Stoichiometry: Calculations with Chemical Formulas and Equations (Chapter 3)
Past Quiz and Test Questions – Answer Key
My answers are in bold faced italics and underlined – to this point I have only included answers up to question 9.

1. Balance the following equations

   \( \text{Al}_2(\text{SO}_4)_3 \text{ (aq)} + 3 \text{Ba(NO}_3)_2 \text{ (aq)} \rightarrow 2 \text{Al(NO}_3)_3 \text{ (aq)} + 3 \text{BaSO}_4 \text{ (s)} \)

   \( \text{C}_3\text{H}_8\text{O}_2 \text{ (R)} + 4 \text{O}_2 \text{ (g)} \rightarrow 3 \text{CO}_2 \text{ (g)} + 4 \text{H}_2\text{O} \text{ (g)} \)

   \( \text{P}_2\text{O}_5 \text{ (g)} + 3 \text{H}_2\text{O} \text{ (R)} \rightarrow 2 \text{H}_3\text{PO}_4 \text{ (aq)} \)

   \( \text{B}_{10}\text{H}_{18} + 12 \text{O}_2 \rightarrow 5 \text{B}_2\text{O}_3 + 9 \text{H}_2\text{O} \)

   \( \text{N}_2\text{ (g)} + 3 \text{H}_2 \text{ (g)} \rightarrow 2 \text{NH}_3 \text{ (g)} \)

2. a. How many grams of Cl are in 113.5-g of C\(_2\)Cl\(_6\)?

   \[ \text{gCl} = \frac{113.5 \text{gC}_2\text{Cl}_6}{1} \times \frac{213 \text{gCl}}{237 \text{gC}_2\text{Cl}_6} = 102.0 \text{gCl} \]

   b. How many atoms of O are in 45.0-g of C\(_{12}\)H\(_{22}\)O\(_{11}\)?

   \[ \text{atomsO} = \frac{45.0 \text{gC}_{12}\text{H}_{22}\text{O}_{11}}{1} \times \frac{11 \times 6.022 \times 10^{23} \text{atomsO}}{342 \text{gC}_{12}\text{H}_{22}\text{O}_{11}} = 8.72 \times 10^{23} \text{atomsO} \]

   c. How many moles of CaBr\(_2\) are required to obtain 4.5 \times 10^{21} \text{ atoms of Br}?

   \[ \text{molCaBr}_2 = \frac{4.5 \times 10^{21} \text{atomsBr}}{1} \times \frac{1 \text{molCaBr}_2}{2 \times 6.022 \times 10^{23} \text{atomsBr}} = 3.7 \times 10^{-3} \text{molCaBr}_2 \]

   d. How many grams of C are required to make 1.35-mol of CH\(_4\)?

   \[ \text{gC} = \frac{1.35 \text{molCH}_4}{1} \times \frac{12 \text{gC}}{1 \text{molCH}_4} = 16.2 \text{gC} \]

   e. How many moles of NH\(_3\) are contained in 250.0-g of NH\(_3\)?

   \[ \text{molNH}_3 = \frac{250.0 \text{gNH}_3}{1} \times \frac{1 \text{molNH}_3}{17 \text{gNH}_3} = 14.7 \text{molNH}_3 \]
3. Consider the following chemical equation.

Given: 20.0 g K₂CO₃  30.0 g AlBr₃

\[ 3 \text{ K}_2\text{CO}_3 (\text{aq}) + 2 \text{ AlBr}_3 (\text{aq}) \rightarrow 6 \text{ KBr (aq)} + \text{ Al}_2(\text{CO}_3)_3 (s) \]

Stoichiometry:

<table>
<thead>
<tr>
<th></th>
<th>3 mol K₂CO₃</th>
<th>2 mol AlBr₃</th>
<th>6 mol KBr</th>
<th>1 mol Al₂(CO₃)₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>K₂CO₃</td>
<td>3 x 138 g</td>
<td>2 x 267 g</td>
<td>6 x 119 g</td>
<td>1 x 234 g</td>
</tr>
<tr>
<td>AlBr₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

20.0-g of K₂CO₃ are reacted with 30.0-g of AlBr₃.

a. Balance the equation.
b. Which of the reactants is the limiting reactant?

\[
\frac{20.0 g \text{K}_2\text{CO}_3}{3 x 138 g \text{K}_2\text{CO}_3} = 0.0483 \quad \frac{30.0 g \text{AlBr}_3}{2 x 267 g \text{AlBr}_3} = 0.0562
\]

Since the stoichiometric ratio involving K₂CO₃ is less than that involving AlBr₃, K₂CO₃ is the limiting reactant.

c. How many grams of Al₂(CO₃)₃ can be formed?

\[
? g \text{Al}_2(\text{CO}_3)_3 = \frac{20.0 g \text{K}_2\text{CO}_3 \times 1 x 234 g \text{Al}_2(\text{CO}_3)_3}{3 x 138 g \text{K}_2\text{CO}_3} = 11.3 g \text{Al}_2(\text{CO}_3)_3
\]

Note this calculation is based on the limiting reactant, not the excess.

d. Suppose the reaction is carried out and 7.5-g of Al₂(CO₃)₃ are formed. What is the percent yield?

\[
\% \text{yield} = \frac{7.5 g \text{Al}_2(\text{CO}_3)_3 \text{Produced}}{11.3 g \text{Al}_2(\text{CO}_3)_3 \text{Calculated}} \times 100\% = 66.4\% \text{yield}
\]

e. How much of the excess reactant remains after the reaction if all of the limiting reactant is consumed?

First, figure out how much of the excess reactant, AlBr₃, was used.

\[
? g \text{AlBr}_3 \text{used} = \frac{20.0 g \text{K}_2\text{CO}_3}{1} \times \frac{2 x 267 g \text{AlBr}_3}{3 x 138 g \text{K}_2\text{CO}_3} = 25.8 g \text{AlBr}_3 \text{used}
\]

Excess AlBr₃ = 30.0 g – 25.8 g = 4.2 g of excess reactant left over after the reaction.
4. Determine the empirical and molecular formulas for epinephrine, a hormone secreted into the blood stream during times of danger or stress. Its elemental composition is 59.0% C, 7.1% H, 26.2% O, and 7.7% N. Its molecular weight is 183.

\[
\begin{align*}
\text{C}: & \quad \frac{59.0}{12} = 4.92 \quad \text{Divide by smallest}(0.55): \quad \frac{4.92}{0.55} = 8.95 \\
\text{H}: & \quad \frac{7.1}{1} = 7.1 \quad \frac{7.1}{0.55} = 12.91 \\
\text{O}: & \quad \frac{26.2}{16} = 1.6375 \quad \frac{1.6375}{0.55} = 2.98 \\
\text{N}: & \quad \frac{7.7}{14} = 0.55 \quad \frac{0.55}{0.55} = 1
\end{align*}
\]

Since these are all close to an integer, round to the nearest integer for a ratio of: 9:13:3:1 or \( \text{C}_9\text{H}_{13}\text{O}_3\text{N} \) for the empirical formula. Since the formula weight of the empirical formula is 183 and matches the molecular weight, the molecular formula is also \( \text{C}_9\text{H}_{13}\text{O}_3\text{N} \).

5. Balance each of the following equations.

\[
\begin{align*}
\text{C}_{12}\text{H}_{22}\text{O}_{11} \text{ (s) } & + 12 \text{ O}_2 \quad \text{----> } 12 \text{ CO}_2 \text{ (g) } + 11 \text{ H}_2\text{O} \text{ (R)} \\
\text{Mg} \text{ (s) } & + 2 \text{ HCl} \text{ (aq) } \text{----> } \text{MgCl}_2 \text{ (aq) } + \text{ H}_2 \text{ (g)} \\
\text{Na}_2\text{CO}_3 \text{ (aq) } + 2 \text{ HNO}_3 \text{ (aq) } & \text{----> } 2 \text{ NaNO}_3 \text{ (aq) } + \text{ H}_2\text{CO}_3 \text{ (aq)} \\
\text{SO}_3 \text{ (g) } & + \text{ H}_2\text{O} \text{ (R) } \text{----> } \text{H}_2\text{SO}_4 \text{ (aq)} \\
2 \text{ KHCO}_3 \text{ (s) } & \text{----> } \text{K}_2\text{CO}_3 \text{ (s) } + \text{ CO}_2 \text{ (g) } + \text{ H}_2\text{O} \text{ (g)}
\end{align*}
\]
6. a. How many moles of BaBr$_2$ are contained in 150.0-g of BaBr$_2$?

\[ \text{molBaBr}_2 = \frac{150 \text{gBaBr}}{1} \times \frac{1 \text{molBaBr}_2}{217 \text{gBaBr}_2} = 0.691 \text{molBaBr}_2 \]

b. How many atoms of Cl are in 42.6-g of CHCl$_3$?

\[ \text{atomsCl} = \frac{42.6 \text{gCHCl}_3}{1} \times \frac{3 \times 6.022 \times 10^{23} \text{atomsCl}}{119.5 \text{gCHCl}_3} = 6.44 \times 10^{23} \text{atomsCl} \]

c. How many grams of Sr are contained in sufficient SrCl$_2$ to obtain 7.2 x 10$^{22}$ atoms of Cl?

\[ \text{gSr} = \frac{7.2 \times 10^{22} \text{atomsCl}}{1} \times \frac{87.6 \text{gSr}}{2 \times 6.022 \times 10^{23} \text{atomsCl}} = 5.2 \text{gSr} \]

d. How many grams of N are contained in 350 molecules of N$_2$H$_4$?

\[ \text{gN} = \frac{350 \text{moleculesN}_2\text{H}_4}{1} \times \frac{28 \text{gN}}{6.022 \times 10^{23} \text{moleculesN}_2\text{H}_4} = 1.6 \times 10^{-20} \text{gN} \]

e. How many molecules of H$_2$SO$_4$ are contained in 75.0-g of H$_2$SO$_4$?

\[ \text{moleculesH}_2\text{SO}_4 = \frac{75.0 \text{gH}_2\text{SO}_4}{1} \times \frac{6.022 \times 10^{23} \text{moleculesH}_2\text{SO}_4}{98 \text{gH}_2\text{SO}_4} = 4.61 \times 10^{23} \text{moleculesH}_2\text{SO}_4 \]
7.  a. How many grams of N are there in 15.0-g of NH₄NO₃?

\[
g_N = \frac{15.0 \text{ g} \text{NH}_4\text{NO}_3}{1} \times \frac{28 \text{ g} \text{N}}{80 \text{ g} \text{NH}_4\text{NO}_3} = 5.25 \text{ gN}
\]

b. How many molecules of water are in one gallon - (or 3.784 L). Assume the density of water is 1 g/mL.

\[
m_{\text{moleculesH}_2\text{O}} = \frac{3.784 \text{ L}}{1} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ g}}{1 \text{ mL}} \times \frac{6.022 \times 10^{23} \text{ moleculesH}_2\text{O}}{18 \text{ gH}_2\text{O}} = 1.266 \times 10^{26} \text{ moleculesH}_2\text{O}
\]

c. How many moles of C₂H₆ are in 115.0-g of C₂H₆?

\[
m_{\text{molC}_2\text{H}_6} = \frac{115.0 \text{ gC}_2\text{H}_6}{1} \times \frac{1 \text{ molC}_2\text{H}_6}{30 \text{ gC}_2\text{H}_6} = 3.833 \text{ molC}_2\text{H}_6
\]

d. How many atoms of oxygen are contained in 245.0-g of KMnO₄?

\[
m_{\text{atomsO}} = \frac{245.0 \text{ gKMnO}_4}{1} \times \frac{4 \times 6.022 \times 10^{23} \text{ atomsO}}{158 \text{ gKMnO}_4} = 3.736 \times 10^{24} \text{ atomsO}
\]

e. How many grams of BaCl₂ are required to obtain 450 atoms of Cl?

\[
m_{\text{gramsBaCl}_2} = \frac{450 \text{ atomsCl}}{1} \times \frac{208 \text{ gramsBaCl}_2}{2 \times 6.022 \times 10^{23} \text{ atomsCl}} = 7.8 \times 10^{-20} \text{ gramsBaCl}_2
\]
8. Hydrazine, N₂H₄, is used extensively as a rocket fuel. It can be prepared from the reaction:

\[ \text{Given:} \quad 15.0\text{-g NH}_3 \quad 17.5\text{-g OCl}^- \]

\[ 2 \text{mol NH}_3 + 1 \text{mol OCl}^- \rightarrow \text{N}_2\text{H}_4 + \text{Cl}^- + \text{H}_2\text{O} \]

**Stoichiometry:**

<table>
<thead>
<tr>
<th>2 mol NH₃</th>
<th>1 mol OCl⁻</th>
<th>1 mol N₂H₄</th>
<th>1 mol Cl⁻</th>
<th>1 mol H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 x 17 g NH₃</td>
<td>1 x 51.5 g OCl⁻</td>
<td>1 x 32 g N₂H₄</td>
<td>1 x 35.5 g Cl⁻</td>
<td>1 x 18 g H₂O</td>
</tr>
</tbody>
</table>

Suppose the reaction is carried out with 15.0-g of NH₃ and 17.5-g of OCl⁻.

**a.** Which is the limiting reactant?

\[
\text{Compare:} \quad \frac{15.0\text{gNH}_3}{2\times17\text{gNH}_3} = 0.441 \quad \frac{17.5\text{gOCl}^-}{1\times51.5\text{gOCl}^-} = 0.340
\]

*Since the stoichiometric ratio for OCl⁻ is less than that for NH₃, OCl⁻ is the limiting reactant.*

**b.** How many grams hydrazine could be formed in the reaction above?

**Based on limiting reactant:**

\[
?\text{gN}_2\text{H}_4 = \frac{17.5\text{gOCl}^-}{1} \times \frac{1\times32\text{gN}_2\text{H}_4}{1\times51.5\text{gOCl}^-} = 10.9\text{gN}_2\text{H}_4
\]

**c.** In the above reaction, how many grams of the excess reactant would be expected to remain after the reaction is complete.

**First, find out how many grams of the excess reactant, NH₃, reacted.**

\[
?\text{gNH}_3\text{reacted} = \frac{17.5\text{gOCl}^-}{1} \times \frac{2\times17\text{gNH}_3}{1\times51.5\text{gOCl}^-} = 11.6\text{gNH}_3
\]

*The remaining amount is the starting amount minus the reacted amount:*

\[ 15.0\text{-g} - 11.6\text{-g} = 3.4\text{g NH}_3 \text{ remain} \]
9. The reason hydrofluoric acid, HF(aq), cannot be stored in glass bottles and can be used in etching glass is its reaction with silica, SiO₂.

\[
\begin{align*}
15.0\text{-g SiO}_2 & \quad 25.0\text{-g HF} \\
\text{SiO}_2 (s) & + \quad 6 \text{ HF (aq)} \quad \text{W} \quad \text{H}_2\text{SiF}_6 (aq) & + \quad 2 \text{ H}_2\text{O (l)} \\
1 \text{ mol SiO}_2 & \quad 6 \text{ mol HF} & \quad 1 \text{ mol H}_2\text{SiF} & \quad 2 \text{ mol H}_2\text{O} \\
1 \times 60 \text{ g SiO}_2 & \quad 6 \times 20 \text{ g HF} & \quad 1 \times 144 \text{ g H}_2\text{SiF} & \quad 2 \times 18 \text{ g H}_2\text{O}
\end{align*}
\]

Suppose 15.0-g of SiO₂ are reacted with 25.0-g of HF.

a. Which is the limiting reactant?

**Using conversion factor approach:**

\[
\text{? g HF required to use up SiO}_2 = \frac{15.0\text{gSiO}_2}{1} \times \frac{6\times20\text{gHF}}{1\times60\text{gSiO}_2} = 30.0\text{g HF required}
\]

Since we only have 25.0-g of HF, we will run out and HF will be the limiting reactant.

Alternatively, look at the two ratios formed from the equation above:

\[
\frac{15.0\text{gSiO}_2}{1\times60\text{gSiO}_2} = 0.25 \quad \frac{25.0\text{gHF}}{6\times20\text{gHF}} = 0.208
\]

Since the HF ratio is smaller, it is limiting using this approach, too, which is fortunate.

b. How many grams of H₂SiF₆ could be produced in the reaction?

\[
\text{? g H}_2\text{SiF}_6 \text{ formed} = \frac{25.0\text{gHF}}{1} \times \frac{1\times144\text{g H}_2\text{SiF}_6}{6\times20\text{gHF}} = 30.0\text{g H}_2\text{SiF}_6 \text{ formed}
\]

Notice this calculation is based on the limiting reactant, HF. It is also an absolute coincidence that the grams of HF required and the grams of H₂SiF₆ are the same.

c. If the experiment actually produces 22.5-g of H₂SiF₆ what is the percent yield?

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{22.5\text{g}}{30.0\text{g}} \times 100\% = 75.0\% \text{ yield}
\]

d. How many grams of the excess reactant remain at the end of the reaction?

\[
\text{? g SiO}_2\text{used} = \frac{25.0\text{gHF}}{1} \times \frac{1\times60\text{gSiO}_2}{6\times20\text{gHF}} = 12.5\text{g SiO}_2 \text{ used}
\]

\#g SiO₂ remaining at end = 15.0 g started with – 12.5 g used = 2.5 g SiO₂
10. The following reaction occurs during the production of ammonia.

\[ 55.0 \text{ g } N_2 \quad 20.0 \text{ g } H_2 \]

\[ N_2 (g) + 3 \text{ H}_2 (g) \rightarrow 2 \text{ NH}_3 (g) \]

\[
\begin{align*}
1 \text{ mol } N_2 & \quad 3 \text{ mol } H_2 & \quad 2 \text{ mol } NH_3 \\
1 \times 28 \text{ g } N_2 & \quad 3 \times 2 \text{ g } H_2 & \quad 2 \times 17 \text{ g } NH_3
\end{align*}
\]

a. Balance the equation.

b. 55.0-g of N₂ are mixed with 20.0-g of H₂. Which reactant is the limiting reactant?

First, find out how much H₂ is required to react with the N₂, or vice versa if you wish.

\[ ? \text{ g } H_2 \text{ required to react with } 55.0 \text{ g } N_2 = \frac{55.0 \text{ g } N_2}{1} \times \frac{3 \times 2 \text{ g } H_2}{1 \times 28 \text{ g } N_2} = 11.8 \text{ g } H_2 \text{ required} \]

Since we have 20.0 g of H₂ and only need 11.8 g, we have excess hydrogen and the nitrogen limits the reaction.

c. How many grams of ammonia could be produced from the reaction in part b?

Using the limiting reactant, N₂:

\[ ? \text{ g } NH_3 \text{ produced } = \frac{55.0 \text{ g } N_2}{1} \times \frac{2 \times 17 \text{ g } NH_3}{1 \times 28 \text{ g } N_2} = 66.8 \text{ g } NH_3 \text{ produced} \]

d. How many grams of the excess reactant would remain at the end of the reaction?

From part b, we know that 11.8 g of H₂ were used in the reaction. To find the excess, simply subtract 11.8 g from the starting amount.

\[ \# \text{ g } H_2 \text{ left over } = 20.0 \text{ g } - 11.8 \text{ g } = 12.2 \text{ g } H_2 \]
11. Balance each of the following equations by placing coefficients in the appropriate places.

\[ 2 \text{C}_6\text{H}_6 (\text{R}) + 15 \text{O}_2 \rightarrow 12 \text{CO}_2 (\text{g}) + 6 \text{H}_2\text{O} (\text{R}) \]

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

\[ \text{Au}_2\text{S}_3(\text{s}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{Au}(\text{s}) + 3 \text{H}_2\text{S} (\text{g}) \]

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

\[ 2 \text{Eu} (\text{s}) + 6 \text{HF} (\text{g}) \rightarrow 2 \text{EuF}_3 (\text{s}) + 3 \text{H}_2 (\text{g}) \]

12. Find the number of moles and molecules in 17.50-g of CO$_2$. Be sure to show your work.

\[ \text{? moles} = \frac{17.50 \text{g CO}_2}{1} \times \frac{1\text{mol CO}_2}{44 \text{g CO}_2} = 0.3977 \text{mol CO}_2 \]

\[ \text{? molecules} = \frac{17.50 \text{g CO}_2}{1} \times \frac{6.022 \times 10^{23} \text{molecules CO}_2}{44 \text{g CO}_2} = 2.395 \times 10^{23} \text{molecules CO}_2 \]

13. Find the number of grams and moles in 4.5 x $10^{22}$ formula units of KBr.

14. The reusable booster rockets of the space shuttle program use a mixture of aluminum and ammonium perchlorate for fuel. A possible equation for this reaction is:

\[ 3 \text{Al} (\text{s}) + 3 \text{NH}_4\text{ClO}_4 (\text{s}) \rightarrow \text{Al}_2\text{O}_3 (\text{s}) + \text{AlCl}_3 (\text{s}) + 3 \text{NO} (\text{g}) + 6 \text{H}_2\text{O} (\text{g}) \]

<table>
<thead>
<tr>
<th></th>
<th>1000 g NH$_4$ClO$_4$</th>
</tr>
</thead>
<tbody>
<tr>
<td>3Al (s)</td>
<td>3 mol Al</td>
</tr>
<tr>
<td>3NH$_4$ClO$_4$ (s)</td>
<td>3 mol NH$_4$ClO$_4$</td>
</tr>
<tr>
<td>Al$_2$O$_3$ (s)</td>
<td>1 mol Al$_2$O$_3$</td>
</tr>
<tr>
<td>AlCl$_3$ (s)</td>
<td>1 mol AlCl$_3$</td>
</tr>
<tr>
<td>3 NO (g)</td>
<td>3 mol NO</td>
</tr>
<tr>
<td>6 H$_2$O (g)</td>
<td>6 mol H$_2$O</td>
</tr>
</tbody>
</table>

3 x 27 g Al 3 x 117.5 g NH$_4$ClO$_4$ 1 x 102 g Al$_2$O$_3$ 1 x 133.5 g AlCl$_3$ 3 x 30 g NO 6 x 18 g H$_2$O

a. What mass of aluminum is required to react with every 1000-g of ammonium perchlorate burned?

\[ \text{? g Al} = \frac{1000 \text{g NH}_4\text{ClO}_4}{1} \times \frac{3 \times 27 \text{g Al}}{3 \times 117.5 \text{g NH}_4\text{ClO}_4} = 229.8 \text{g Al} \]

b. What mass of NO is produced when the 1000-g of ammonium perchlorate is burned?

\[ \text{? g NO} = \frac{1000 \text{g NH}_4\text{ClO}_4}{1} \times \frac{3 \times 30 \text{g NO}}{3 \times 117.5 \text{g NH}_4\text{ClO}_4} = 255.3 \text{g NO produced} \]
15. a. How many N atoms are contained in 75.0-g of N₂H₄ known as the rocket fuel hydrazine?

\[ \text{Natoms} = \frac{75.0 \text{gN}_2\text{H}_4}{1} \times \frac{2 \times 6.022 \times 10^{23} \text{Natoms}}{32 \text{gN}_2\text{H}_4} = 2.82 \times 10^{24} \text{Natoms} \]

b. How many hydrogen atoms are contained in 100 molecules of H₂O₂?

hydrogen atoms in each molecule, 100 molecules = 2 x 100 hydrogen atoms = 200 hydrogen atoms

c. How many moles of CHF₂Cl are contained in 150.0-g of CHF₂Cl?

\[ \text{molCHF}_2\text{Cl} = \frac{150.0 \text{gCHF}_2\text{Cl}}{1} \times \frac{1 \text{molCHF}_2\text{Cl}}{86.5 \text{gCHF}_2\text{Cl}} = 1.734 \text{molCHF}_2\text{Cl} \]
16. Ammonia can be converted to nitric oxide, NO, by the reaction below:

\[
\begin{align*}
15.0 \text{ g NH}_3 & \quad 25.0 \text{ g O}_2 \\
4 \text{ mol NH}_3 & \quad 5 \text{ mol O}_2 & \quad 4 \text{ mol NO} & \quad 6 \text{ mol H}_2\text{O}
\end{align*}
\]

\[
\begin{align*}
4 \times 17 \text{ g NH}_3 & \quad 5 \times 32 \text{ g O}_2 & \quad 4 \times 30 \text{ g NO} & \quad 6 \times 18 \text{ g H}_2\text{O}
\end{align*}
\]

a. 15.0-g of NH\textsubscript{3} are reacted with 25.0-g of O\textsubscript{2}. Which is the limiting reactant?

**Figure out how many grams of O\textsubscript{2} are required to use up all of the NH\textsubscript{3}**.

\[
\begin{align*}
? \text{ g O}_2 &= \frac{15.0 \text{ g NH}_3}{4 \times 17 \text{ g NH}_3} \times \frac{5 \times 32 \text{ g O}_2}{4 \times 32 \text{ g O}_2} = 35.3 \text{ g O}_2 \text{ required}
\end{align*}
\]

Since we need 35.3 g of O\textsubscript{2} to react with all of the NH\textsubscript{3} and only have 25.0 g of O\textsubscript{2}, we will run out of O\textsubscript{2}. It is limiting.

b. How many grams of NO could be formed in the reaction of part a?

\[
\begin{align*}
? \text{ g NO} &= \frac{25.0 \text{ g O}_2}{1} \times \frac{4 \times 30 \text{ g NO}}{5 \times 32 \text{ g O}_2} = 18.8 \text{ g NO}
\end{align*}
\]

c. How many grams of the excess reactant would remain after the reaction of part a?

**First, find out how many grams of the excess reactant, NH\textsubscript{3}, are used**.

\[
\begin{align*}
? \text{ g NH}_3 \text{ used} &= \frac{25.0 \text{ g O}_2}{1} \times \frac{4 \times 17 \text{ g NH}_3}{5 \times 32 \text{ g O}_2} = 10.6 \text{ g NH}_3 \text{ used}
\end{align*}
\]

Amount of NH\textsubscript{3} remaining = start amount of NH\textsubscript{3} – amount of NH\textsubscript{3} used

\[
\begin{align*}
&= 15.0 \text{ g} – 10.6 \text{ g} = 4.4 \text{ g NH}_3
\end{align*}
\]

d. If 15.0-g of NO is formed in the reaction of part a, what is the percent yield?

\[
\begin{align*}
\% \text{ yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \\
&= \frac{15.0 \text{ g NO}}{18.8 \text{ g NO}} \times 100\% = 79.8\%
\end{align*}
\]
17. Balance each of the following equations and classify it as combustion, decomposition, combination, methathesis, and/or oxidation-reduction. Note that some equations may fall into more than one category. State all categories that apply in each case.

<table>
<thead>
<tr>
<th>Equation</th>
<th>Classification</th>
</tr>
</thead>
<tbody>
<tr>
<td>( 2 \text{C(gr)} + \text{___O}_2 (g) \rightarrow \text{2CO (g)} )</td>
<td>combination, oxidation-reduction</td>
</tr>
<tr>
<td>( \text{NH}_3 (g) + \text{___HCl (g)} \rightarrow \text{___NH}_4\text{Cl (s)} )</td>
<td>combination</td>
</tr>
<tr>
<td>( \text{Cr(s)} + \text{NiSO}_4 \text{ (aq)} \rightarrow \text{Ni (s)} + \text{CrSO}_4 \text{ (aq)} )</td>
<td>oxidation-reduction</td>
</tr>
<tr>
<td>( \text{HC}_2\text{H}_3\text{O}_2 \text{ (aq)} + \text{NaOH (aq)} \rightarrow \text{HOH (aq)} + \text{NaC}_2\text{H}_3\text{O}_2 \text{ (aq)} )</td>
<td>metathesis</td>
</tr>
<tr>
<td>( \text{2H}_2\text{O (aq)} \rightarrow \text{2H}_2 (g) + \text{___O}_2 (g) )</td>
<td>decomposition, oxidation-reduction</td>
</tr>
</tbody>
</table>

18. Answer each of the following questions. Show your work to receive full credit.

a. How many moles are contained in 14.5-g of \( \text{CH}_3\text{Br} \)?

\[
? \text{mol CH}_3\text{Br} = \frac{14.5 \text{ g CH}_3\text{Br}}{95 \text{ g CH}_3\text{Br}} \times 1 \text{ mol CH}_3\text{Br} = 0.153 \text{ mol CH}_3\text{Br}
\]

b. How many oxygen atoms are contained in 5 molecules of \( \text{C}_2\text{H}_8\text{O}_2 \)?

\[
? \text{mol O atoms} = \frac{5 \text{ molecules C}_2\text{H}_8\text{O}_2}{1} \times \frac{2 \text{ O atoms}}{1 \text{ molecule C}_2\text{H}_8\text{O}_2} = 10 \text{ O atoms}
\]

or just think about it – 2 O atoms per molecule and 5 molecules = 10 O atoms

c. How many grams are contained in 0.125-mol of \( \text{Ba(NO}_3)_2 \)?

\[
? \text{g Ba(NO}_3)_2 = \frac{0.125 \text{ moles Ba(NO}_3)_2}{1} \times \frac{261 \text{ g Ba(NO}_3)_2}{1 \text{ mole Ba(NO}_3)_2} = 32.6 \text{ g Ba(NO}_3)_2
\]

d. What is the molar mass of \( \text{MgSO}_4\text{C}_2\text{H}_2\text{O} \)?

\[
\text{molar mass} = 1 \times 24.3 + 1 \times 32 + 4 \times 16 + 7 \times 18 = 246.3 \text{ g/mol}
\]

e. How many grams of carbon are contained in 25.0-g of \( \text{CH}_2\text{Cl}_2 \)?

\[
? \text{g C} = \frac{25.0 \text{ g CH}_2\text{Cl}_2}{1} \times \frac{1 \times 12 \text{ g C}}{85 \text{ g CH}_2\text{Cl}_2} = 3.52 \text{ g C}
\]
19. Balance each of the following chemical equations by placing the appropriate coefficients on the underlines.

\[ _{\text{TiCl}_4} + 6 \text{H}_2\text{O} \rightarrow 6 \text{TiO}_2 + 6 \text{HCl} \]

\[ _{\text{P}_4} + 10 \text{Cl}_2 \rightarrow 4 \text{PCl}_5 \]

\[ _{\text{Al}_2\text{O}_3} + 3 \text{H}_2\text{SeO}_4 \rightarrow 6 \text{Al}_2(\text{SeO}_4)_3 + 3 \text{H}_2\text{O} \]

\[ 2 \text{C}_2\text{H}_6 + 7 \text{O}_2 \rightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O} \]

\[ _{\text{Ca}_3(\text{PO}_4)_2} + 2 \text{H}_2\text{SO}_4 \rightarrow 6 \text{Ca(H}_2\text{PO}_4)_2 + 2 \text{CaSO}_4 \]

20. Nitromethane, CH\textsubscript{3}NO\textsubscript{2}, reacts with oxygen to form carbon dioxide, water, and nitrogen dioxide.

a. Write the balanced equation for this reaction.

\[ 4 \text{CH}_3\text{NO}_2 + 7 \text{O}_2 \rightarrow 4 \text{CO}_2 + 6 \text{H}_2\text{O} + 4 \text{NO}_2 \]

b. How many moles of oxygen would be required to react with 8-mol of nitromethane?

*It takes 7 moles of O\textsubscript{2} for every 4 mol of nitromethane, so 8 mol of nitromethane would require 8 mol \times \frac{7}{4} = 14 mol O\textsubscript{2}.*

21. The $\beta$-blocker drug, timolol, is expected to reduce the need for heart bypass surgery. Its composition by mass is 47.2 %C, 6.55 %H, 13.0 %N, 25.9 %O, and 7.43 % S. Its molar mass is 432-g/mol.

a. What is the empirical formula for timolol?

Divide by smallest:

\[
\begin{align*}
\text{C}: & \quad 47.2/12 = 3.93 \quad 3.93/0.232 = 16.9 \quad \text{C}_{17}\text{H}_{28}\text{N}_4\text{O}_7\text{S} \\
\text{H}: & \quad 6.55/1 = 6.55 \quad 6.55/0.232 = 28.2 \\
\text{N}: & \quad 13.0/14 = 0.929 \quad 0.929/0.232 = 4.00 \quad \text{Mass of this formula} = 432 \\
\text{O}: & \quad 25.9/16 = 1.619 \quad 1.619/0.232 = 6.98 \\
\text{S}: & \quad 7.43/32 = 0.232 \quad 0.232/0.232 = 1
\end{align*}
\]

b. What is the molecular formula for timolol? Since the molar mass = mass of the empirical formula, the empirical formula is the molecular formula:

\[ \text{C}_{17}\text{H}_{28}\text{N}_4\text{O}_7\text{S} \]
22. Find the percent mass of each element in C$_9$H$_9$N.

\[
\begin{align*}
\%C &= \frac{9 \times 12}{9 \times 12 + 9 \times 1 + 1 \times 14} \times 100\% = 82.4\%C \\
\%H &= \frac{9 \times 1}{9 \times 12 + 9 \times 1 + 1 \times 14} \times 100\% = 6.87\%H \\
\%N &= \frac{1 \times 14}{9 \times 12 + 9 \times 1 + 1 \times 14} \times 100\% = 10.7\%N
\end{align*}
\]

23. 

a. How many grams of NH$_3$ are contained in 2.45 x $10^{-3}$ mol of NH$_3$?

\[
\text{? g } NH_3 = \frac{2.45 \times 10^{-3} \text{ mol } NH_3}{1} \times \frac{17 \text{ g } NH_3}{1 \text{ mol } NH_3} = 0.0417 \text{ g } NH_3
\]

b. How many atoms of C are contained in 0.193-g of C$_{12}$H$_{22}$O$_{11}$?

\[
\text{? atoms C} = \frac{0.193 \text{ g } C_{12}H_{22}O_{11}}{1} \times \frac{12 \times 6.022 \times 10^{23} \text{ atoms C}}{342 \text{ g } C_{12}H_{22}O_{11}} = 4.08 \times 10^{21} \text{ atoms C}
\]

c. How many moles of PCl$_3$ are contained in 75.0-g of PCl$_3$?

\[
\text{? mol PCl}_3 = \frac{75.0 \text{ g } PCl_3}{1} \times \frac{1 \text{ mol } PCl_3}{137.5 \text{ g } PCl_3} = 0.545 \text{ mol } PCl_3
\]

d. How many atoms of oxygen are contained in 35 molecules of CO$_2$?

\[
\text{? atoms O} = \frac{35 \text{ molecules } CO_2}{1} \times \frac{2 \text{ atoms O}}{1 \text{ molecule } CO_2} = 70 \text{ atoms O}
\]

e. How many molecules of H$_2$O are contained in 35.0-g of water?

\[
\text{? molecules } H_2O = \frac{35.0 \text{ g } H_2O}{1} \times \frac{6.022 \times 10^{23} \text{ molecules } H_2O}{18 \text{ g } H_2O} = 1.17 \times 10^{24} \text{ molecules } H_2O
\]

24. 

a. How many grams of BaI$_2$ are contained in 1.25-mol of BaI$_2$?

\[
\text{? g } BaI_2 = \frac{1.25 \text{ mol } BaI_2}{1} \times \frac{391 \text{ g } BaI_2}{1 \text{ mol } BaI_2} = 489 \text{ g } BaI_2
\]

b. How many oxygen atoms are contained in 10.0-g of water?

\[
\text{? O atoms} = \frac{10.0 \text{ g } H_2O}{1} \times \frac{1 \times 6.022 \times 10^{23} \text{ O atoms}}{18 \text{ g } H_2O} = 3.35 \times 10^{23} \text{ O atoms}
\]

c. How many moles of NH$_4$NO$_3$ are contained in 25.0-g of NH$_4$NO$_3$?

\[
\text{? mol } NH_4NO_3 = \frac{25.0 \text{ g } NH_4NO_3}{1} \times \frac{1 \text{ mol } NH_4NO_3}{80 \text{ g } NH_4NO_3} = 0.312 \text{ mol } NH_4NO_3
\]