Stoichiometry Worksheet 2:

Percent Yield
For each of the problems:

a. Write the balanced chemical equation
b. Identify the given (with units) and what you want to find (with units)
c. Show set up with units. Check sig figs, give final answer with units and label.
1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?

Equation: \[ 2 \text{H}_2\text{O} \rightarrow \text{O}_2 + 2\text{H}_2 \]

Before:

Change:

After:
Change grams to moles!

Only moles go in the BCA table!

- 36 g x $\frac{1 \text{ mole } \text{H}_2\text{O}}{18.02 \text{ g}} = 2.0 \text{ mole }$
1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?

Equation: \(2 \text{H}_2\text{O} \rightarrow \text{O}_2 + 2\text{H}_2\)

Before:

\begin{align*}
2.0 & \quad 0 \\
0 & \quad 0
\end{align*}

Change:

After:
Calculate the change!

\[
\begin{align*}
2.0 \text{ mol } H_2O & \times \frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O} = 2 \text{ mol } H_2 \\
2.0 \text{ mol } H_2O & \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} = 1 \text{ mol } O_2
\end{align*}
\]

Equation: \( 2 \text{ H}_2\text{O} \rightarrow O_2 + 2\text{H}_2 \)

Before: 
\[
\begin{array}{c}
2.0 \\
0 \\
0
\end{array}
\]

Change: 

After: 

1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?

Equation: \[ 2 \text{H}_2\text{O} \rightarrow \text{O}_2 + 2\text{H}_2 \]

<table>
<thead>
<tr>
<th></th>
<th>(2 \text{H}_2\text{O})</th>
<th>(\rightarrow)</th>
<th>(\text{O}_2)</th>
<th>+</th>
<th>(2\text{H}_2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Before:</td>
<td>2.0</td>
<td>0</td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change:</td>
<td>-2.0</td>
<td>1.0</td>
<td>2.0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>After:</td>
<td>0</td>
<td>1.0</td>
<td>2.0</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Change moles to grams!

- $2.0 \text{ mole} \times \frac{2.02 \text{ g}}{1 \text{ mole } \text{H}_2} = 4.0 \text{ g } \text{H}_2$
2. Recall that liquid sodium reacts with chlorine gas to produce sodium chloride. You want to produce 581 g of sodium chloride. How many grams of sodium are needed?

Equation: 2 Na + Cl₂ → 2 NaCl

Before:
Change:
After:
Change grams to moles!

- $581 \text{ g of NaCl} \times \frac{1 \text{ mole H}_2\text{O}}{58.44 \text{ g NaCl}} = 9.94 \text{ mole H}_2\text{O}$

$58.44 \text{ g}$

$\text{Na: } 22.99$

$\text{Cl: } +35.45$

$\text{NaCl: } 58.44\text{g}$
2. Recall that liquid sodium reacts with chlorine gas to produce sodium chloride. You want to produce 581 g of sodium chloride. How many grams of sodium are needed?

<table>
<thead>
<tr>
<th>Equation:</th>
<th>2 Na + Cl₂ → 2 NaCl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Before:</td>
<td>?</td>
</tr>
<tr>
<td>Change:</td>
<td>9.94</td>
</tr>
<tr>
<td>After:</td>
<td>9.94</td>
</tr>
</tbody>
</table>
\[ 9.94 \text{ mol NaCl} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol NaCl}} = 4.97 \text{ mol Cl}_2 \]

\[ 9.94 \text{ mol NaCl} \times \frac{2 \text{ mol Na}}{2 \text{ mol NaCl}} = 9.94 \text{ mol Na} \]

Equation: \[ 2 \text{ Na} + \text{ Cl}_2 \rightarrow 2 \text{ NaCl} \]

Before: 
- NaCl: 9.94 mol
- Cl: 0 mol

Change: 
- NaCl: -9.94 mol
- Cl: -4.97 mol

After: 
- NaCl: 0 mol
- Cl: 9.94 mol
Change moles to grams!

- $9.94 \text{ mole Na} \times \frac{22.99 \text{ g}}{1 \text{ mole Na}} = 228 \text{ g Na}$
3. You eat 180.0 g of glucose (90 M&Ms). If glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, reacts with oxygen gas to produce carbon dioxide and water, how many grams of oxygen will you have to breathe in to burn the glucose?

Equation: $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$

Before:

Change:

After:
Change grams to moles!

- 180.0 g of $\text{C}_6\text{H}_{12}\text{O}_6$ x $\frac{1\text{ mole H}_2\text{O}}{180.18\text{ g}}$ = 0.9990 mole

6C : 6(12.01)
12H : 12(1.01)
6O : +6(16.00)
180.18 g
3. You eat 180.0 g of glucose (90 M&Ms). If glucose, C₆H₁₂O₆, reacts with oxygen gas to produce carbon dioxide and water, how many grams of oxygen will you have to breathe in to burn the glucose?

Equation: \[ \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} \]

Before: 0.9990 XS 0 0

Change: 

After:
\[
\begin{align*}
0.9990 \text{ mol} \times \frac{6 \text{ mol} O_2}{1 \text{ mol}} &= 5.994 \text{ mol} O_2 \\
0.9990 \text{ mol} \times \frac{6 \text{ mol} CO_2}{1 \text{ mol}} &= 5.994 \text{ mol} CO_2 \\
0.9990 \text{ mol} \times \frac{6 \text{ mol} H_2O}{1 \text{ mol}} &= 5.994 \text{ mol} H_2O
\end{align*}
\]

<table>
<thead>
<tr>
<th>Equation:</th>
<th>$C_6H_{12}O_6$ + 6$O_2$ $\rightarrow$ 6$CO_2$ + 6$H_2O$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Before:</td>
<td>0.9990 XS 0 0</td>
</tr>
<tr>
<td>Change:</td>
<td>-0.9990 -5.994 5.994 5.994</td>
</tr>
<tr>
<td>After:</td>
<td>0 XS 5.994 5.994</td>
</tr>
</tbody>
</table>
Change moles to grams!

- $5.994 \text{ mole } O_2 \times \frac{32.00 \text{ g}}{1 \text{ mole } O_2} = 191.8 \text{ g } O_2$
4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?

Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?

Equation: \( \text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \)

Before:

Change:

After:
Change grams to moles!

- $4.61\text{ g of Zn} \times \frac{1\text{ mole Zn}}{65.38\text{ g}} = 0.0705\text{ mole}$
4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?

Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?

Equation: \[ \text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]

<table>
<thead>
<tr>
<th>Before:</th>
<th>0.0705</th>
<th>XS</th>
<th>0</th>
<th>0</th>
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</thead>
<tbody>
<tr>
<td>Change:</td>
<td>-0.0705</td>
<td>-0.141</td>
<td>0.0705</td>
<td>0.0705</td>
</tr>
<tr>
<td>After:</td>
<td>0</td>
<td>XS</td>
<td>0.0705</td>
<td>0.0705</td>
</tr>
</tbody>
</table>
Change moles to grams!

- $0.0705 \text{ mole } \text{ZnCl}_2 \times \frac{136.28 \text{ g}}{1 \text{ mole } \text{ZnCl}_2} = 9.61 \text{ g } \text{ZnCl}_2$
4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?

Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?

Equation: Zn + 2 HCl → ZnCl₂ + H₂

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<tbody>
<tr>
<td>Before:</td>
<td>0.0705</td>
<td>XS</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change:</td>
<td>-0.0705</td>
<td>-0.141</td>
<td>0.0705</td>
<td>0.0705</td>
</tr>
<tr>
<td>After:</td>
<td>0</td>
<td>XS</td>
<td>0.0705</td>
<td>0.0705</td>
</tr>
</tbody>
</table>
Find %Yield!

- $0.0705 \text{ mole ZnCl}_2 \times \frac{136.28 \text{ g}}{1 \text{ mole ZnCl}_2} = 9.61 \text{ g ZnCl}_2$

%Yield = \frac{\text{ACTUAL}}{\text{THEORETICAL}} \times 100

= \frac{8.56 \text{ g}}{9.61 \text{ g}} \times 100 = 89.1\% \text{ yield}$
5. Determine the mass of carbon dioxide that should be produced in the reaction between 3.74 g of carbon and excess O$_2$. What is the % yield if 11.34 g of CO$_2$ is recovered?

<table>
<thead>
<tr>
<th>Equation:</th>
<th>C + O$_2$ → CO$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Before:</td>
<td>0.311 XS 0</td>
</tr>
<tr>
<td>Change:</td>
<td>- 0.311 - 0.311 0.311</td>
</tr>
<tr>
<td>After:</td>
<td>0 XS 0.311</td>
</tr>
</tbody>
</table>

After: 0 XS 0.311
Calculations

\[
3.74 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g}} = 0.311 \text{ mol C}
\]

\[
0.311 \text{ mol C} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} = 0.311 \text{ mol CO}_2
\]

\[
0.311 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 13.7 \text{ g CO}_2
\]

\[
\frac{11.34 \text{ g CO}_2}{13.7 \text{ g CO}_2} \times 100\% = 82.8\% \text{ yield}
\]
6. In the reaction between excess K(s) and 4.28 g of O₂(g), potassium oxide is formed. What mass would you expect to form (theoretical yield)? If 17.36 g of K₂O is actually produced, what is the percent yield?

Equation: 4K + O₂ → 2K₂O

Before:

Change:

After:
6. In the reaction between excess K(s) and 4.28 g of O₂(g), potassium oxide is formed. What mass would you expect to form (theoretical yield)? If 17.36 g of K₂O is actually produced, what is the percent yield?

Equation: 4K + O₂ → 2K₂O

| Before: | XS | 0.134 | 0 |
| Change: | -0.536 | -0.134 | 0.266 |
| After:  | XS | 0 | 0.266 |
Calculations

\[
\begin{align*}
4.28 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} &= 0.134 \text{ mol } O_2 \\
0.134 \text{ mol } O_2 \times \frac{2 \text{ mol } K_2O}{1 \text{ mol } O_2} &= 0.268 \text{ mol } K_2O \\
0.268 \text{ mol } K_2O \times \frac{94.20 \text{ g } K_2O}{1 \text{ mol } K_2O} &= 25.2 \text{ g } K_2O \\
\frac{17.36 \text{ g } K_2O}{25.2 \text{ g } K_2O} \times 100\% &= 68.9 \% \text{ yield}
\end{align*}
\]
7. **Determine the mass of carbon dioxide** one could expect to form (and the percent yield) for the reaction between excess CH$_4$ and 11.6 g of O$_2$ if 5.38 g of carbon dioxide gas is produced along with some water vapor.

**Equation:**

$$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$$

**Before:**

<table>
<thead>
<tr>
<th></th>
<th>CH$_4$</th>
<th>2O$_2$</th>
<th>→</th>
<th>CO$_2$</th>
<th>+</th>
<th>2H$_2$O</th>
</tr>
</thead>
<tbody>
<tr>
<td>XS</td>
<td>0.363</td>
<td>0</td>
<td></td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
</tbody>
</table>

**Change:**

**After:**

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>0.363</td>
</tr>
</tbody>
</table>
7. Determine the mass of carbon dioxide one could expect to form (and the percent yield) for the reaction between excess CH\(_4\) and 11.6 g of O\(_2\) if 5.38 g of carbon dioxide gas is produced along with some water vapor.

Equation: \[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

<table>
<thead>
<tr>
<th></th>
<th>Before:</th>
<th>Change:</th>
<th>After:</th>
</tr>
</thead>
<tbody>
<tr>
<td>XS</td>
<td>0.363</td>
<td>-0.363</td>
<td>0.363</td>
</tr>
<tr>
<td>0</td>
<td>0</td>
<td>0.181</td>
<td>0.181</td>
</tr>
</tbody>
</table>

After: XS 0 0.181 0.363
Calculations

\[ 11.6gO_2 \times \frac{1molO_2}{32.00gO_2} = 0.363molO_2 \]

\[ 0.363molO_2 \times \frac{1molCO_2}{2molO_2} = 0.181molCO_2 \]

\[ 0.181molCO_2 \times \frac{44.01gCO_2}{1molCO_2} = 7.98gCO_2 \]

\[ \frac{5.38\ g\ CO_2}{7.98\ g\ CO_2} \times 100\% = 67.4\%\ yield \]
8. Determine the mass of water vapor you would expect to form (and the percent yield) in the reaction between 15.8 g of NH$_3$ and excess oxygen to produce water and nitric oxide (NO). The mass of water actually formed is 21.8 g.

Equation: $4\text{NH}_3 + 5\text{O}_2 \rightarrow 6\text{H}_2\text{O} + 4\text{NO}$

Before: 0.928 XS 0 0
Change: 

After:
8. Determine the mass of water vapor you would expect to form (and the percent yield) in the reaction between 15.8 g of NH₃ and excess oxygen to produce water and nitric oxide (NO). The mass of water actually formed is 21.8 g.

Equation: 4NH₃ + 5O₂ → 6H₂O + 4NO

<table>
<thead>
<tr>
<th>Before:</th>
<th>Change:</th>
<th>After:</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.928 XS 0</td>
<td>- 0.928 -1.16</td>
<td>0 XS 1.39</td>
</tr>
<tr>
<td>0</td>
<td>0.928</td>
<td>0.928</td>
</tr>
</tbody>
</table>
Calculations

\[
15.8 \text{g} \text{NH}_3 \times \frac{1 \text{mol} \text{NH}_3}{17.04 \text{g} \text{NH}_3} = 0.928 \text{mol} \text{NH}_3
\]

\[
0.928 \text{mol} \text{NH}_3 \times \frac{6 \text{mol} \text{H}_2\text{O}}{4 \text{mol} \text{NH}_3} = 1.39 \text{mol} \text{H}_2\text{O}
\]

\[
1.39 \text{mol} \text{H}_2\text{O} \times \frac{18.02 \text{g} \text{H}_2\text{O}}{1 \text{mol} \text{H}_2\text{O}} = 25.1 \text{g} \text{H}_2\text{O}
\]

\[
\frac{21.8 \text{g} \text{H}_2\text{O}}{25.1 \text{g} \text{H}_2\text{O}} \times 100\% = 86.9\% \text{ yield}
\]