3.79 Define limiting reagent and excess reagent.
Limiting reagent is the reactant that is depleted; excess reagent is a reagent that is not depleted.

What is the significance of the limiting reagent in predicting the amount of the product obtained in a reaction? In a reaction, since the reaction stops when one reactant is depleted, the amount of limiting reagent determines the amount of each product possible.

Can there be a limiting reagent if only one reactant is present? If there is only one reactant, then technically it must be the limiting reagent since it will be depleted.

3.82 \( \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \)
From diagram, one sees that there are 3 moles of \( \text{N}_2 \) and 10 moles of \( \text{H}_2 \)

Fill in the “ICF” table to help answer all questions:

<table>
<thead>
<tr>
<th></th>
<th>( \text{N}_2 )</th>
<th>( 3\text{H}_2 )</th>
<th>( 2\text{NH}_3 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>3 mc</td>
<td>10 mc</td>
<td>0 mc</td>
</tr>
<tr>
<td>Change</td>
<td>−3 mc</td>
<td>−9 mc</td>
<td>+6 mc</td>
</tr>
<tr>
<td>Final</td>
<td>0 mc</td>
<td>1 mc</td>
<td>6 mc</td>
</tr>
</tbody>
</table>

What is the limiting reagent? Explain. There are 3 molecules of \( \text{N}_2 \) in the diagram. We need 9 molecules of \( \text{H}_2 \) to react completely with 3 molecules of \( \text{N}_2 \). Since there are 10 molecules of \( \text{H}_2 \), \( \text{H}_2 \) is in excess.

Number of moles of product produced: When 3 molecules of \( \text{N}_2 \) react, 6 molecules of \( \text{NH}_3 \) will be produced.

Number of moles of excess reagent left: 9 molecules of \( \text{H}_2 \) will react with 3 molecules of \( \text{N}_2 \), leaving 1 molecule of \( \text{H}_2 \) in excess.

3.83 Nitric oxide reacts with oxygen gas to form nitrogen dioxide, a dark-brown gas:
\( 2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \)
In one experiment 0.886 mole of NO is mixed with 0.503 mole of \( \text{O}_2 \). Calculate which of the two reactants is the limiting reagent. Calculate also the number of moles of \( \text{NO}_2 \) produced.

\[
\frac{0.886 \text{ mol NO}}{2} = 0.443 \text{ mol Rxn}; \quad \frac{0.503 \text{ mol O}_2}{1} = 0.503 \text{ mol Rxn}; \quad \text{NO is the limiting reagent}
\]

\[
\text{mol NO}_2 = 0.443 \text{ mol Rxn} \times \frac{2 \text{ mol NO}_2}{1 \text{ mol Rxn}} = 0.886 \text{ mol NO}_2
\]

Here is an ICF table, to see the information in an organized fashion:

<table>
<thead>
<tr>
<th></th>
<th>2 NO</th>
<th>+</th>
<th>0.503 mol O₂</th>
<th>→</th>
<th>2 NO₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.886 mol</td>
<td>0.503 mol</td>
<td>0 mol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>−0.886 mol</td>
<td>−0.443 mol</td>
<td>+0.886 mol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final</td>
<td>0 mol</td>
<td>0.060 mol</td>
<td>0.886 mol</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

3.88 Why is the actual yield of a reaction almost always smaller than the theoretical yield? The actual yield of a reaction is usually less than the theoretical yield because not all reactants form products (reversible reactions or side products), and not all product may be recovered.
3.84 The depletion of the ozone in the stratosphere has been a matter of great concern among scientists in recent years. \( \text{O}_3 + \text{NO} \rightarrow \text{O}_2 + \text{NO}_2 \) If 0.740 g of \( \text{O}_3 \) reacts with 0.679 g of \( \text{NO} \), how many grams of \( \text{NO}_2 \) produced?

\[
\frac{0.740 \text{ g O}_3 \times \frac{1 \text{ mol O}_3}{48.00 \text{ g O}_3}}{1} = 0.0154 \text{ mol Rxn} ; \quad \frac{0.679 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}}}{1} = 0.0226 \text{ mol Rxn}
\]

Therefore \( \text{O}_3 \) is the limiting reagent.

\[
? \text{ g NO}_2 = 0.0154 \text{ mol Rxn} \times \frac{1 \text{ mol NO}_2}{1 \text{ mol Rxn}} = 0.0154 \text{ mol NO}_2
\]

\[
	ext{mol NO remaining} = 0.0226 \text{ mol NO} - 0.0154 \text{ mol NO} = 0.0072 \text{ mol NO}
\]

3.92 Ethylene \((\text{C}_2\text{H}_4)\), can be prepared by heating hexane \((\text{C}_6\text{H}_{14})\) at 800°C: \( \text{C}_6\text{H}_{14} \rightarrow \text{C}_2\text{H}_4 + \text{other products} \). If the yield of ethylene production is 42.5 percent, what mass of hexane must be reacted to produce 481 g of ethylene?

\[
42.5\% \text{ yield} = \frac{481 \text{ g C}_2\text{H}_4}{\text{theoretical yield}} \times 100\% ; \quad \text{theoretical yield C}_2\text{H}_4 = 1.13 \times 10^3 \text{ g C}_2\text{H}_4
\]

\[
1.13 \times 10^3 \text{ g C}_2\text{H}_4 \times \frac{1 \text{ mol C}_2\text{H}_4}{28.05 \text{ g C}_2\text{H}_4} \times \frac{1 \text{ mol C}_6\text{H}_{14}}{1 \text{ mol C}_2\text{H}_4} = \frac{86.15 \text{ g C}_6\text{H}_{14}}{1 \text{ mol C}_6\text{H}_{14}} = \frac{3.47 \times 10^3 \text{ g C}_6\text{H}_{14}}{1 \text{ mol C}_6\text{H}_{14}}
\]

3.94 Disulfur dichloride is used in the vulcanization of rubber. It is prepared by heating sulfur in an atmosphere of chlorine. What is the theoretical yield of \( \text{S}_2\text{Cl}_2 \) when 4.06 g of \( \text{S}_8 \) is heated with 6.24 g of \( \text{Cl}_2 \), \( \text{S}_8(l) + 4\text{Cl}_2(g) \rightarrow 4\text{S}_2\text{Cl}_2(l) \)?

\[
\frac{4.06 \text{ g S}_8 \times \frac{1 \text{ mol S}_8}{256.6 \text{ g S}_8}}{1} = 0.0158 \text{ mol Rxn} ; \quad \frac{6.24 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2}}{4} = 0.0220 \text{ mol Rxn}
\]

\[
? \text{ g S}_2\text{Cl}_2 = 0.0158 \text{ mol Rxn} \times \frac{4 \text{ mol S}_2\text{Cl}_2}{1 \text{ mol Rxn}} \times \frac{135.04 \text{ g S}_2\text{Cl}_2}{1 \text{ mol S}_2\text{Cl}_2} = \frac{8.55 \text{ g S}_2\text{Cl}_2}{1 \text{ mol S}_2\text{Cl}_2}
\]

If the actual yield of \( \text{S}_2\text{Cl}_2 \) is 6.55 g, what is the percent yield?

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{6.55 \text{ g}}{8.55 \text{ g}} \times 100\% = \frac{76.6\%}{1 \text{ mol S}_2\text{Cl}_2}
\]

3.108 \( \text{Fe}_2\text{O}_3 + 3 \text{ CO} \rightarrow 2 \text{ Fe} + 3 \text{ CO}_2 \)

One obtains \( 1.64 \times 10^3 \) kg Fe from \( 2.62 \times 10^3 \) kg of an impure sample of \( \text{Fe}_2\text{O}_3 \). What is % purity of sample? [Hint: Find the mass of pure \( \text{Fe}_2\text{O}_3 \) needed to produce the amount of Fe obtained]

\[
1.64 \times 10^3 \text{ kg Fe} \times \frac{\frac{1 \text{ kg-mol Fe}}{55.85 \text{ kg Fe}} \times \frac{1 \text{ kg-mol Fe}_2\text{O}_3}{2 \text{ kg-mol Fe}} \times \frac{159.7\text{ kg Fe}_2\text{O}_3}{1 \text{ kg-mol Fe}_2\text{O}_3}}{1} = 2.34 \times 10^3 \text{ kg Fe}_2\text{O}_3 \text{ (pure)}
\]

\[
\text{percent purity Fe}_2\text{O}_3 = \frac{2.34 \times 10^3 \text{ kg pure Fe}_2\text{O}_3}{2.62 \times 10^3 \text{ kg impure Fe}_2\text{O}_3} \times 100\% = \frac{89.6\%}{1 \text{ mol Fe}_2\text{O}_3}
\]