Limiting Reactants and Percent Yield

1. Aluminum metal reacts with chlorine gas in a synthesis reaction.
   a. Write the balanced equation for this reaction.

   \[ 2 \text{Al(s)} + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3 \]

   b. If 15.2 g of aluminum reacts with 39.1 g of chlorine, identify the limiting reactant.

   \[
   \frac{15.2 \text{ g Al}}{27.27 \text{ g Al} \text{ mol}^{-1}} = 0.558 \text{ mol Al} \\
   \frac{39.1 \text{ g Cl}_2}{70.90 \text{ g Cl}_2 \text{ mol}^{-1}} = 0.551 \text{ mol Cl}_2 \\
   \text{mol Cl}_2 < \text{mol Al} \\
   \text{Al is limiting reactant (excess Cl}_2\text{)}
   \]

   c. Determine the mass in grams of the product formed.

   \[
   \frac{0.558 \text{ mol Al}}{1 \text{ mol AlCl}_3} \times 133.34 \text{ g AlCl}_3 = 75.12 \text{ g AlCl}_3
   \]

   d. Determine the mass in grams of excess reactant remaining when the reaction is complete.

   \[
   \frac{0.551 \text{ mol Cl}_2 - 0.558 \text{ mol Al}}{0.558 \text{ mol AlCl}_3} \times 39.1 \text{ g Cl}_2 = 49.02 \text{ g Cl}_2
   \]

2. \[ 3\text{Fe(s)} + 2\text{O}_2(g) \rightarrow \text{Fe}_3\text{O}_4(s) \]
   a. When 13.54 g of O\(_2\) is mixed with 12.21 g of Fe, which is the limiting reactant?

   \[
   \frac{13.54 \text{ g O}_2}{32 \text{ g O}_2 \text{ mol}^{-1}} = 0.423 \text{ mol O}_2 \\
   \frac{12.21 \text{ g Fe}}{56 \text{ g Fe} \text{ mol}^{-1}} = 0.218 \text{ mol Fe} \\
   \text{mol Fe < mol O}_2 \\
   \text{O}_2 \text{ is limiting reactant (excess Fe)}
   \]

   b. What mass in grams of iron oxide is produced?

   \[
   \frac{0.218 \text{ mol Fe}}{1 \text{ mol Fe}_3\text{O}_4} \times 159.71 \text{ g Fe}_3\text{O}_4 = 55.85 \text{ g Fe}_3\text{O}_4
   \]

   c. What mass in grams of excess reactant remains when the reaction is complete?

   \[
   \frac{0.423 \text{ mol O}_2 - 0.218 \text{ mol Fe}}{0.218 \text{ mol Fe}_3\text{O}_4} \times 32 \text{ g O}_2 = 4.61 \text{ g O}_2
   \]

   d. Kelly performed this reaction in a lab and made 15.88 g of Fe\(_3\)O\(_4\). What was her percent yield?

   \[
   \left( \frac{15.88 \text{ g Fe}_3\text{O}_4}{16.87 \text{ g Fe}_3\text{O}_4 \text{ yield}} \right) \times 100 = 94.0 \text{ % yield}
   \]
3. Diborane pentoxide is useful in devices such as respirators because it reacts with the dangerous gas carbon monoxide, CO, to produce relatively harmless CO₂ according to the following equation:

\[ \text{I}_2\text{O}_5 + 5\text{CO} \rightarrow \text{I}_2 + 5\text{CO}_2 \]

a. In testing a respirator, 2.00 g of carbon monoxide gas is passed through diborane pentoxide. Upon analyzing the results, it is found that 3.17 g of I₂ was produced. Calculate the percent yield of the reaction.

\[ \% = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \]

b. Assuming that the yield in part a resulted because some of the CO did not react, calculate the mass of CO that passed through.

\[ \frac{3.17 \text{ g I}_2}{255.8 \text{ g CO}_2} \times \frac{5 \text{ mol CO}}{1 \text{ mol I}_2} \times \frac{28.0 \text{ g CO}}{1 \text{ mol CO}} = 1.75 \text{ g CO reacted} \]

4. Cu(s) + Cl₂(g) → CuCl₂(s)

a. If 12.5 g of Cu reacts with excess chlorine, calculate the theoretical yield of CuCl₂.

\[ \frac{12.5 \text{ g Cu}}{63.5 \text{ g Cu}} \times \frac{1 \text{ mol Cu}}{1 \text{ mol CuCl₂}} \times \frac{127 \text{ g CuCl₂}}{1 \text{ mol CuCl₂}} = 26.4 \text{ g CuCl₂} \]

b. If only 25.4 g of CuCl₂ was produced, what is the percent yield?

\[ \frac{25.4 \text{ g CuCl₂}}{26.4 \text{ g CuCl₂}} \times 100 = 96.2 \% \]

5. Sodium hyperchlorite, NaClO, the main ingredient in household bleach, is produced by bubbling chlorine gas through a strong lye (sodium hydroxide, NaOH) solution. The following equation shows the reaction that occurs.

\[ 2 \text{NaOH} + \text{Cl}_2 \rightarrow \text{NaCl} + \text{NaClO} + \text{H}_2\text{O} \]

a. What is the percent yield of the reaction if 1.2 kg of Cl₂ reacts to form 0.90 kg of NaClO?

\[ \frac{1.2 \text{ kg Cl}_2}{70.9 \text{ g Cl}_2} \times \frac{1 \text{ mol Cl}_2}{1 \text{ mol NaClO}} \times \frac{74.4 \text{ g NaClO}}{1 \text{ mol NaClO}} = 1259.9 \text{ g NaClO} \]

\[ \frac{900 \text{ g NaClO}}{1259.9 \text{ g NaClO}} \times 100 = 71.4 \% \]

6. The percent yield for the reaction PCl₅ + Cl₂ → PCl₃ is 83.2%. What mass of PCl₃ is expected from the reaction of 73.7 g of PCl₅ with excess chlorine?

\[ \frac{73.7 \text{ g PCl₅}}{137.5 \text{ g PCl₅}} \times \frac{1 \text{ mol PCl₅}}{1 \text{ mol PCl₃}} \times \frac{208.2 \text{ g PCl₃}}{1 \text{ mol PCl₃}} = 111.75 \text{ g PCl₃} \]

\[ \frac{73.7 \text{ g PCl₅}}{137.5 \text{ g PCl₅}} \times \frac{1 \text{ mol PCl₅}}{1 \text{ mol PCl₃}} \times \frac{208.2 \text{ g PCl₃}}{1 \text{ mol PCl₃}} = 111.75 \text{ g PCl₃} \]
Conservation of Mass Worksheet

Example:
A compound containing carbon and hydrogen is analyzed. When a 1.2543 gram sample is burned completely in excess oxygen, 3.671 grams of CO₂(g) is formed. What is the empirical formula of this compound?

a) Write a skeleton equation for this reaction. If you do not know the formula for every compound, put information about what you do know in parenthesis.

\[ \text{C}_x \text{H}_y + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

(1.2543 g) (excess) \rightarrow 3.671 g)

b) How many grams of carbon are in 3.671 g of CO₂? Where did all of this carbon come from? How much carbon was in the original sample of the unknown compound?

\[ \text{mol CO}_2 \times \frac{12 \text{ g C}}{1 \text{ mol CO}_2} \times \frac{1 \text{ mol C}}{12 \text{ g C}} = 0.106 \text{ g C} \]

You could also use percent composition.

All the carbon came from the hydrogen reaction. There was only C in the original unknown compound.

c) What is the percent composition of carbon and hydrogen in the unknown compound?

\[ \frac{12.54 \text{ g C}}{12.54 \text{ g C} + 1.00 \text{ g H}} = 79.17 \% \text{ Carbon} \]

\[ \frac{1.00 \text{ g H}}{12.54 \text{ g C} + 1.00 \text{ g H}} = 20.83 \% \text{ Hydrogen} \]

d) Determine the empirical formula of this compound, now that you know the percent composition.

\[ \frac{79.17 \text{ g C}}{12 \text{ g C}} = \frac{6.64 \text{ mol C}}{1 \text{ mol C}} \]

\[ \frac{20.83 \text{ g H}}{1 \text{ mol H}} = \frac{2.01 \text{ mol H}}{1 \text{ mol H}} \]

\[ \text{Empirical formula } = \text{CH}_3 \]

e) The molar mass of the compound is determined to be 30.08 g/mol. What is the correct molecular formula of the unknown compound?

\[ \frac{30.08 \text{ g}}{M_{	ext{Empirical}}} = \frac{15.04 \text{ g}}{2 \text{ mol H}} \]

\[ 2 \text{ mol } \text{C}_2 \text{H}_6 \]

ONE
A compound containing carbon and hydrogen is analyzed. Combustion of a 16.81 g sample of this compound produces 38.91 g of CO₂(g). What is the empirical formula of this compound? What is the molecular formula if the molar mass of the compound is determined to be 38.03 g/mol?

TWO
A compound containing carbon and hydrogen is analyzed. Combustion of a 0.213 g sample of this compound yields 0.2132 g of water. What is the empirical formula of this compound? What is the molecular formula if the molar mass of the compound is 81.15 g/mol?

THREE
A compound containing carbon, hydrogen, and oxygen is analyzed. Combustion of a 3.4 g sample of this compound produces 8.79 g of CO₂(g) and 2.7843 g of water vapor. What is the empirical formula of this compound? What is the molecular formula if the molar mass of the compound is 44.06 g/mol?
conservation of mass worksheet continued...

1. \( C_2H_4 + O_2 \rightarrow CO_2 + H_2O \)

\[
16.81g \quad 38.9g
\]

\[
\frac{38.9g \text{ CO}_2}{44g \text{ CO}_2} = \frac{1 \text{ mol CO}_2}{1 \text{ mol CO}_2} = 0.887 \text{ mol CO}_2
\]

\[
\frac{12g \text{ C}}{1 \text{ mol CO}_2} = \frac{10.62g \text{ C}}{0.887 \text{ mol CO}_2} = 11.92g \text{ C}
\]

\[
\frac{2g \text{ H}_2}{1 \text{ mol CO}_2} = \frac{3.643g \text{ H}}{0.887 \text{ mol CO}_2} = 4.119g \text{ H}
\]

\[
\text{C} = \frac{10.62g \text{ C}}{16.81g \text{ C}_2H_4} \times 100 = 63.27\% \text{ C}
\]

\[
\text{H} = 100\% - 63.27 = 36.83\% \text{ H}
\]

\[
\frac{63.2g \text{ C}}{1 \text{ mol C}} = \frac{5.27 \text{ mol C}}{5.27 \text{ mol C}} = 1 \text{ mol C}
\]

\[
\frac{36.83g \text{ H}}{1 \text{ mol H}} = \frac{36.43g \text{ H}}{0.887 \text{ mol C}} = 7 \text{ mol H}
\]

\[
\text{CH}_7 = 19.089 \text{ g} \quad \frac{38.0g}{19.089 \text{ g/mol}} = 2 \text{ (CH}_7\text{) = [C}_2\text{H}_5\text{]}
\]

2. \( C_2H_4 + O_2 \rightarrow CO_2 + H_2O \)

\[
0.213g \quad 0.2132g
\]

\[
\frac{0.2132g \text{ H}_2O}{1 \text{ mol H}_2O} = \frac{2 \text{ mol H}}{1 \text{ mol H}_2O} = \frac{1.019g \text{ H}}{0.213g \text{ C}_2H_4}
\]

\[
0.0239g \text{ H} \times 100 = 11.27\% \text{ H}
\]

\[
\frac{11.29g \text{ H}}{1 \text{ mol H}} = \frac{7.39 \text{ mol C}_2H_5}{1 \text{ mol H}}
\]

\[
\text{H} = \frac{11.27\%}{100} = 88.87\% \text{ C}
\]

\[
\frac{88.89g \text{ C}}{12.01g \text{ C}} = \frac{7.39 \text{ mol C}_2H_5}{7.39 \text{ mol C}_2H_5}
\]

\[
\text{C} = \frac{1 \text{ mol C}}{2 \text{ mol C}} = 2 \text{ mol C}
\]

\[
\text{C}_2H_5 = 27.05 \text{ g} \quad \frac{81.15g}{27.05g} = 3 \text{ (C}_2\text{H}_5\text{) = [C}_2\text{H}_5\text{]}
\]

molecular formula: \( \text{C}_6\text{H}_6 \)
\[ C_{x}H_{y}O_{z} + O_{2} \rightarrow CO_{2} + H_{2}O \]

\[
\begin{align*}
6.79g CO_{2} & \div 1mol CO_{2} \div 1mol C = 12 g C = 1.859 g C \\
2.7843 g H_{2}O & \div 1mol H_{2}O \div 2mol H = 1.019 g H = 0.312 g H
\end{align*}
\]

C : 1.859 g C \times 100 = 54.97% C
3.4 g C_{x}H_{y}O_{z}

H : 0.312 g H \times 100 = 9.27% H
3.4 g C_{x}H_{y}O_{z}

O : 100% - 54.97% - 9.27% = 36.47% O

\[
\begin{align*}
54.4g C & \div 12 g C = 4.53 \text{ mol } C = 2 \text{ mol } C \\
9.2g H & \div 1 g H = 9.11 \text{ mol } H = 4 \text{ mol } H
\end{align*}
\]

Empirical

\[
C_{2}H_{4}O
\]

\[
36.4g O \div 16g O = 2.28 \text{ mol } O = 1 \text{ mol } O
\]

\[ C_{2}H_{4}O = 44.06g/\text{mol} \quad \text{molecules} \quad \text{Empirical formulas are the same.} \]
AP Chemistry

Percent Composition and Molecular Formula Worksheet

1) What's the empirical formula of a molecule containing 65.5% carbon, 6.5% hydrogen, and 29.0% oxygen?

\[ \frac{65.5\text{ g C}}{12.01\text{ g C}} = 5.45 \text{ mol C} \]

\[ \frac{5.5 \text{ g H}}{1.008\text{ g H}} = 5.46 \text{ mol H} \]

\[ \frac{2.9 \text{ g O}}{15.999\text{ g O}} = 1.81 \text{ mol O} \]

\[ \text{C}_3\text{H}_7\text{O}_3 \]

2) If the molar mass of the compound in problem 1 is 110 grams/mole, what's the molecular formula?

\[ \frac{110 \text{ g mol}}{55.07\text{ g mol C}_6\text{H}_8\text{O}_6} = 2 \text{ (C}_3\text{H}_7\text{O}_3) \Rightarrow \text{C}_6\text{H}_8\text{O}_6 \]

3) What's the empirical formula of a molecule containing 18.7% lithium, 16.3% carbon, and 65.0% oxygen?

\[ \frac{18.7\text{ g Li}}{7.01\text{ g Li}} = 2.66 \text{ mol Li} \]

\[ \frac{16.3\text{ g C}}{12.01\text{ g C}} = 1.36 \text{ mol C} \]

\[ \frac{65.0\text{ g O}}{15.999\text{ g O}} = 3.49 \text{ mol O} \]

\[ \text{Li}_2\text{CO}_3 \]

4) If the molar mass of the compound in problem 3 is 73.8 grams/mole, what's the molecular formula?

\[ \text{Li}_2\text{CO}_3 = 73.8 \text{ g/mol} \]

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