Consider the balanced chemical equation (i.e., catalytic oxidation of ammonia) such as
\[ \text{4 NH}_3 (g) + 5 \text{O}_2 (g) \rightarrow 4 \text{NO} (g) + 6 \text{H}_2 \text{O} (g) \]
not only tells how many molecules of each kind are involved in a reaction, it also indicates the amount of each substance that is involved. Equation \ref{1} (represented molecularly by the image below it) says that 4 \text{NH}_3 \text{ molecules} can react with 5 \text{O}_2 \text{ molecules} to give 4 \text{NO} \text{ molecules} and 6 \text{H}_2 \text{O} \text{ molecules}. It also says that 4 \text{mol NH}_3 \text{ would react with 5 mol O}_2 \text{ yielding 4 mol NO and 6 mol H}_2 \text{O}.

The balanced equation does more than this, though. It also tells us that \(2 \cdot 4 = 8 \text{mol NH}_3\) will react with \(2 \cdot 5 = 10 \text{mol O}_2\), and that \(\frac{1}{2} \cdot 4 = 2 \text{mol NH}_3\) requires only \(\frac{1}{2} \cdot 5 = 2.5 \text{mol O}_2\). In other words, the equation indicates that exactly 5 \text{mol O}_2 must react for every 4 \text{mol NH}_3 consumed. For the purpose of calculating how much \text{O}_2 is required to react with a certain amount of \text{NH}_3 therefore, the significant information contained in Equation \ref{1} is the ratio
\[ \frac{5 \text{mol O}_2}{4 \text{mol NH}_3} \]
We shall call such a ratio derived from a balanced chemical equation a stoichiometric ratio and give it the symbol \(S\). Thus, for Equation \ref{1},
\[
\frac{\text{5 mol O}_2}{\text{4 mol NH}_3}\]
The word stoichiometric comes from the Greek words stoicheion, “element,” and metron, “measure.” Hence the stoichiometric ratio measures one element (or compound) against another.

Example \PageIndex{1}: Stoichiometric Ratios
Derive all possible stoichiometric ratios from Equation \ref{1}.

Solution
Any ratio of amounts of substance given by coefficients in the equation may be used:
\[
\text{S} \left( \frac{\text{NH}_3}{\text{O}_2} \right) = \frac{4 \text{mol NH}_3}{5 \text{mol O}_2} \\
\text{S} \left( \frac{\text{O}_2}{\text{NO}} \right) = \frac{5 \text{mol O}_2}{4 \text{mol NO}} \\
\text{S} \left( \frac{\text{NH}_3}{\text{NO}} \right) = \frac{4 \text{mol NH}_3}{4 \text{mol NO}} \\
\text{S} \left( \frac{\text{O}_2}{\text{H}_2 \text{O}} \right) = \frac{5 \text{mol O}_2}{6 \text{mol } \text{H}_2 \text{O}} \\
\text{S} \left( \frac{\text{NH}_3}{\text{H}_2 \text{O}} \right) = \frac{4 \text{mol NH}_3}{6 \text{mol } \text{H}_2 \text{O}} \\
\text{S} \left( \frac{\text{NO}}{\text{H}_2 \text{O}} \right) = \frac{4 \text{mol NO}}{6 \text{mol } \text{H}_2 \text{O}}
\]
There are six more stoichiometric ratios, each of which is the reciprocal of one of these. [Equation (3) gives one of them. ]

When any chemical reaction occurs, the amounts of substances consumed or produced are related by the appropriate stoichiometric ratios. Using Equation (2) as an example, this means that the ratio of the amount of O₂ consumed to the amount of NH₃ consumed must be the stoichiometric ratio S(O₂/NH₃):

\[
\frac{n_{\text{O}_2\text{ consumed}}}{n_{\text{NH}_3\text{ consumed}}} = \text{S}
\left(\frac{\text{O}_2}{\text{NH}_3}\right) = \frac{\text{5 mol O}_2}{\text{4 mol NH}_3}
\]  

\[\label{9}\]

Similarly, the ratio of the amount of H₂O produced to the amount of NH₃ consumed must be S(H₂O/NH₃):

\[
\frac{n_{\text{H}_2\text{O produced}}}{n_{\text{NH}_3\text{ consumed}}} = \text{S}\left(\frac{\text{H}_2\text{O}}{\text{NH}_3}\right) = \frac{\text{6 mol H}_2\text{O}}{\text{4 mol NH}_3}
\]

\[\label{10}\]

In general we can say that

\[
\text{Stoichiometric ratio } \left(\frac{X}{Y}\right)=\frac{\text{amount of X consumed or produced}}{\text{amount of Y consumed or produced}}
\]

\[\label{11}\]

or, in symbols,

\[
\text{S}\left(\frac{X}{Y}\right)= \frac{n_{\text{X consumed or produced}}}{n_{\text{Y consumed or produced}}}
\]

\[\label{12}\]

Note that in the word Equation (11) and the symbolic Equation (12), \(X\) and \(Y\) may represent any reactant or any product in the balanced chemical equation from which the stoichiometric ratio was derived. No matter how much of each reactant we have, the amounts of reactants consumed and the amounts of products produced will be in appropriate stoichiometric ratios.

**Example \(\PageIndex{2}\): Ratio of Water**

Find the amount of water produced when 3.68 mol NH₃ is consumed according to Equation (10).

**Solution**

The amount of water produced must be in the stoichiometric ratio S(H₂O/NH₃) to the amount of ammonia consumed:

\[
\begin{align}
\text{S}\left(\frac{\text{H}_2\text{O}}{\text{NH}_3}\right) &= \frac{n_{\text{H}_2\text{O produced}}}{n_{\text{NH}_3\text{ consumed}}} \\
&= \frac{3.68 \text{ mol NH}_3 \cdot \frac{6 \text{ mol H}_2\text{O}}{4 \text{ mol NH}_3}}{\text{5.52 mol H}_2\text{O}} \\
&= 5.52 \text{ mol H}_2\text{O}
\end{align}
\]

In general we can say that

\[
\text{Stoichiometric ratio } \left(\frac{X}{Y}\right)=\frac{\text{amount of X consumed or produced}}{\text{amount of Y consumed or produced}}
\]

\[\label{11}\]

or, in symbols,

\[
\text{S}\left(\frac{X}{Y}\right)= \frac{n_{\text{X consumed or produced}}}{n_{\text{Y consumed or produced}}}
\]

\[\label{12}\]

Note that in the word Equation (11) and the symbolic Equation (12), \(X\) and \(Y\) may represent any reactant or any product in the balanced chemical equation from which the stoichiometric ratio was derived. No matter how much of each reactant we have, the amounts of reactants consumed and the amounts of products produced will be in appropriate stoichiometric ratios.
This is a typical illustration of the use of a stoichiometric ratio as a conversion factor. Example \(\PageIndex{2}\) is analogous to Examples 1 and 2 from Conversion Factors and Functions, where density was employed as a conversion factor between mass and volume. Example \(\PageIndex{2}\) is also analogous to Examples 2.4 and 2.6, in which the Avogadro constant and molar mass were used as conversion factors. As in these previous cases, there is no need to memorize or do algebraic manipulations with Equation \((\ref{9})\) when using the stoichiometric ratio. Simply remember that the coefficients in a balanced chemical equation give stoichiometric ratios, and that the proper choice results in cancellation of units. In road-map form

\[
\text{amount of X consumed or produced} \overset{\text{stoichiometric ratio X/Y}}{\longleftrightarrow} \text{amount of Y consumed or produced}\]

or symbolically.

\[
n_{\text{X consumed or produced}} \overset{S(\text{X/Y})}{\longleftrightarrow} n_{\text{Y consumed or produced}}\]

When using stoichiometric ratios, be sure you always indicate moles of what. You can only cancel moles of the same substance. In other words, 1 mol NH\(_3\) cancels 1 mol NH\(_3\) but does not cancel 1 mol H\(_2\)O.

The next example shows that stoichiometric ratios are also useful in problems involving the mass of a reactant or product.

Example \(\PageIndex{3}\): Mass Produced

Calculate the mass of sulfur dioxide (SO\(_2\)) produced when 3.84 mol O\(_2\) is reacted with FeS\(_2\) according to the equation

\[
\text{4FeS}_2 + 11\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 + 8\text{SO}_2
\]

Solution

The problem asks that we calculate the mass of SO\(_2\) produced. As we learned in Example 2 of The Molar Mass, the molar mass can be used to convert from the amount of SO\(_2\) to the mass of SO\(_2\). Therefore this problem in effect is asking that we calculate the amount of SO\(_2\) produced from the amount of O\(_2\) consumed. This is the same problem as in Example 2. It requires the stoichiometric ratio

\[
\text{S}(\frac{\text{SO}_2}{\text{O}_2}) = \frac{8 \text{ mol SO}_2}{11 \text{ mol O}_2}
\]

The amount of SO\(_2\) produced is then

\[
\begin{align*}
n_{\text{SO}_2 \text{ produced}} & = n_{\text{O}_2 \text{ consumed}} \cdot \text{ conversion factor} \\
& = 3.84 \text{ mol O}_2 \cdot \frac{8 \text{ mol SO}_2}{11 \text{ mol O}_2} \\
& = 2.79 \text{ mol SO}_2
\end{align*}
\]

The mass of SO\(_2\) is

\[
\begin{align*}
m_{\text{SO}_2} & = n_{\text{SO}_2 \text{ produced}} \cdot \text{ molar mass of SO}_2 \\
& = 2.79 \text{ mol SO}_2 \cdot 64.06 \text{ g SO}_2/\text{mol SO}_2 \\
& = 177.0 \text{ g SO}_2
\end{align*}
\]
With practice this kind of problem can be solved in one step by concentrating on the units. The appropriate stoichiometric ratio will convert moles of \( \text{O}_2 \) to moles of \( \text{SO}_2 \) and the molar mass will convert moles of \( \text{SO}_2 \) to grams of \( \text{SO}_2 \). A schematic road map for the one-step calculation can be written as

\[
\begin{align*}
  n_{\text{O}_2} & \xrightarrow{\text{S}(\text{SO}_2/\text{O}_2)} n_{\text{SO}_2} \\
  n_{\text{SO}_2} & \xrightarrow{M_{\text{SO}_2}} m_{\text{SO}_2}
\end{align*}
\]

Thus

\[
\begin{align*}
  m_{\text{SO}_2} &= 3.84 \text{ mol O}_2 \cdot \frac{8 \text{ mol SO}_2}{11 \text{ mol O}_2} \cdot \frac{64.06 \text{ g}}{1 \text{ mol SO}_2} = 179 \text{ g}
\end{align*}
\]

These calculations can be organized as a table, with entries below the respective reactants and products in the chemical equation. You may verify the additional calculations.

<table>
<thead>
<tr>
<th>( \text{FeS}_2 )</th>
<th>( \text{O}_2 )</th>
<th>( \text{Fe}_2\text{O}_3 )</th>
<th>( \text{SO}_2 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>m (g)</td>
<td>168</td>
<td>123</td>
<td>111</td>
</tr>
<tr>
<td>M (g/mol)</td>
<td>120.0</td>
<td>32.0</td>
<td>159.7</td>
</tr>
<tr>
<td>n (mol)</td>
<td>1.40</td>
<td>3.84</td>
<td>0.698</td>
</tr>
</tbody>
</table>

The chemical reaction in this example is of environmental interest. Iron pyrite (\( \text{FeS}_2 \)) is often an impurity in coal, and so burning this fuel in a power plant produces sulfur dioxide (\( \text{SO}_2 \)), a major air pollutant. Our next example also involves burning a fuel and its effect on the atmosphere.

Example \( \PageIndex{4} \): Mass of Oxygen

What mass of oxygen would be consumed when \( 3.3 \times 10^{15} \) g, 3.3 Pg (petagrams), of octane (\( \text{C}_8\text{H}_{18} \)) is burned to produce \( \text{CO}_2 \) and \( \text{H}_2\text{O} \)?

**Solution**

First, write a balanced equation

\[
\text{2C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}
\]

The problem gives the mass of \( \text{C}_8\text{H}_{18} \) burned and asks for the mass of \( \text{O}_2 \) required to combine with it. Thinking the problem through before trying to solve it, we realize that the molar mass of octane could be used to calculate the amount of octane consumed. Then we need a stoichiometric ratio to get the amount of \( \text{O}_2 \) consumed. Finally, the molar mass of \( \text{O}_2 \) permits calculation of the mass of \( \text{O}_2 \). Symbolically

\[
\text{mass of } \text{O}_2 = \text{mass of } \text{C}_8\text{H}_{18} \cdot \frac{25 \text{ mol } \text{O}_2}{1 \text{ mol } \text{C}_8\text{H}_{18}} \cdot \frac{32 \text{ g } \text{O}_2}{1 \text{ mol } \text{O}_2}
\]

\[
\text{mass of } \text{O}_2 = 3.3 \times 10^{15} \text{ g } \text{C}_8\text{H}_{18} \cdot \frac{25}{1} \cdot \frac{32}{1} \text{ g } \text{O}_2
\]

\[
\text{mass of } \text{O}_2 = 2.79 \times 10^{16} \text{ g } \text{O}_2
\]
\[ m_{\text{C}_{8}\text{H}_{18}} \xrightarrow{M_{\text{C}_{8}\text{H}_{18}}} n_{\text{C}_{8}\text{H}_{18}} \xrightarrow{S\text{(SO}_{2}\text{/C}_{8}\text{H}_{18})}} n_{\text{O}_{2}} \xrightarrow{M_{\text{O}_{2}}} m_{\text{O}_{2}} \]

\[
\begin{align*}
m_{\text{O}_{2}} &= 3.3 \cdot 10^{15} \text{ g} \cdot \frac{1 \text{ mol C}_{8}\text{H}_{18}}{114 \text{ g}} \cdot \frac{25 \text{ mol O}_{2}}{2 \text{ mol C}_{8}\text{H}_{18}} \cdot \frac{32 \cdot 0.00 \text{ g}}{1 \text{ mol O}_{2}} \\
&= 1.2 \cdot 10^{16} \text{ g}
\end{align*}
\]

Thus 12 Pg (petagrams) of O\(_2\) would be needed.

The large mass of oxygen obtained in this example is an estimate of how much O\(_2\) is removed from the earth’s atmosphere each year by human activities. Octane, a component of gasoline, was chosen to represent coal, gas, and other fossil fuels. Fortunately, the total mass of oxygen in the air (1.2 \times 10^{21} \text{ g}) is much larger than the yearly consumption. If we were to go on burning fuel at the present rate, it would take about 100 000 years to use up all the O\(_2\). Actually we will consume the fossil fuels long before that! One of the least of our environmental worries is running out of atmospheric oxygen.

---

**Contributors and Attributions**

- Ed Vitz (Kutztown University), **John W. Moore** (UW-Madison), **Justin Shorb** (Hope College), **Xavier Prat-Resina** (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.