I. Stoichiometry
1) Given the balanced equation \(2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)\), how many moles of \(\text{O}_2\) are produced from twelve moles of \(\text{KClO}_3\)?

\[
12 \text{ mol KClO}_3 \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = 18 \text{ mol O}_2
\]

2) Using the equation from problem 1, how many moles of \(\text{O}_2\) are produced from 14 moles of \(\text{KCl}\)?

\[
14 \text{ mol KCl} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KCl}} = 21 \text{ mol O}_2
\]

3) Calculate the number of grams of \(\text{NO}_2\) that are produced from 4 moles of \(\text{NO}\) given the equation \(2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)\).

\[
4 \text{ mol NO} \times \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}} \times \frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} = 184.04 \text{ g NO}_2
\]

4) Calculate the mass of \(\text{O}_2\) produced from the decomposition of 75.0 g of \(\text{KClO}_3\) given the equation \(2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)\).

\[
75 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \times \frac{32 \text{ g O}_2}{1 \text{ mol O}_2} = 29.38 \text{ g O}_2
\]

5) Calculate the mass of \(\text{Ag}\) needed to react with \(\text{Cl}_2\) to produce 84 g of silver chloride. The equation for this reaction is \(2\text{Ag} + \text{Cl}_2 \rightarrow 2\text{AgCl}\).

\[
84 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.36 \text{ g AgCl}} \times \frac{2 \text{ mol Ag}}{2 \text{ mol AgCl}} \times \frac{107.91 \text{ g Ag}}{1 \text{ mol Ag}} = 63.23 \text{ g Ag}
\]

6) How many liters of carbon monoxide at STP are needed to react with 4.80 g of \(\text{O}_2\) to produce \(\text{CO}_2\)? The equation for this reaction is \(2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g)\).

\[
4.80 \text{ g O}_2 \times \frac{2 \text{ mol CO}}{32 \text{ g O}_2} \times \frac{22.4 \text{ L CO}}{1 \text{ mol CO}} = 6.72 \text{ L CO}
\]

7) A volume of 7.5 L of hydrogen gas at STP was produced from the single-replacement reaction of zinc with nitric acid (\(\text{Zn} + 2\text{HNO}_3 \rightarrow \text{Zn(NO}_3)_2 + \text{H}_2\)). Calculate the mass of \(\text{Zn}\) needed for this reaction.

\[
7.5 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 21.89 \text{ g Zn}
\]
8) Calculate the number of liters of oxygen gas needed to produce 15.0 liters of N\(_2\)O\(_3\). Assume all gases are at the same conditions of temperature and pressure and react according to this balanced equation: 2N\(_2\)(g) + 3O\(_2\)(g) \rightarrow 2N\(_2\)O\(_3\)(g).

\[
15.0 \text{ L } N_2O_3 \times \frac{1 \text{ mol } N_2O_3}{2 \text{ mol } N_2O_3} \times \frac{3 \text{ mol } O_2}{1 \text{ mol } O_2} = 22.5 \text{ L } O_2
\]

----or----

\[
15.0 \text{ L } N_2O_3 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } N_2O_3} = 22.5 \text{ L } O_2
\]

9) Given the balanced equation 2H\(_2\) + O\(_2\) \rightarrow 2H\(_2\)O, how many molecules of water are produced from 2.0 \(\times\) 10\(^{23}\) molecules of oxygen?

\[
2.0 \times 10^{23} \text{ molecules } O_2 \times \frac{2 \text{ molecules } H_2O}{1 \text{ molecule } O_2} = 4.0 \times 10^{23} \text{ molecules } H_2O
\]

II. Limiting Reagent and Percent Yield

10) How many moles of water can be made from 4 moles of oxygen gas and 5 moles of hydrogen gas? What is the limiting reagent? (5 mol H\(_2\)O, LR is H\(_2\))

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<thead>
<tr>
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<th>End</th>
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</thead>
<tbody>
<tr>
<td>2H(_2) + O(_2)</td>
<td>2H(_2)O</td>
</tr>
<tr>
<td>5 mol</td>
<td>4 mol</td>
</tr>
<tr>
<td>9.38 mol</td>
<td>0 mol</td>
</tr>
<tr>
<td>-2.5 mol</td>
<td></td>
</tr>
<tr>
<td>1.5 mol</td>
<td>5 mol</td>
</tr>
</tbody>
</table>

11) Calculate the mass of water produced from the reaction of 50 g of H\(_2\) and 300 g of O\(_2\). What is the limiting reagent? (337.88 g H\(_2\)O, LR is O\(_2\))

\[
50 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.02 \text{ g } H_2} = 24.75 \text{ mol } H_2 \quad 300 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32 \text{ g } O_2} = 9.38 \text{ mol } O_2
\]

<table>
<thead>
<tr>
<th>Start</th>
<th>End</th>
</tr>
</thead>
<tbody>
<tr>
<td>2H(_2) + O(_2)</td>
<td>2H(_2)O</td>
</tr>
<tr>
<td>24.75 mol</td>
<td>9.38 mol</td>
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<td>-18.75 mol</td>
<td></td>
</tr>
<tr>
<td>6 mol</td>
<td>18.75 mol</td>
</tr>
</tbody>
</table>

\[
18.75 \text{ mol } H_2O \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 337.88 \text{ g } H_2O
\]
Chemistry--Unit 5: Stoichiometry

Practice Problems

12) The burning of 18.0 g of carbon produces 55.0 g of carbon dioxide according to the reaction \( \text{C} + \text{O}_2 \rightarrow \text{CO}_2 \). What is the theoretical yield of \( \text{CO}_2 \)? Calculate the percent yield of \( \text{CO}_2 \).

\[
18.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 65.96 \text{ g CO}_2
\]

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \quad \% \text{ yield CO}_2 = \frac{55 \text{ g CO}_2}{65.96 \text{ g CO}_2} \times 100 \quad \% \text{ yield CO}_2 = 83.38\%
\]

13) Calculate the percent yield of \( \text{Cl}_2(g) \) in the electrolytic decomposition of hydrogen chloride if 25.8 g of \( \text{HCl} \) produces 13.6 g of chlorine gas. The balanced equation for this reaction is \( 2\text{HCl} \rightarrow \text{H}_2 + \text{Cl}_2 \).

\[
25.8 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol HCl}} \times \frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} = 25.09 \text{ g Cl}_2
\]

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \quad \% \text{ yield Cl}_2 = \frac{13.6 \text{ g Cl}_2}{25.09 \text{ g Cl}_2} \times 100 \quad \% \text{ yield Cl}_2 = 54.22\%
\]

14) One method for reclaiming silver metal from silver chloride results in a 94.6% yield. Calculate the actual mass of silver that can be produced in this reaction if 100.0 g of \( \text{AgCl} \) is converted to silver metal. \( 2\text{AgCl(s)} \rightarrow 2\text{Ag(s)} + \text{Cl}_2(g) \)

\[
100 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.36 \text{ g AgCl}} \times \frac{2 \text{ mol Ag}}{2 \text{ mol AgCl}} \times \frac{107.91 \text{ g Ag}}{1 \text{ mol Ag}} = 75.27 \text{ g Ag}
\]

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \quad 94.6\% = \frac{\text{actual g Ag}}{75.27 \text{ g Ag}} \times 100 \quad \text{actual g Ag} = 71.21 \text{ g Ag}
\]

15) What is the actual amount of \( \text{MgO} \) produced when excess \( \text{CO}_2 \) reacts with 42.8 g of \( \text{Mg(s)} \)? The percent yield of \( \text{MgO(s)} \) for this reaction is 81.7%. \( 2\text{Mg(s)} + \text{CO}_2(g) \rightarrow 2\text{MgO(s)} + \text{C(s)} \)

\[
42.8 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{2 \text{ mol MgO}}{2 \text{ mol Mg}} \times \frac{40.31 \text{ g MgO}}{1 \text{ mol MgO}} = 70.97 \text{ g MgO}
\]

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \quad 81.7\% = \frac{\text{actual g MgO}}{70.97 \text{ g MgO}} \times 100 \quad \text{actual g MgO} = 57.98 \text{ g MgO}
\]
Chemistry--Unit 5: Stoichiometry
Practice Problems

III. Titrations

16) Titration reveals that 11.6 mL of 3.0 M sulfuric acid, $H_2SO_4$, are required to neutralize the sodium hydroxide in 25.00 mL of NaOH solution. What is the molarity of the NaOH solution? The titration reaction is $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2HOH(l)$.

$$11.6 \text{ mL } H_2SO_4 \times \frac{1 \text{ L } H_2SO_4}{1000 \text{ mL } H_2SO_4} \times \frac{3 \text{ mol } H_2SO_4}{1 \text{ L } H_2SO_4} \times \frac{2 \text{ mol } NaOH}{1 \text{ mol } H_2SO_4} = 0.0696 \text{ mol } NaOH$$

$$\frac{0.0696 \text{ mol } NaOH}{0.025 \text{ L } NaOH} = 2.78 \text{ M } NaOH$$

17) A titration of 10.0 mL of HCl requires 15.5 mL of 0.50 M NaOH. What is the molarity of the HCl? The titration reaction is HCl$(aq) + NaOH(aq) \rightarrow NaCl(aq) + HOH(l)$.

$$15.5 \text{ mL } NaOH \times \frac{1 \text{ L } NaOH}{1000 \text{ mL } NaOH} \times \frac{0.5 \text{ mol } NaOH}{1 \text{ L } NaOH} \times \frac{1 \text{ mol } HCl}{1 \text{ mol } NaOH} = 0.00775 \text{ mol } HCl$$

$$\frac{0.00775 \text{ mol } HCl}{0.010 \text{ L } HCl} = 0.775 \text{ M } HCl$$

18) If 48.9 mL of 0.750 M HCl are needed in a titration to neutralize a solution of NaOH, how many grams of solid NaOH were dissolved in water to make the solution? The titration reaction is HCl$(aq) + NaOH(aq) \rightarrow NaCl(aq) + HOH(l)$.

$$48.9 \text{ mL } HCl \times \frac{1 \text{ L } HCl}{1000 \text{ mL } HCl} \times \frac{0.75 \text{ mol } HCl}{1 \text{ L } HCl} \times \frac{1 \text{ mol } NaOH}{1 \text{ mol } HCl} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 1.47 \text{ g NaOH}$$

19) A solid acid with a molar mass of 56.98 g/mol can be used to titrate an unknown NaOH solution. If $5.48 \times 10^{-3}$ g of this acid are used, what volume, in mL, of 0.00250 M NaOH would be needed to reach the equivalence point? Assume a mole:mole ratio of acid:base to be 2:1.

$$5.48 \times 10^{-3} \text{ g acid} \times \frac{1 \text{ mol acid}}{56.98 \text{ g acid}} \times \frac{1 \text{ mol } NaOH}{2 \text{ mol acid}} \times \frac{1 \text{ L } NaOH}{0.00250 \text{ mol } NaOH} \times \frac{1000 \text{ mL } NaOH}{1 \text{ L } NaOH} = 19.23 \text{ mL } NaOH$$