AP Chemistry

Key

Limiting Reactants and Percent Yield

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

   $2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_3$

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

   $\frac{15.2\,g\,Al}{} \cdot \frac{1\,mol\,Al}{26.98\,g\,Al} \cdot \frac{2\,mol\,AlCl_3}{2\,mol\,Al} \cdot \frac{133.33\,g\,AlCl_3}{1\,mol\,AlCl_3} = 75.12\,g\,AlCl_3$

   $\frac{39.1\,g\,Cl_2}{} \cdot \frac{1\,mol\,Cl_2}{70.90\,g} \cdot \frac{2\,mol\,AlCl_3}{3\,mol\,Cl_2} \cdot \frac{133.33\,g\,AlCl_3}{1\,mol\,AlCl_3} = 49.02\,g\,AlCl_3$

   ☆ Limiting Reactant is Cl_2

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

   $\boxed{49.02\,g\,AlCl_3}$

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

   $\frac{49.02\,g\,AlCl_3}{} \cdot \frac{1\,mol\,AlCl_3}{133.33\,g\,AlCl_3} \cdot \frac{2\,mol\,Al}{2\,mol\,AlCl_3} \cdot \frac{26.98\,g\,Al}{1\,mol\,Al} = 9.92\,g\,Al$ was used

   15.2g (started) − 9.92g (used) = 5.28 g Al left over

   e. If you made 50.0g of $AlCl_3$ in the lab, what is your percent yield?

   % yield = $\frac{\text{actual}}{\text{theoret.}} \times 100 = \frac{50.0\,g}{49.02\,g} \times 100 \approx \boxed{102\%}$

   f. If the percent yield were 83.65%, how much could be expected to be made in the experiment?

   $83.65\% = \frac{exp.}{49.02\,g} \times 100$ $\boxed{exp = 41.0\,g}$

2. $3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$
   a. When 13.54 g of O_2 is mixed with 12.21 g of Fe, which is the limiting reactant?

   $\frac{13.54\,g\,O_2}{} \cdot \frac{1\,mol\,O_2}{32.00\,g\,O_2} \cdot \frac{1\,mol\,Fe_3O_4}{2\,mol\,O_2} \cdot \frac{231.55\,g\,Fe_3O_4}{1\,mol\,Fe_3O_4} = 48.99\,g\,Fe_3O_4$

   $\frac{12.21\,g\,Fe}{} \cdot \frac{1\,mol\,Fe}{55.85\,g\,Fe} \cdot \frac{1\,mol\,Fe_3O_4}{3\,mol\,Fe} \cdot \frac{231.55\,g\,Fe_3O_4}{1\,mol\,Fe_3O_4} = 16.87\,g\,Fe_3O_4$

   Fe is limiting

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

   $\boxed{16.9\,g}$

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

   $\frac{16.87\,g\,Fe_3O_4}{} \cdot \frac{1\,mol\,Fe_3O_4}{231.55\,g\,Fe_3O_4} \cdot \frac{2\,mol\,O_2}{1\,mol\,Fe_3O_4} \cdot \frac{32\,g\,O_2}{1\,mol\,O_2} = 4.66\,g$ used

   13.54 g (start) − 4.66 g (used) = 8.88 g Left-over

   d. Kelly performed completed this reaction in a lab and made15.88 g of Fe_3O_4, what was her percent yield?

   $\frac{15.88\,g}{16.9\,g} \times 100 = 94.0\%$ yield

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3. Diiodine pentoxide is useful in devices such as respirators because it reacts with the dangerous gas carbon monoxide, CO, to produce relatively harmless CO_2 according to the following equation:

$I_2O_5 + 5CO \rightarrow I_2 + 5CO_2$

a. In testing a respirator, 2.00g of carbon monoxide gas is passed through diiodine pentoxide. Upon analyzing the results, it is found that 3.17g of I_2 was produced. Calculate the percent yield of the reaction.

$$\frac{2.00\text{ g CO}}{} \cdot \frac{1\text{ mol CO}}{28\text{ g CO}} \cdot \frac{1\text{ mol } I_2}{5\text{ mol CO}} \cdot \frac{253.8\text{ g } I_2}{1\text{ mol } I_2} = 3.63\text{ g } I_2$$

$$\% = \frac{\text{act}}{\text{theor.}} \times 100 = \frac{3.17\text{ g}}{3.63\text{ g}} \times 100 = \boxed{87.3\%}$$

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.

$$\frac{3.17\text{ g } I_2}{} \cdot \frac{1\text{ mol } I_2}{253.8\text{ g } I_2} \cdot \frac{5\text{ mol CO}}{1\text{ mol } I_2} \cdot \frac{28\text{ g CO}}{1\text{ mol CO}} = 1.75\text{ g CO Reacted}$$

$$\underset{\text{(starting)}}{2.00\text{ g CO}} - \underset{\text{(reacted)}}{1.75\text{ g CO}} = \boxed{0.25\text{ g CO unreacted / passed thru}}$$

4. $Cu(s) + Cl_2(g) \rightarrow CuCl_2(s)$

a. If 12.5 g Cu reacts with excess chlorine, calculate the theoretical yield of $CuCl_2$.

$$\frac{12.5\text{ g Cu}}{} \cdot \frac{1\text{ mol Cu}}{63.55\text{ g}} \cdot \frac{1\text{ mol } CuCl_2}{1\text{ mol Cu}} \cdot \frac{134.45\text{ g } CuCl_2}{1\text{ mol } CuCl_2} = \boxed{26.4\text{ g}}$$

b. If ony 25.4g of $CuCl_2$ was produced, what is the percent yield?

$$\frac{25.4\text{ g}}{26.4\text{ g}} \times 100 = \boxed{96.2\%}$$

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.

$2\,NaOH + Cl2 \rightarrow NaCl + NaClO + H_2O$

a. What is the percent yield of the reaction if 1.2kg of Cl_2 reacts to form .90kg of NaClO?

$$\frac{1.2\text{ kg } Cl_2}{} \cdot \frac{1000\text{ g}}{1\text{ kg}} \cdot \frac{1\text{ mol } Cl_2}{70.9\text{ g } Cl_2} \cdot \frac{1\text{ mol NaClO}}{1\text{ mol } Cl_2} \cdot \frac{74.44\text{ g NaClO}}{1\text{ mol NaClO}} = 1259.9\text{ g NaClO}$$

$$\frac{900\text{ g}}{1259.9\text{ g}} \times 100 = \boxed{71.4\%}$$

6. The percent yield for the reaction $PCl_3 + Cl_2 \rightarrow PCl_5$ is 83.2%. What mass of PCl_5 is expected from the reaction of 73.7 g of PCl_3 with excess chlorine?

$$\frac{73.7\text{ g } PCl_3}{} \cdot \frac{1\text{ mol } PCl_3}{137.32\text{ g } PCl_3} \cdot \frac{1\text{ mol } PCl_5}{1\text{ mol } PCl_3} \cdot \frac{208.22\text{ g } PCl_5}{1\text{ mol } PCl_5} = 111.75\text{ g } PCl_5 \cdot 0.832 = \boxed{93.0\text{ g } PCl_5}$$

AP Chemistry

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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 $CO_2(g)$ 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.

$$\underset{(1.2543\,g)}{C_xH_y} + \underset{(excess)}{O_2} \longrightarrow \underset{(3.671\,g)}{CO_2} + H_2O$$

b) How many grams of carbon are in 3.671 g of CO_2? Where did all of this carbon come from? How much carbon was in the original sample of the unknown compound?

$$\frac{3.671\,g\,CO_2}{} \left| \frac{1\,mol\,CO_2}{44\,g\,CO_2} \right| \frac{1\,mol\,C}{1\,mol\,CO_2} \left| \frac{12\,g\,C}{1\,mol\,C} \right. = \boxed{1.00\,g\,C}$$ you could also use percent comp.

All the carbon came from the hydrocarbon

There was 1.00 g C in the original unknown compound.

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

$$C: \frac{1.00\,g\,C}{1.2543\,g\,C_xH_y} \times 100 = \boxed{79.7\%\ \text{Carbon}}$$

$$H: 100\% - 79.7\% = \boxed{20.3\%\ \text{Hydrogen}}$$

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

$$\frac{79.7\,g\,C}{} \left| \frac{1\,mol\,C}{12.01\,g\,C} \right. = \frac{6.64\,mol}{6.64\,mol} = 1\,mol\,C \qquad \boxed{CH_3}$$

$$\frac{20.3\,g\,H}{} \left| \frac{1\,mol\,H}{1.01\,g\,H} \right. = \frac{20.1\,mol}{6.64\,mol} = 3\,mol\,H$$

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

Empirical weight = 15.04 g

$$\frac{30.08\,g/mol}{15.04\,g/mol} = 2\ (CH_3) = \boxed{C_2H_6}$$

ONE
A compound containing carbon and hydrogen is analyzed. Combustion of a 16.81 g sample of this compound produces 38.91 g of $CO_2(g)$. 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 6.79 g of $CO_2(g)$ 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?

Key

Conservation of Mass worksheet continued...

① $C_xH_y + O_2 \longrightarrow CO_2 + H_2O$

C_xH_y: 16.81g CO_2: 38.9g

$$\frac{38.9g\ CO_2}{} \cdot \frac{1mol\ CO_2}{44g\ CO_2} \cdot \frac{1mol\ C}{1mol\ CO_2} \cdot \frac{12g\ C}{1mol\ C} = 10.62g\ C$$

$$C: \frac{10.62g\ C}{16.81g\ C_xH_y} \times 100 = 63.2\%\ C$$

$H: 100\% - 63.2\% = 36.8\%\ H$

$$\frac{63.2g\ C}{} \cdot \frac{1mol\ C}{12g\ C} = \frac{5.27mol\ C}{5.27mol} = 1mol\ C$$

$$\frac{36.8gH}{} \cdot \frac{1molH}{1.01gH} = \frac{36.43mol\ H}{5.27mol} = 7mol\ H$$

$\boxed{\text{Empirical Form } CH_7}$

$CH_7 = 19.08g$ $\frac{38.03g/mol}{19.08g/mol} = 2(CH_7) = \boxed{C_2H_{14}}$

② $C_xH_y + O_2 \longrightarrow CO_2 + H_2O$

C_xH_y: 0.213g H_2O: 0.2132g

$$\frac{0.2132g\ H_2O}{} \cdot \frac{1mol\ H_2O}{18g\ H_2O} \cdot \frac{2mol\ H}{1mol\ H_2O} \cdot \frac{1.01g\ H}{1mol\ H} = 0.0239g\ H$$

$$\frac{0.0239g\ H}{0.213g\ C_xH_y} \times 100 = 11.2\%\ H \quad \frac{11.2g\ H}{} \cdot \frac{1mol\ H}{1.01gH} = \frac{11.1molH}{7.39mol} = 1.5$$

$$100\% - 11.2\% = 88.8\%\ C \quad \frac{88.8g\ C}{} \cdot \frac{1mol\ C}{12.01g\ C} = \frac{7.39mol\ C}{7.39mol} = 1$$

$H = 1.5\ mol \times 2 = 3mol\ H$

$C = 1\ mol \times 2 = 2\ mol\ C$

$\boxed{C_2H_3}$ ← empirical

$C_2H_3 = 27.05g$ $\frac{81.15g}{27.05g} = 3(C_2H_3)$

$\boxed{\text{molecular Formula: } C_6H_9}$

(3) $C_xH_yO_z + O_2 \longrightarrow CO_2 + H_2O$

3.4g 6.79g 2.7843g

$\frac{6.79g\ CO_2}{} \left| \frac{1mol\ CO_2}{44g\ CO_2} \right| \frac{1mol\ C}{1mol\ CO_2} \left| \frac{12g\ C}{1mol\ C} \right. = 1.85g\ C$

$\frac{2.7843g\ H_2O}{} \left| \frac{1mol\ H_2O}{18g\ H_2O} \right| \frac{2mol\ H}{1mol\ H_2O} \left| \frac{1.01g\ H}{1mol\ H} \right. = 0.312g\ H$

$C: \frac{1.85g\ C}{3.4g\ C_xH_yO_z} \times 100 = 54.4\%\ C$

$H: \frac{0.312\ gH}{3.4g\ C_xH_yO_z} \times 100 = 9.2\%\ H$

$O: 100\% - 54.4\% - 9.2\% = 36.4\%\ O$

$\frac{54.4g\ C}{} \left| \frac{1mol\ C}{12.01g\ C} \right. = \frac{4.53}{2.28} mol\ C = 2\ mol\ C$

$\frac{9.2g\ H}{} \left| \frac{1mol\ H}{1.01g\ H} \right. = \frac{9.11}{2.28} mol\ H = 4\ mol\ H$

$\frac{36.4g\ O}{} \left| \frac{1mol\ O}{16g\ O} \right. = \frac{2.28}{2.28} mol\ O = 1mol\ O$

Empirical
C_2H_4O

$C_2H_4O = 44.06g/mol$ ⇐ molecular + 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, 5.5% hydrogen, and 29.0% oxygen?

$\frac{65.5\,g\,C}{} \left| \frac{1\,mol\,C}{12.01\,g\,C} \right. = \frac{5.45\,mol\,C}{1.81\,mol} = 3\,mol$

$\frac{5.5\,g\,H}{} \left| \frac{1\,mol\,H}{1.01\,g\,H} \right. = \frac{5.45\,mol\,H}{1.81\,mol} = 3\,mol$

$\frac{29.0\,g\,O}{} \left| \frac{1\,mol\,O}{16\,g\,O} \right. = \frac{1.81\,mol\,O}{1.81\,mol} = 1\,mol\,O$

$\boxed{C_3H_3O}$

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

$C_3H_3O = 55.06\,g$

$\frac{110\,g/mol}{55.06\,g/mol} = 2\,(C_3H_3O) \Rightarrow \boxed{C_6H_6O_2}$

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

$\frac{18.7\,g\,Li}{} \left| \frac{1\,mol\,Li}{6.94\,g\,Li} \right. = \frac{2.69\,mol\,Li}{1.36\,mol} = 2\,mol\,Li$

$\frac{16.3\,g\,C}{} \left| \frac{1\,mol\,C}{12.01\,g} \right. = \frac{1.36\,mol\,C}{1.36\,mol} = 1\,mol\,C$

$\frac{65.0\,g\,O}{} \left| \frac{1\,mol\,O}{16\,g\,O} \right. = \frac{4.06\,mol\,O}{1.36\,mol} = 3\,mol\,O$

$\boxed{Li_2CO_3}$

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

$Li_2CO_3 = 73.89$

molecular and Empirical are the same

Li_2CO_3