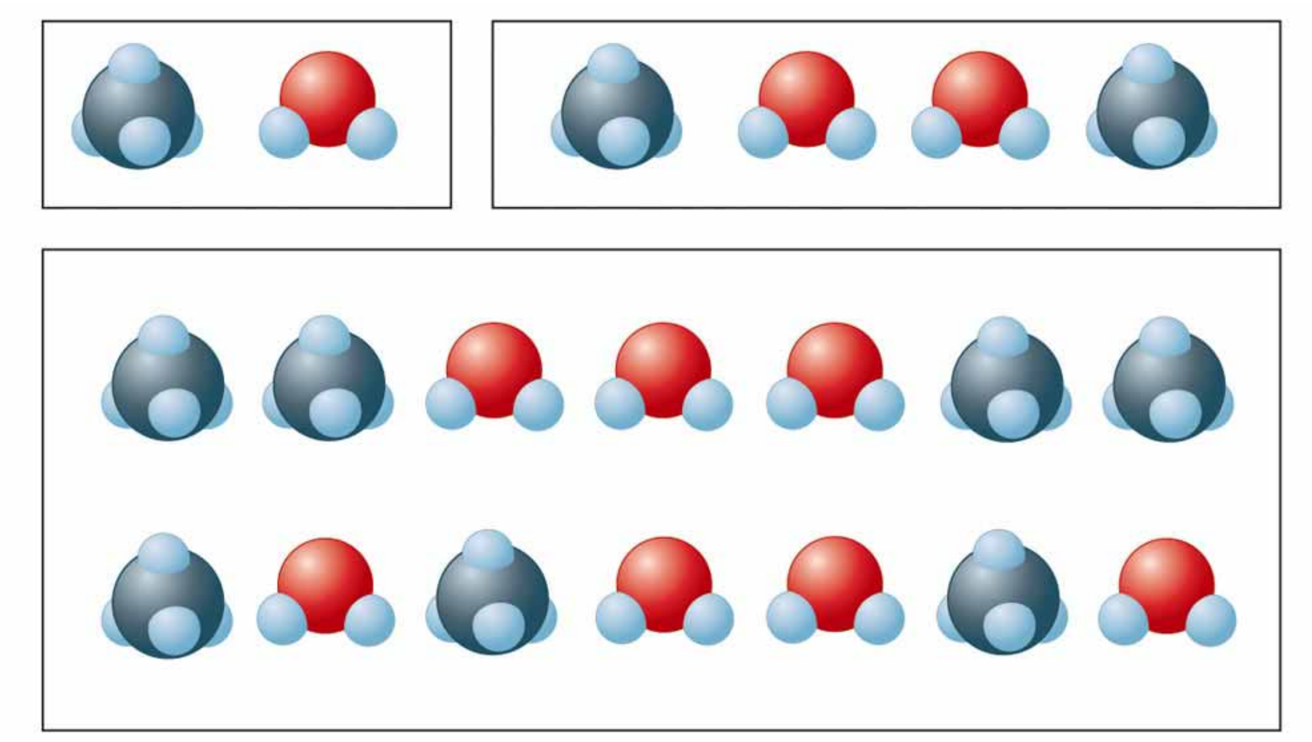
$$CH_4(g) + H_2O(g) \longrightarrow 3H_2(g) + CO(g)$$



- We first need to find the number of moles of methane molecules in 2.50×10^3 kg (2.50×10^6 g) of methane:

$$1.56 \times 10^5 \text{ mol } CH_4 \text{ molecules}$$

Figure 3.9



Three different stoichiometric mixtures of methane and water, which react one-to-one.

- We first need to find the number of moles of methane molecules in 2.50×10^3 kg (2.50×10^6 g) of methane:

$$1.56 \times 10^5 \text{ mol } CH_4 \text{ molecules}$$

- If 2.50×10^3 kilograms of methane is mixed with 2.81×10^3 kilograms of water, both reactants will “run out” at the same time.
- If 2.50×10^3 kilograms of methane is mixed with 3.00×10^3 kilograms of water, the methane will be consumed before the water runs out.

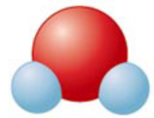
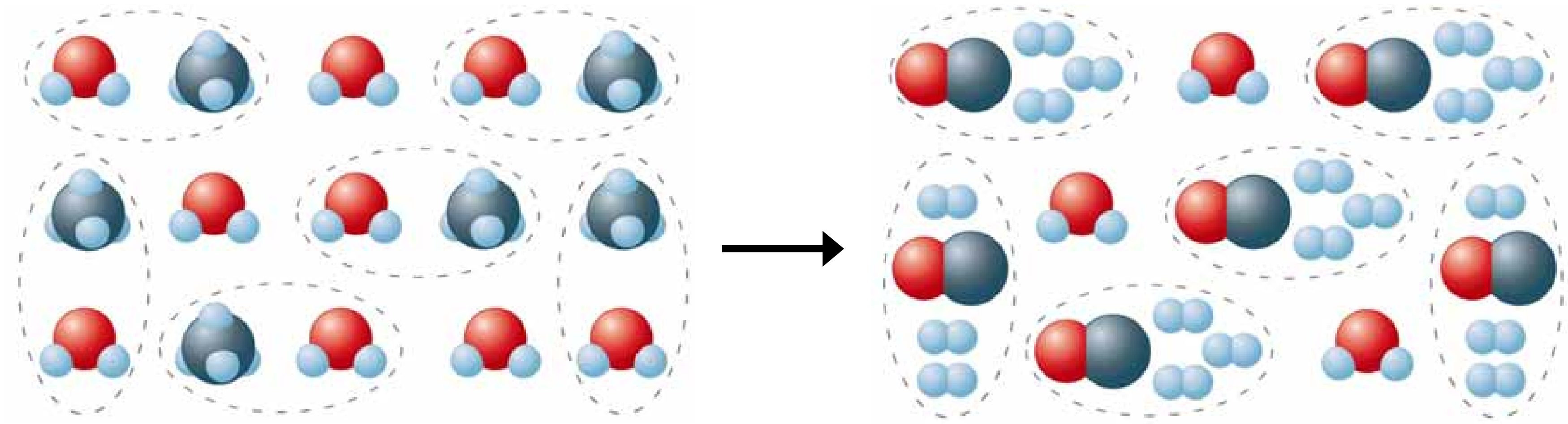
- The water will be in excess; that is, there will be more water molecules than methane molecules in the reaction mixture.
- First picture the mixture of CH_4  and H_2O  molecules as shown in Fig. 3.10.
- Then imagine that groups consisting of one CH_4 molecule and one H_2O molecule (Fig. 3.10) will react to form three H_2 and one CO molecules (Fig. 3.11).
- Notice that products can form only when both CH_4 and H_2O are available to react.

Figure 3.10 & 3.11

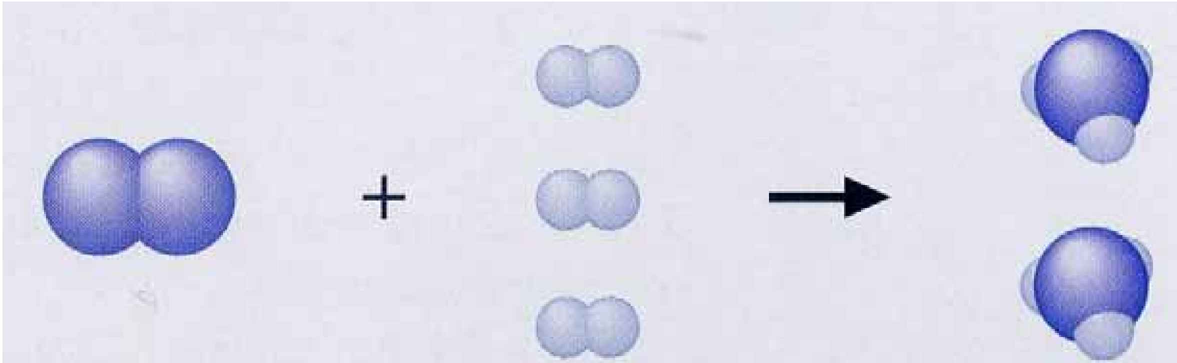


A mixture of CH_4 and H_2O molecules.

Methane and water have reacted to form products according to the equation $CH_4 + H_2O \rightarrow 3H_2 + CO$.

- Thus the number of products that can form is limited by the methane.
- This brings us to the concept of the **limiting reactant** (or **limiting reagent**), which is the reactant that is consumed first and that therefore limits the amounts of products that can be formed.
- To further explore the idea of a limiting reactant, consider the ammonia synthesis reaction:

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

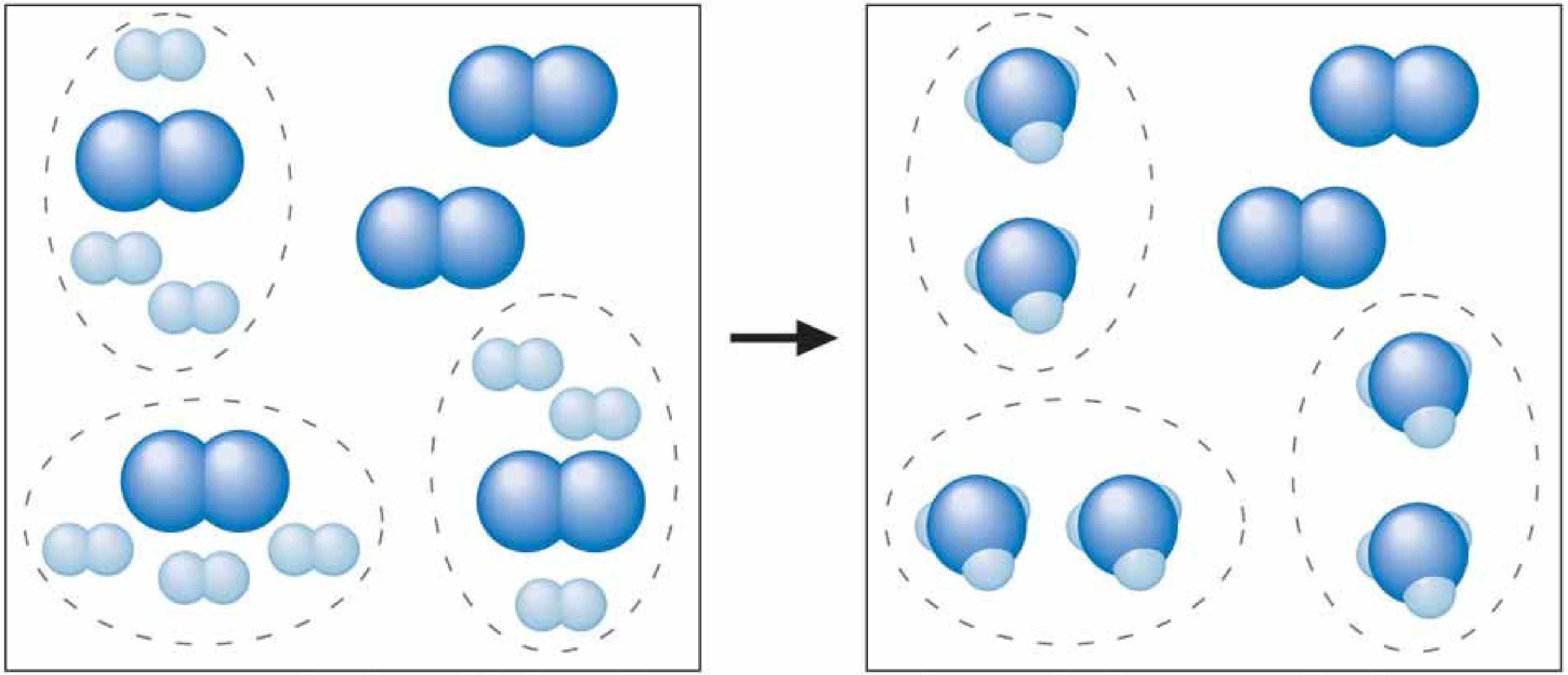


- Figure 3.12 shows that 3 of the N_2 molecules react with the 9 H_2 molecules to produce 6 NH_3 molecules:

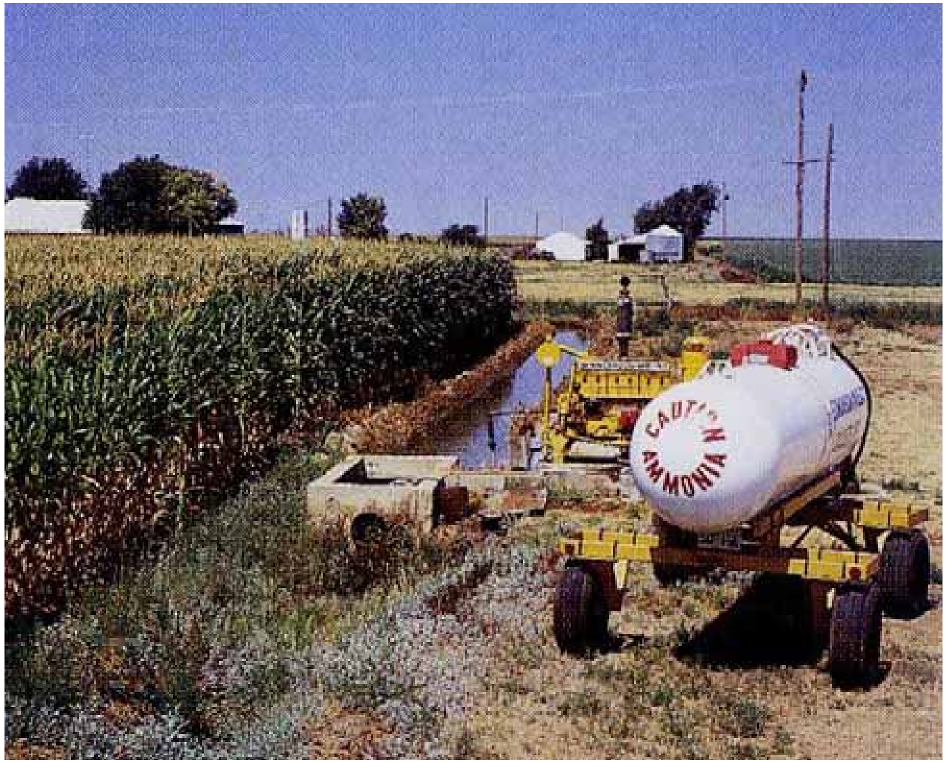
$$3N_2 + 9H_2 \longrightarrow 6NH_3$$

- This leaves 2 N_2 molecules unreacted — the nitrogen is in excess.

Figure 3.12



Hydrogen and nitrogen react to form ammonia according to the equation $N_2 + 3H_2 \rightarrow 2NH_3$.



Ammonia is dissolved in irrigation water to provide fertilizer for a field of corn.

- The most important point here is this: *The limiting reactant limits the amount of product that can form.*
- The reaction that actually occurred was

$$3N_2(g) + 9H_2(g) \longrightarrow 6NH_3(g)$$

- not

$$5N_2(g) + 15H_2(g) \longrightarrow 10NH_3(g)$$

- Thus 6 NH_3 were formed, not 10 NH_3, because the H_2, not the N_2, was limiting.

Sample Exercise 3.18

Stoichiometry: Limiting Reactant

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(Ⅱ) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1 g of NH_3 is reacted with 90.4 g of CuO, which is the limiting reactant? How many grams of N_2 will be formed?

- The amount of a product formed when the limiting reactant is completely consumed is called the **theoretical yield** of that product.
- The *actual yield* of product is often given as a percentage of the theoretical yield. This is called the **percent yield**:

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \text{percent yield}$$

Sample Exercise 3.19

Calculating Percent Yield

Methanol (CH_3OH), also called methyl alcohol, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combination of gaseous carbon monoxide and hydrogen. Suppose 68.5 kg $CO(g)$ is reacted with 8.60 kg $H_2(g)$. calculate the theoretical yield of methanol. If 3.57×10^4 g CH_3OH is actually produced, what is the percent yield of methanol?

Methanol is used as a fuel in Indianapolis-type racing cars.

Solving a Stoichiometry Problem Involving Masses of Reactants and Products

1. **Write and balance the equation for the reaction.**
2. **Convert the known masses of substances to moles.**
3. **Determine which reactant is limiting.**
4. **Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.**
5. **Convert from moles to grams, using the molar mass.**

- This process is summarized in the diagram below: