Stoichiometry

Section 11.1 What is stoichiometry?

In your textbook, read about stoichiometry and the balanced equation.

For each statement below, write true or false.

1. The study of the quantitative relationships between the amounts of reactants used and the amounts of products formed by a chemical reaction is called stoichiometry.

2. Stoichiometry is based on the law of conservation of mass.

3. In any chemical reaction, the mass of the products is less than the mass of the reactants.

4. The coefficients in a chemical equation represent not only the number of individual particles but also the number of moles of particles.

5. The mass of each reactant and product is related to its coefficient in the balanced chemical equation for the reaction by its molar mass.

Complete the table below, using information represented in the chemical equation for the combustion of methanol, an alcohol.

methanol + oxygen → carbon dioxide + water

\[ 2\text{CH}_3\text{OH(l)} + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O(g)} \]

<table>
<thead>
<tr>
<th>Substance</th>
<th>Molar Mass (g/mol)</th>
<th>Number of Molecules</th>
<th>Number of Moles (mol)</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>6. Methanol</td>
<td>32.05</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>7. Oxygen gas</td>
<td>32.00</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>8. Carbon dioxide</td>
<td>44.01</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>9. Water</td>
<td>18.02</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

10. What are the reactants? __________________________________________

11. What are the products? __________________________________________

12. What is the total mass of the reactants? __________________________

13. What is the total mass of the products? __________________________

14. How do the total masses of the reactants and products compare? ____________________
Section 11.1 continued

In your textbook, read about mole ratios.

Answer the questions about the following chemical reaction.

\[
\text{sodium + iron(III) oxide} \rightarrow \text{sodium oxide + iron} \\
6\text{Na(s)} + \text{Fe}_2\text{O}_3(s) \rightarrow 3\text{Na}_2\text{O(s)} + 2\text{Fe(s)}
\]

15. What is a mole ratio?

______________________________________________________________________________________________
______________________________________________________________________________________________

16. How is a mole ratio written?

______________________________________________________________________________________________
______________________________________________________________________________________________
______________________________________________________________________________________________
______________________________________________________________________________________________

17. Predict the number of mole ratios for this reaction. __________________

18. What are the mole ratios for this reaction?

19. What is the mole ratio relating sodium to iron? __________________

20. What is the mole ratio relating iron to sodium? __________________

21. Which mole ratio has the largest value? _____________________________
Section 11.2 Stoichiometric Calculations

In your textbook, read about mole-to-mole conversion.

Read the following passage and then solve the problems. In the equation that follows each problem, write in the space provided the mole ratio that can be used to solve the problem. Complete the equation by writing the correct value on the line provided.

The reaction of sodium peroxide and water produces sodium hydroxide and oxygen gas. The following balanced chemical equation represents the reaction.

\[ 2\text{Na}_2\text{O}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{NaOH}(\text{s}) + \text{O}_2(\text{g}) \]

1. How many moles of sodium hydroxide are produced when 1.00 mol sodium peroxide reacts with water?

\[ 1.00 \text{ mol Na}_2\text{O}_2 \times = \text{ mol NaOH} \]

2. How many moles of oxygen gas are produced when 0.500 mol Na\textsubscript{2}O\textsubscript{2} reacts with water?

\[ 0.500 \text{ mol Na}_2\text{O}_2 \times = \text{ mol O}_2 \]

3. How many moles of sodium peroxide are needed to produce 1.00 mol sodium hydroxide?

\[ 1.00 \text{ mol NaOH} \times = \text{ mol Na}_2\text{O}_2 \]

4. How many moles of water are required to produce 2.15 mol oxygen gas in this reaction?

\[ 2.15 \text{ mol O}_2 \times = \text{ mol H}_2\text{O} \]

5. How many moles of water are needed for 0.100 mol of sodium peroxide to react completely in this reaction?

\[ 0.100 \text{ mol Na}_2\text{O}_2 \times = \text{ mol H}_2\text{O} \]

6. How many moles of oxygen are produced if the reaction produces 0.600 mol sodium hydroxide?

\[ 0.600 \text{ mol NaOH} \times = \text{ mol O}_2 \]
Section 11.2  continued

In your textbook, read about mole-to-mass and mass-to-mass conversions.

Solving a mass-to-mass problem requires the four steps listed below. The equations in the boxes show how the four steps are used to solve an example problem. After you have studied the example, solve the problems below, using the four steps.

Example problem: How many grams of carbon dioxide are produced when 20.0 g acetylene (C₂H₂) is burned?

<table>
<thead>
<tr>
<th>Step 1</th>
<th>Write a balanced chemical equation for the reaction.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Step 2</td>
<td>Determine the number of moles of the known substance, using mass-to-mole conversion.</td>
</tr>
<tr>
<td>Step 3</td>
<td>Determine the number of moles of the unknown substance, using mole-to-mole conversion.</td>
</tr>
<tr>
<td>Step 4</td>
<td>Determine the mass of the unknown substance, using mole-to-mass conversion.</td>
</tr>
</tbody>
</table>

Solution

\[
\begin{align*}
2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) & \rightarrow 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g) \\
20.0 \text{ g C}_2\text{H}_2 & \rightarrow 1 \text{ mol C}_2\text{H}_2 \\
& \rightarrow \frac{1 \text{ mol C}_2\text{H}_2}{26.04 \text{ g C}_2\text{H}_2} \\
& = 0.768 \text{ mol C}_2\text{H}_2 \\
0.768 \text{ mol C}_2\text{H}_2 & \times \frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_2} \\
& = 1.54 \text{ mol CO}_2 \\
1.54 \text{ mol CO}_2 & \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \\
& = 67.8 \text{ g CO}_2
\end{align*}
\]
7. In some mole-to-mass conversions, the number of moles of the known substance is given. In those conversions, which step of the above solution is not necessary? ______________________

8. In a blast furnace, iron and carbon monoxide are produced from the reaction of iron(III) oxide (Fe$_2$O$_3$) and carbon. How many grams of iron are formed when 150 g iron(III) oxide reacts with an excess of carbon?

9. Solid sulfur tetrafluoride (SF$_4$) and water react to form sulfur dioxide and an aqueous solution of hydrogen fluoride. How many grams of water are necessary for 20.0 g sulfur tetrafluoride to react completely?
Section 11.3 Limiting Reactants

In your textbook, read about why reactions stop and how to determine the limiting reactant.

Study the diagram showing a chemical reaction and the chemical equation that represents the reaction. Then complete the table. Show your calculations for questions 25–27 in the space below the table.

\[
\text{O}_2 + 2\text{NO} \rightarrow 2\text{NO}_2
\]

The molar masses of \( \text{O}_2 \), \( \text{NO} \), and \( \text{NO}_2 \) are 32.00 g/mol, 30.01 g/mol, and 46.01 g/mol, respectively.

<table>
<thead>
<tr>
<th>Amount of ( \text{O}_2 )</th>
<th>Amount of ( \text{NO} )</th>
<th>Amount of ( \text{NO}_2 )</th>
<th>Limiting Reactant</th>
<th>Amount and Name of Excess Reactant</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 molecule</td>
<td>2 molecules</td>
<td>2 molecules</td>
<td>none</td>
<td>none</td>
</tr>
<tr>
<td>4 molecules</td>
<td>4 molecules</td>
<td>4 molecules</td>
<td>NO</td>
<td>2 molecules ( \text{O}_2 )</td>
</tr>
<tr>
<td>2 molecules</td>
<td>8 molecules</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1.00 mol</td>
<td>2.00 mol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4.00 mol</td>
<td>4.00 mol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>5.00 mol</td>
<td>7.00 mol</td>
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<td></td>
<td></td>
</tr>
<tr>
<td>1.00 mol</td>
<td>4.00 mol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.500 mol</td>
<td>0.200 mol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>32.00 g</td>
<td>60.02 g</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>16.00 g</td>
<td>80.00 g</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>10.00 g</td>
<td>20.00 g</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Section 11.4 Percent Yield

In your textbook, read about the yields of products.

Study the diagram and the example problem.

\[
\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% 
\]

Example Problem: The following chemical equation represents the production of gallium oxide, a substance used in the manufacturing of some semiconductor devices.

\[4\text{Ga}(s) + 3\text{O}_2(g) \rightarrow 2\text{Ga}_2\text{O}_3(s)\]

In one experiment, the reaction yielded 7.42 g of the oxide from a 7.00-g sample of gallium. Determine the percent yield of this reaction. The molar masses of Ga and Ga\(_2\)O\(_3\) are 69.72 g/mol and 187.44 g/mol, respectively.

Use the information in the diagram and example problem to evaluate each value or expression below. If the value or expression is correct, write correct. If it is incorrect, write the correct value or expression.

1. actual yield: unknown

2. mass of reactant: 7.00 g Ga

3. number of moles of reactant: \(7.00 \text{ g Ga} \times \frac{69.72 \text{ g Ga}}{1 \text{ mol Ga}}\)

4. number of moles of product: \(0.100 \text{ mol Ga} \times \frac{2 \text{ mol Ga}_2\text{O}_3}{1 \text{ mol Ga}}\)

5. theoretical yield: \(0.0500 \text{ mol Ga}_2\text{O}_3 \times \frac{187.44 \text{ g Ga}_2\text{O}_3}{1 \text{ mol Ga}_2\text{O}_3}\)

6. percent yield: \(\frac{9.37 \text{ g Ga}_2\text{O}_3}{7.42 \text{ g Ga}_2\text{O}_3} \times 100\%\)
Study Guide - Chapter 11 – Stoichiometry

Section 11.1 What is stoichiometry?

1. true
2. true
3. false
4. true
5. true
6. 2, 2, 64.10
7. 3, 3, 96.00
8. 2, 2, 88.02
9. 4, 4, 72.08
10. methanol and oxygen gas
11. carbon dioxide and water
12. 160.10 g
13. 160.10 g
14. They are equal.
15. A mole ratio is a ratio between the numbers of moles of any two substances in a balanced chemical equation.
16. A mole ratio is written for two substances in a balanced chemical equation as a fraction by placing the number of moles of one substance in the numerator and the number of moles of another substance in the denominator.
17. 12
18.

<table>
<thead>
<tr>
<th>Fe₂O₃</th>
<th>Na</th>
<th>Na₂O</th>
<th>Fe</th>
<th>H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol</td>
<td>6 mol</td>
<td>6 mol Na</td>
<td>2 mol Fe</td>
<td>6 mol Na</td>
</tr>
<tr>
<td>6 mol Na</td>
<td>1 mol Fe₂O₃</td>
<td>3 mol Na₂O</td>
<td>2 mol Fe</td>
<td>3 mol Fe₂O₃</td>
</tr>
<tr>
<td>6 mol Na</td>
<td>5 mol Na₂O</td>
<td>1 mol Fe₂O₃</td>
<td>3 mol Na₂O</td>
<td>2 mol Fe</td>
</tr>
<tr>
<td>6 mol Na</td>
<td>2 mol Fe</td>
<td>1 mol Fe₂O₃</td>
<td>2 mol Fe</td>
<td>3 mol Na₂O</td>
</tr>
</tbody>
</table>

19. 6 mol Na/2 mol Fe
20. 2 mol Fe/6 mol Na
21. 6 mol Na/1 mol Fe₂O₃
<table>
<thead>
<tr>
<th>Amount of O₂</th>
<th>Amount of NO</th>
<th>Amount of NO₂</th>
<th>Limiting Reactant</th>
<th>Excess Reactant</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 molecule</td>
<td>2 molecules</td>
<td>2 molecules</td>
<td>none</td>
<td>none</td>
</tr>
<tr>
<td>4 molecules</td>
<td>4 molecules</td>
<td>4 molecules</td>
<td>NO</td>
<td>2 molecules O₂</td>
</tr>
<tr>
<td>2 molecules</td>
<td>8 molecules</td>
<td>1. 4 molecules</td>
<td>2. O₂</td>
<td>3. 4 molecules NO</td>
</tr>
<tr>
<td>1.00 mol</td>
<td>2.00 mol</td>
<td>4. 2.00 mol</td>
<td>5. none</td>
<td>6. none</td>
</tr>
<tr>
<td>4.00 mol</td>
<td>4.00 mol</td>
<td>7. 4.00 mol</td>
<td>8. NO</td>
<td>9. 2.00 mol O₂</td>
</tr>
<tr>
<td>5.00 mol</td>
<td>7.00 mol</td>
<td>10. 7.00 mol</td>
<td>11. NO</td>
<td>12. 1.50 mol O₂</td>
</tr>
<tr>
<td>1.00 mol</td>
<td>4.00 mol</td>
<td>13. 2.00 mol</td>
<td>14. O₂</td>
<td>15. 2.00 mol NO</td>
</tr>
<tr>
<td>0.500 mol</td>
<td>0.200 mol</td>
<td>16. 0.200 mol</td>
<td>17. NO</td>
<td>18. 0.400 mol O₂</td>
</tr>
<tr>
<td>32.00 g</td>
<td>60.02 g</td>
<td>19. 92.02 g</td>
<td>20. none</td>
<td>21. none</td>
</tr>
<tr>
<td>16.00 g</td>
<td>80.00 g</td>
<td>22. 46.01 g</td>
<td>23. O₂</td>
<td>24. 50.12 g NO</td>
</tr>
<tr>
<td>10.00 g</td>
<td>20.00 g</td>
<td>25. 28.76 g</td>
<td>26. O₂</td>
<td>27. 1.24 g NO</td>
</tr>
</tbody>
</table>

0.370 mol H₂O = 18.02 g H₂O/1 mol H₂O
balanced equation mole ratio = 2 mol NO/1 mol O₂
10.00 g O₂/1 mol O₂ = 0.3125 mol O₂
20.00 g NO/1 mol NO/30.01 = 0.6664 mol NO
actual mole ratio = 0.6664 mol NO/0.3125 mol O₂ = 2.132 mol NO/1.000 mol O₂

Because the actual mole ratio of NO:O₂ is larger than the balanced equation mole ratio of NO:O₂, there is an excess of NO; O₂ is the limiting reactant.

Mass of NO used = 0.3125 mol O₂ × 2 mol NO/1 mol O₂ = 0.6250 mol NO
0.6250 mol NO × 30.01 g NO/1 mol NO = 18.76 g NO
Mass of NO₂ produced = 0.6250 mol NO × 46.01 g NO₂/1 mol NO = 28.76 g NO₂
Excess NO = 20.00 g NO − 18.76 g NO = 1.24 g NO

Section 11.4 Percent Yield
1. 7.42 g Ga₂O₃
2. correct
3. 7.00 g Ga × 1 mol Ga/69.72 g Ga
4. 0.100 mol Ga × 2 mol Ga₂O₃/4 mol Ga
5. correct
6. 7.42 Ga₂O₃/9.37 g Ga₂O₃ × 100