P.T.

1 mole = molar mass in grams = \( 6.02 \times 10^{23} \) atoms or molecules

Convert

77.0 g Na to mol

\[ 77.0 \text{g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{g Na}} = 3.35 \text{ mol Na} \]

How many atoms is this

\[ 77.0 \text{ g Na} \times \frac{6.02 \times 10^{23} \text{ atoms Na}}{22.99 \text{ g Na}} = 2.02 \times 10^{24} \text{ atoms Na} \]
for molecules

\[
1 \text{ mol} = \text{ molar mass in g} = 6.02 \times 10^{23} \text{ molecules}
\]

\[
\text{(add average mass for each atom) = (formula units)}
\]

How many molecules \( \text{H}_2\text{O} \) are there in a 0.500 L bottle of water?

\[
\begin{align*}
0.500 \text{ L} & \times 1000 \text{ g H}_2\text{O} \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \\
&= 1.67 \times 10^{25} \text{ molecules H}_2\text{O}
\end{align*}
\]
Convert:

2.54 x 10^{22} \text{ atoms of Cr to mol}

4.32 \text{ mol NaCl to grams}

0.786 \text{ mol carbon dioxide to grams}

2.67 \text{ g lithium carbonate to mol}

1.000 \text{ atom of C-12 to grams}
Convert:

2.54 x 10^{22} \text{ atoms of Cr} \text{ to mol}

\[
2.54 \times 10^{22} \text{ atoms Cr} \times \frac{1 \text{ mol Cr}}{6.02 \times 10^{23} \text{ atoms Cr}} = 0.0422 \text{ mol Cr}
\]

4.32 mol NaCl \text{ to grams}

\[
4.32 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 252 \text{ g Na}
\]

0.786 mol carbon dioxide \text{ to grams}

\[
0.786 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 34.6 \text{ g CO}_2
\]

2.67 g lithium carbonate \text{ to mol}

\[
2.67 \text{ g Li}_2\text{CO}_3 \times \frac{1 \text{ mol Li}_2\text{CO}_3}{73.89 \text{ g Li}_2\text{CO}_3} = 0.0361 \text{ mol Li}_2\text{CO}_3
\]

1 atom of C-12 \text{ to grams}

\[
1.00 \text{ atom }^{12}\text{C} \times \frac{12.00 \text{ g }^{12}\text{C}}{6.02 \times 10^{23} \text{ atoms}} = 1.99 \times 10^{-23} \text{ g}
\]
How many carbonate ions are there in 2.23 g of iron (III) carbonate?

How many oxygen atoms are there in the same sample?
How many carbonate ions are there in 2.23 g of iron (III) carbonate?

\[
2.23 \text{ g Fe}_2(\text{CO}_3)_3 \times \frac{1 \text{ mol Fe}_2(\text{CO}_3)_3}{291.73 \text{ g Fe}_2(\text{CO}_3)_3} \times \frac{3 \text{ mol CO}_3^{2-}}{1 \text{ mol Fe}_2(\text{CO}_3)_3} \times \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol CO}_3^{2-}} = 1.38 \times 10^{22} \text{ ions CO}_3^{2-}
\]

How many oxygen atoms are there in the same sample?

\[
1.38 \times 10^{22} \text{ ions CO}_3^{2-} \times \frac{3 \text{ O atoms}}{1 \text{ CO}_3^{2-}} = 4.14 \times 10^{22} \text{ O atoms}
\]
% Composition and empirical formulas

To find the % composition of an element in a compound, do the following:

\[
\% \text{ comp} = \frac{\text{mass element in compound}}{\text{total mass}}
\]

Find the % composition of ammonium sulfate.

\[
\begin{align*}
\% \text{N} &= \frac{2(14.01)\text{ g N}}{132.17\text{ g (NH}_4\text{)}_2\text{SO}_4} = 0.2120 \text{ or } 21.20\% \text{ N} \\
\% \text{H} &= \frac{8(1.01)\text{ g H}}{132.17\text{ g (NH}_4\text{)}_2\text{SO}_4} = 0.0611 \text{ or } 6.11\% \text{ H} \\
\% \text{S} &= \frac{32.07\text{ g S}}{132.17\text{ g (NH}_4\text{)}_2\text{SO}_4} = 0.2426 \text{ or } 24.26\% \text{ S} \\
\% \text{O} &= \frac{4(16.00)\text{ g O}}{132.17\text{ g (NH}_4\text{)}_2\text{SO}_4} = 0.4842 \text{ or } 48.42\% \text{ O}
\end{align*}
\]
We can find empirical formulas given % composition or the mass of each element in a sample of the compound.

- % to mass
- mass to moles
- / by smallest
- x 'till whole

Vitamin C is composed of 40.92 % C, 4.58 % H, and 54.50 % O. Find the empirical formula.

1) % to mass (take a 100g sample)
2) mass to mol
3) / by smallest
4) x 'till whole

\[ \frac{40.92 \text{ g C}}{12.01 \text{ g C}} = 3.407 \text{ mol C} \]
\[ \frac{4.58 \text{ g H}}{1.008 \text{ g H}} = 4.53 \text{ mol H} \]
\[ \frac{54.50 \text{ g O}}{16.00 \text{ g O}} = 3.406 \text{ mol O} \]

\[ \frac{3.407 \text{ mol C}}{3.406} : \frac{4.53 \text{ mol H}}{3.406} : \frac{3.406 \text{ mol O}}{3.406} \]

1.000 mol C : 1.33 mol H : 1.000 mol O
\[ x \text{ all by } 3 \]
3 mol C : 4 mol H : 3 mol O

Empirical formula is \( \text{C}_3\text{H}_4\text{O}_3 \)

The molar mass of vitamin C is 176.14 g/mol. What is the molecular formula?

\[ \frac{\text{molar mass compound}}{\text{empirical formula mass}} = \text{whole #} \]

\[ \frac{176.14 \text{ g/mol}}{88.07 \text{ g/mol}} \approx 2 \]

so the molecular formula is \( \text{C}_6\text{H}_8\text{O}_6 \)
3.50 a. What is the empirical formula of the compound with the following composition?
40.1 % C
6.6 % H
53.3 % O

\[
\begin{align*}
40.1 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} &= 3.34 \text{ mol C} \\
6.6 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} &= 6.5 \text{ mol H} \\
53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 3.33 \text{ mol O} \\
\end{align*}
\]

\[
\frac{3.34 \text{ mol C}}{3.33} : \frac{6.5 \text{ mol H}}{3.33} : \frac{3.33 \text{ mol O}}{3.33}
\]

1.00 mol C : 2.0 mol H : 1.00 mol O

\[
\text{of C}_2\text{H}_2\text{O}
\]
Formula of a hydrate lab:

Some ionic salts are hydrates. This means that there are water molecules in the ionic lattice.

If we heat a hydrate, we can thermally decompose it (all of the water will driven out of the lattice) into the anhydrous salt.

\[ \text{BaCl}_2 \cdot 2\text{H}_2\text{O} \xrightarrow{\Delta} \text{BaCl}_2\text{(s)} + 2\text{H}_2\text{O}_{(g)} \]

Your goal is to find what is the mole ratio between the ionic compound and the water.

\[ \text{MgSO}_4 \cdot x\text{H}_2\text{O} \]

\[ \text{CuSO}_4 \cdot x\text{H}_2\text{O} \]
Q: How do you find the theoretical % H2O?
A: Mass H2O/Mass hydrated salt
<table>
<thead>
<tr>
<th></th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Empty crucible</td>
<td>21.320 g</td>
</tr>
<tr>
<td>Crucible and Mg (before heating)</td>
<td>21.432 g</td>
</tr>
<tr>
<td>Mass of magnesium</td>
<td>0.112 g</td>
</tr>
<tr>
<td>Crucible and combustion product</td>
<td>21.468 g</td>
</tr>
<tr>
<td>Mass of combustion product</td>
<td>0.148 g</td>
</tr>
</tbody>
</table>
1. Determine the mass of the magnesium used. (3 data)
   \[21.432 \text{ g} - 21.320 \text{ g} = 0.112 \text{ g} \text{ Mg}\]

2. Determine the number of moles of magnesium used.
   \[0.112 \text{ g} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g} \text{ Mg}} = 4.61 \times 10^{-3} \text{ mol Mg}\]

3. Determine the mass of the magnesium oxide (combustion product) formed. (5 data)
   \[21.468 \text{ g} - 21.320 \text{ g} = 0.148 \text{ g} \text{ MgO}\]

4. Determine the mass of oxygen that combined with the magnesium.
   \[0.148 \text{ g} \text{ MgO} - 0.112 \text{ g} \text{ Mg} = 0.036 \text{ g O}\]

5. Calculate the number of moles of oxygen that reacted.
   \[0.036 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.25 \times 10^{-3} \text{ mol O}\]

6. Calculate the ratio between moles of magnesium used and moles of oxygen used. Express this in a small whole number form.
   \[\frac{4.61 \times 10^{-3} \text{ mol Mg}}{2.25 \times 10^{-3} \text{ mol O}} = \frac{2.05 \text{ mol Mg}}{1 \text{ mol O}}\]

7. Based on your experimental data, what is the empirical formula of magnesium oxide formed.
   \[\text{Mg}_2\text{O}\]
Allicin is the compound responsible for the characteristic smell of garlic. An analysis of the compound gives the following percent composition by mass: C: 44.4%; H: 6.21%; S: 39.5%; O: 9.86%. Calculate its empirical formula. (Type your answer using the format CH2 for CH₂.)

What is its molecular formula given that its molar mass is about 162 g?
Allicin is the compound responsible for the characteristic smell of garlic. An analysis of the compound gives the following percent composition by mass: C: 44.4%; H: 6.21%; S: 39.5%; O: 9.86%. Calculate its empirical formula. (Type your answer using the format CH2 for CH2.)

\[
\begin{align*}
44.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} &= 3.70 \text{ mol C} \\
6.21 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} &= 6.15 \text{ mol H} \\
39.5 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} &= 1.23 \text{ mol S} \\
9.86 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 0.616 \text{ mol O}
\end{align*}
\]

\[
\begin{align*}
3.70 \text{ mol C} : 6.15 \text{ mol H} : 1.23 \text{ mol S} : 0.616 \text{ mol O}
\end{align*}
\]

\[
\begin{align*}
0.616 : 0.616 : 0.616 : 0.616
\end{align*}
\]

\[
\begin{align*}
6.00 \text{ mol C} : 9.98 \text{ mol H} : 2.00 \text{ mol S} : 1.00 \text{ mol O}
\end{align*}
\]

\[
\begin{align*}
e_f. \quad \text{C}_6 \text{H}_{10} \text{S}_2 \text{O}
\end{align*}
\]

\[
\begin{align*}
\text{molar mass} &= \frac{162}{162.3} \approx 1 \\
\text{e.f. mass} &= \frac{162}{162.3} \\
\therefore \text{e.f. is the m.f.} \quad \text{C}_6 \text{H}_{10} \text{S}_2 \text{O}
\end{align*}
\]
C: 44.4%; H: 6.21%; S: 39.5%; O: 9.86%.

\[ \frac{44.4\text{ g C}}{12.01\text{ g C}} \times \frac{1\text{ mol C}}{1\text{ mol C}} = 3.70\text{ mol C} \]
\[ \frac{6.21\text{ g H}}{1.01\text{ g H}} \times \frac{1\text{ mol H}}{1\text{ mol H}} = 6.15\text{ mol H} \]
\[ \frac{39.5\text{ g S}}{32.07\text{ g S}} \times \frac{1\text{ mol S}}{1\text{ mol S}} = 1.23\text{ mol S} \]
\[ \frac{9.86\text{ g O}}{16.00\text{ g O}} \times \frac{1\text{ mol O}}{1\text{ mol O}} = 0.616\text{ mol O} \]

\[ \frac{3.70\text{ mol C}}{1.616} : \frac{6.15\text{ mol H}}{1.616} : \frac{1.23\text{ mol S}}{1.616} : \frac{0.616\text{ mol O}}{1.616} \]

\[ = 6.01\text{ mol C} : 9.98\text{ mol H} : 2.00\text{ mol S} : 1.00\text{ mol O} \]

\[ \text{C}_6\text{H}_{10}\text{S}_2\text{O} \]

What is its molecular formula given that its molar mass is about 162 g?

\[ \frac{162\text{ g/mol}}{162.30\text{ g/mol}} = 1 \]

**: et. is the m.f.**
Combust 11.5 g ethanol
Collect 22.0 g CO₂ and 13.5 g H₂O

\[
\% \text{ comp of CO}_2 = \frac{22.0 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2}}{6.00 \text{ g C}} = 6.00 \text{ g C}
\]

\[
13.5 \text{ g H}_2\text{O} \times \frac{2.02 \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 1.51 \text{ g H}
\]

mass O = 11.5 g - (6.00 + 1.51) in ethanol = 4.00 g O

Now we can find the e.f.

\[
6.00 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.500 \text{ mol C}
\]

\[
1.51 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 1.50 \text{ mol H}
\]

\[
4.00 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.250 \text{ mol O}
\]

\[
\frac{0.500 \text{ mol C}}{0.250} : \frac{1.50 \text{ mol H}}{0.250} : \frac{0.250 \text{ mol O}}{0.250}
\]

2 mol C : 6 mol H : 1 mol O

\[
\text{e.f. C}_2\text{H}_6\text{O}
\]
Lysine, an essential amino acid in the human body, contains C, H, O, and N. In one experiment, the complete combustion of 2.175 g of lysine gave 3.94 g CO₂ and 1.89 g H₂O. In a separate experiment, 1.873 g of lysine gave 0.436 g NH₃. (Type your answer using the format CO₂ for CO₂.) (a) Calculate the empirical formula of lysine.

(b) The approximate molar mass of lysine is 150 g. What is the molecular formula of the compound?
Hemoglobin (C_{298}H_{466}N_{81}O_{82}S_{8}Fe_{4}) is the oxygen carrier in blood.

(a) Calculate its molar mass.

\[
\text{g mL}^{-1} = 65332.28
\]

(b) An average adult has about 5.0 liters of blood. Every milliliter of blood has approximately \(5.0 \times 10^9\) erythrocytes, or red blood cells, and every red blood cell has about \(2.8 \times 10^8\) hemoglobin molecules. Calculate the mass of hemoglobin molecules in grams in an average adult.

\[
5000 \text{ mL blood} \times \frac{5.0 \times 10^9 \text{ eryth}}{1 \text{ mL}} \times \frac{2.8 \times 10^8 \text{ hem}}{1 \text{ eryth}} \times \frac{65300 \text{ g}}{6.02 \times 10^{23} \text{ hem}} = \text{g hemoglobin}
\]
The formula for rust can be represented by Fe₂O₃. How many moles of Fe are present in 38.8 g of the compound?

\[
38.8 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} = 0.486 \text{ mol Fe}
\]
Calculate the number of cations and anions in each of the following compounds.

(b) 5.30 g of Na$_2$SO$_4$

\[
\begin{align*}
5.30 \text{g Na}_2\text{SO}_4 & \times \frac{1 \text{mol Na}_2\text{SO}_4}{192.06 \text{g Na}_2\text{SO}_4} \times \frac{2 \text{mol Na}^+}{1 \text{mol Na}_2\text{SO}_4} \times \frac{6.02 \times 10^{23} \text{Na}^+ \text{ions}}{1 \text{mol Na}^+} = \\
& = 4.49 \times 10^{23} \text{Na}^+ \text{ions} \\

5.30 \text{g Na}_2\text{SO}_4 & \times \frac{1 \text{mol Na}_2\text{SO}_4}{192.06 \text{g Na}_2\text{SO}_4} \times \frac{1 \text{mol SO}_4^{2-}}{1 \text{mol Na}_2\text{SO}_4} \times \frac{6.02 \times 10^{23} \text{SO}_4^{2-} \text{ions}}{1 \text{mol SO}_4^{2-}} = \\
& = 2.25 \times 10^{23} \text{SO}_4^{2-} \text{ions} 
\end{align*}
\]
The natural abundances of the two stable isotopes of hydrogen (hydrogen and deuterium) are $^1_1$H: 99.985\% and $^2_1$H: 0.015\%. Assume that water exists as either H$_2$O or D$_2$O. Calculate the number of D$_2$O molecules in exactly 500. mL of water. (Density = 1.00 g/mL.)

\[
500 \text{ mL H}_2\text{O} \times \frac{1 \text{ g}}{1 \text{ mL}} = 500 \text{ g H}_2\text{O}
\]

\[
500 \text{ g H}_2\text{O} \times \frac{6.02 \times 10^{23} \text{ molecules}}{18.02 \text{ g}} = 1.67 \times 10^{23} \text{ molecules}
\]

\[
(0.00015)(1.67 \times 10^{23}) = 2.5 \times 10^{21} \text{ molecules D}_2\text{O}
\]
How many grams of sulfur (S) are needed to react completely with 464 g of mercury (Hg) to form HgS?

\[ \frac{200.57 \text{ g Hg}}{232.66 \text{ g HgS}} = 86.2\% \text{ Hg} \]

\[ 0.862 \times 464 \text{ g Hg} = 538 \text{ g HgS} \]

\[ 74 \text{ g S} = 538 \text{ g HgS} - 464 \text{ g Hg} \]

\[ 464 \text{ g Hg} \times \frac{32.07 \text{ g S}}{200.57 \text{ g Hg}} = 74.2 \text{ g S} \]
Balancing Equations

**Word equation**

magnesium + hydrochloric acid ➔ magnesium chloride + hydrogen

copper (II) nitrate + iron ➔ iron (III) nitrate + copper

propene (C₃H₆) combusts in oxygen to form carbon dioxide and water

iron (III) oxide is reduced by carbon monoxide to make iron and carbon dioxide
Balancing Equations

Word equation

magnesium + hydrochloric acid → magnesium chloride + hydrogen

\[ \text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \]

copper (II) nitrate + iron → iron (III) nitrate + copper

\[ 3\text{Cu(NO}_3\text{)}_2 + 2\text{Fe} \rightarrow 2\text{Fe(NO}_3\text{)}_3 + 3\text{Cu} \]

propene (C₃H₆) combusts in oxygen to form carbon dioxide and water

\[ 2\text{C}_3\text{H}_6 + 9\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} \]

iron (III) oxide is reduced by carbon monoxide to make iron and carbon dioxide

\[ \text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2 \]
Stoichiometry--Allows you to predict the amounts of products that will form or reactants that are needed in a chemical reaction.

In order to do stoichiometry, YOU NEED A BALANCED CHEMICAL EQUATION.
How many grams of silver will form in the reaction if 10.0 g of copper react completely?

\[ 2 \text{AgNO}_3 + \text{Cu} \rightarrow \text{Cu(NO}_3)_2 + 2 \text{Ag} \]

\[ \text{10.0 g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Cu}} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} = 33.9 \text{g Ag} \]

How many grams of silver nitrate are required to react with the 10.0 g Cu?

\[ 2 \text{AgNO}_3 + \text{Cu} \rightarrow \text{Cu(NO}_3)_2 + 2 \text{Ag} \]

\[ \text{10.0 g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mol AgNO}_3}{1 \text{ mol Cu}} \times \frac{169.89 \text{ g AgNO}_3}{1 \text{ mol AgNO}_3} = 53.5 \text{g AgNO}_3 \]
How many grams of hydrogen gas will be produced if 0.250 mol of sodium metal reacts with excess water? (The other product is sodium hydroxide)

\[ 2 \text{Na} + 2\text{H}_2\text{O} \rightarrow \text{H}_2 + 2\text{NaOH} \]

\[
0.250\text{mol} \times \frac{1\text{mol}\text{H}_2}{2\text{mol}\text{Na}} \times \frac{2.02\text{g}\text{H}_2}{1\text{mol}\text{H}_2} = 0.253\text{g}\text{H}_2
\]
Nitrous oxide (N₂O) is also called "laughing gas." It can be prepared by the thermal decomposition of ammonium nitrate (NH₄NO₃). The other product is H₂O. The balanced equation for this reaction is given below.

\[ \text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + 2 \text{H}_2\text{O} \]

How many grams of N₂O are formed if 0.46 mole of NH₄NO₃ is used in the reaction?

\[ 0.46 \text{ mol NH}_4\text{NO}_3 \times \frac{1 \text{ mol N}_2\text{O}}{1 \text{ mol NH}_4\text{NO}_3} \times \frac{44.02 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}} = 20. \]
A common laboratory preparation of oxygen gas is the thermal decomposition of potassium chlorate (KClO₃). Assuming complete decomposition, calculate the number of grams of O₂ gas that can be obtained from 46.0 g of KClO₃. (The products are KCl and O₂.)

\[
2\text{KClO}_3 \rightarrow 3\text{O}_2 + 2\text{KCl}
\]

\[
46.0 \text{g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 18.0 \text{ g O}_2
\]
Limiting Reactant

When the amounts of two or more reactants are provided in the problem it is a limiting reactant problem.

\[ 2 \text{ Pb} + 10 \text{ J} + 2 \text{ Br} \rightarrow 1 \text{ Pb}_2\text{J}_10 \text{ Br}_2 \]

How many peanut butter and jelly sandwiches can you make if you have 13 slices of bread in the cupboard, 5 tbsp of Pb in the jar and 20 tbsp. of jelly in the jar?

\[ \frac{13 \text{ Br}}{2 \text{ Br}} \times \frac{1 \text{ sandwich}}{2 \text{ Br}} = 6.5 \text{ sandwiches} \]

\[ \frac{5 \text{ Pb}}{2 \text{ Pb}} \times \frac{1 \text{ sandwich}}{2 \text{ Pb}} = 2.5 \text{ sandwiches} \]

\[ \frac{20 \text{ J}}{10 \text{ J}} = 2.0 \text{ sandwiches} \]

Jelly is the limiting reactant. The maximum possibly produced theoretical yield is: 2.0 sandwiches.
10.0 grams of silver nitrate are added to 10.0 grams of copper. How many grams silver is produced? (Copper II nitrate is the other product) What is the limiting reactant? How much of the excess reactant remains?

\[
\text{AgNO}_3 + \text{Cu} \rightarrow \text{Cu(NO}_3\text{)}_2 + 2\text{Ag}
\]

\[
10.0_\text{g} \times \frac{1\text{mol AgNO}_3}{169.88_\text{g}} \times \frac{2\text{mol Ag}}{2\text{mol AgNO}_3} \times \frac{107.87_\text{g}}{1\text{mol Ag}} = 6.35_\text{g Ag}
\]

\[
10.0_\text{g} \times \frac{1\text{mol Cu}}{63.55_\text{g}} \times \frac{2\text{mol Ag}}{1\text{mol Cu}} \times \frac{107.87_\text{g}}{1\text{mol Ag}} = 33.9_\text{g Ag}
\]

AgNO\text{_3} is L.R.

Cu in in excess

6.35 g AgNO\text{_3} is theoretical yield.

\[
10.0_\text{g} \times \frac{1\text{mol AgNO}_3}{169.88_\text{g}} \times \frac{1\text{mol Cu}}{2\text{mol AgNO}_3} \times \frac{63.55_\text{g}}{1\text{mol Cu}} = 1.87_\text{g Cu}
\]

10.0 g - 1.87 g = 8.1 g Cu remain.
The compound cisplatin, Pt(NH$_3$)$_2$Cl$_2$, has been studied extensively as an antitumor agent [see Journal of Chemical Education: 54(1977):739]. Cisplatin is synthesized as follows.

K$_2$PtCl$_4$(aq) + 2 NH$_3$(aq) $\rightarrow$ Pt(NH$_3$)$_2$Cl$_2$(s) + 2 KCl (aq)

What mass of cisplatin can be made from 114 g of K$_2$PtCl$_4$ and sufficient NH$_3$?

\[\text{NOT A L.R problem}\]
Mercury and bromine will react with each other to produce mercury(II) bromide.

\[ \text{Hg}(l) + \text{Br}_2(l) \rightarrow \text{HgBr}_2(s) \]

(a) What mass of \( \text{HgBr}_2 \) is produced from the reaction of 9.93 g Hg and 10.9 g \( \text{Br}_2 \)?

\[
\begin{align*}
9.93 \text{ g Hg} & \times \frac{1 \text{ mol Hg}}{200.54 \text{ g Hg}} \times \frac{1 \text{ mol HgBr}_2}{1 \text{ mol Hg}} \times \frac{178.5 \text{ g HgBr}_2}{1 \text{ mol HgBr}_2} = 17.85 \text{ g HgBr}_2 \\
10.9 \text{ g Br}_2 & \times \frac{1 \text{ mol Br}_2}{159.80 \text{ g Br}_2} \times \frac{1 \text{ mol HgBr}_2}{1 \text{ mol Br}_2} \times \frac{178.5 \text{ g HgBr}_2}{1 \text{ mol HgBr}_2} = 24.6 \text{ g HgBr}_2
\end{align*}
\]

What mass of which reactant is left unreacted?

\[
\begin{align*}
9.93 \text{ g Hg} & \times \frac{1 \text{ mol Hg}}{200.54 \text{ g Hg}} \times \frac{1 \text{ mol Br}_2}{1 \text{ mol Hg}} \times \frac{159.80 \text{ g Br}_2}{1 \text{ mol Br}_2} = 7.91 \text{ g Br}_2 \\
10.9 \text{ g Br}_2 - 7.91 \text{ g Br}_2 = 3.0 \text{ g Br}_2
\end{align*}
\]

(b) What mass of \( \text{HgBr}_2 \) is produced from the reaction 6.14 mL of mercury (density = 13.6 g/mL) and 6.14 mL of bromine (density = 3.10 g/mL)?
What you finish with:

\[
\% \text{ Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%
\]

STOICHIOMETRY
Hydrogen fluoride is used in the manufacture of Freons (which destroy ozone in the stratosphere) and in the production of aluminum metal. It is prepared by the following reaction.

$$\text{CaF}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + 2 \text{HF}$$

In one process 6.00 kg of CaF$_2$ are treated with an excess of H$_2$SO$_4$ and yield 2.86 kg of HF. Calculate the percent yield of HF.

\[
6000 \text{g} \times \frac{1 \text{mol} \text{CaF}_2}{78.08 \text{g} \text{CaF}_2} \times \frac{2 \text{mol} \text{HF}}{1 \text{mol} \text{CaF}_2} \times \frac{20.018 \text{g} \text{HF}}{1 \text{mol} \text{HF}} = 3080 \text{g} \text{HF}
\]

\[
\% \text{ Yield} = \frac{2.86 \text{ kg} \text{HF}}{3.08 \text{ kg} \text{HF}} \times 100\% = 92.9\%
\]
The yield in the Haber process is 71.2 %. How many grams of nitrogen are needed to produce 1.00 kg of ammonia?

\[
\begin{align*}
\text{N}_2 + 3\text{H}_2 & \rightarrow 2\text{NH}_3 \\
0.712 &= \frac{1.00 \text{ kg}}{\text{theoret}} \\
\text{theoret} &= \frac{1.00}{0.712} = 1.40 \text{ kg NH}_3 \\
1.40 \times 10^{-3} \text{g NH}_3 & \times \frac{1 \text{ mol NH}_3}{17.04 \text{ g}} \times \frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3} \times \frac{28.02 \text{ g N}_2}{1 \text{ mol}} \\
&= 1150 \text{ g N}_2 \\
&= 1.15 \text{ kg N}_2
\end{align*}
\]
Ethylene ($\text{C}_2\text{H}_4$), an important industrial organic chemical, 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?
Mercury and bromine will react with each other to produce mercury(II) bromide.

Hg(l) + Br₂(l) → HgBr₂(s)

(a) What mass of HgBr₂ is produced from the reaction of 9.93 g Hg and 10.9 g Br₂?

What mass of which reactant is left unreacted?

(b) What mass of HgBr₂ is produced from the reaction of 6.14 mL of mercury (density = 13.6 g/mL) and 6.14 mL bromine (density = 3.10 g/mL)?
The aspirin substitute, acetaminophen (C₈H₉O₂N), is produced by the following three-step syntheses.

I. C₆H₅O₃N(s) + 3 H₂(g) + HCl(aq) → C₆H₈ONCl(s) + 2 H₂O(l)  \[87\%\]

II. C₆H₈ONCl(s) + NaOH(aq) → C₆H₇ON(s) + H₂O(l) + NaCl(aq)  \[98\%\]

III. C₆H₇ON(s) + C₄H₆O₃(l) → C₈H₉O₂N(s) + HC₂H₃O₂(l)

The first two reactions have percent yields of 87% and 98% by mass, respectively. The overall reaction yields 3 mol of acetaminophen product for every 4 mol of C₆H₅O₃N reacted.

(a) What is the percent yield by mass for the overall process?

overall yield = \( \frac{3 \text{ mol C₈H₉O₂N}}{4 \text{ mol C₆H₇ON}} \times 100\% = 75\% \)

(b) What is the percent yield by mass of step III?

\( (0.87) (0.98)(x) = 0.75 \)

\( x = 0.879 \)
The production capacity for acrylonitrile (C$_3$H$_3$N) in the United States is over 2 million pounds per year. Acrylonitrile, the building block for polyacrylonitrile fibers and a variety of plastics, is produced from gaseous propylene, ammonia, and oxygen.

2 C$_3$H$_6$(g) + 2 NH$_3$(g) + 3 O$_2$(g) $\rightarrow$ 2 C$_3$H$_3$N(g) + 6 H$_2$O(g)

(a) What mass of acrylonitrile can be produced from a mixture of 1.04 kg propylene, 1.55 kg of ammonia, and 2.21 kg of oxygen?

(b) What mass of water is produced, and what masses of which starting materials are left in excess?

<table>
<thead>
<tr>
<th>mass of water</th>
<th>starting materials in excess</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>oxygen ✓</td>
</tr>
<tr>
<td></td>
<td>ammonia ✓</td>
</tr>
<tr>
<td></td>
<td>propylene ×</td>
</tr>
</tbody>
</table>

mass of propylene in excess

mass of ammonia in excess

mass of oxygen in excess
Industrial, nitric acid is produced by the Ostwald process represented by the following equations.

\[
4 \text{NH}_3(g) + 5 \text{O}_2(g) \rightarrow 4 \text{NO}(g) + 6 \text{H}_2\text{O}(l)
\]

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

\[
2 \text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(aq) + \text{HNO}_2(aq)
\]

What mass of NH\textsubscript{3} (in g) must be used to produce 1.00 ton of HNO\textsubscript{3} by the above procedure, assuming an 80 percent yield in each step? (1 ton = 2000 lb; 1 lb = 453.6 g.)

\[
1.95 \text{ tons HNO}_3 \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol NH}_3} \times \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}_2} \times \frac{4 \text{ mol H}_2\text{O}}{4 \text{ mol NO}_2} = 9.57 \times 10^5 \text{ g NH}_3
\]