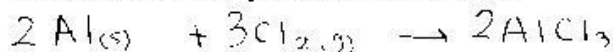


Limiting Reactants and Percent Yield

1. Aluminum metal reacts with chlorine gas in a synthesis reaction.

a. Write the balanced equation for this reaction.



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

$$\frac{15.2\text{ g Al}}{26.98\text{ g Al}} \times \frac{1\text{ mol Al}}{2\text{ mol Al}} \times \frac{2\text{ mol AlCl}_3}{3\text{ mol Cl}_2} \times \frac{133.33\text{ g AlCl}_3}{1\text{ mol AlCl}_3} = 75.12\text{ g AlCl}_3$$

$$\frac{39.1\text{ g Cl}_2}{70.90\text{ g}} \times \frac{1\text{ mol Cl}_2}{3\text{ mol Cl}_2} \times \frac{2\text{ mol AlCl}_3}{2\text{ mol Al}} \times \frac{133.33\text{ g AlCl}_3}{1\text{ mol AlCl}_3} = 49.02\text{ g AlCl}_3$$

* Limiting Reactant is Cl₂ *

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

$$\boxed{49.02\text{ g AlCl}_3}$$

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

$$\frac{49.02\text{ g AlCl}_3}{133.33\text{ g AlCl}_3} \times \frac{1\text{ mol AlCl}_3}{2\text{ mol AlCl}_3} \times \frac{2\text{ mol Al}}{1\text{ mol Al}} \times \frac{26.98\text{ g Al}}{1\text{ mol Al}} = 9.92\text{ g Al used}$$

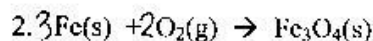
$$\frac{15.2\text{ g}}{\text{(started)}} - \frac{9.92\text{ g}}{\text{(used)}} = \boxed{5.28\text{ g Al left over}}$$

e. If you made 50.0 g of AlCl₃ in the lab, what is your percent yield?

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

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

$$83.65\% = \frac{\text{Exp.}}{49.02\text{ g}} \times 100 \quad \boxed{\text{Exp} = 41.0\text{ g}}$$



a. When 13.54 g of O₂ is mixed with 12.21 g of Fe, which is the limiting reactant?

$$\frac{13.54\text{ g O}_2}{32.00\text{ g O}_2} \times \frac{1\text{ mol O}_2}{2\text{ mol O}_2} \times \frac{1\text{ mol Fe}_3\text{O}_4}{3\text{ mol Fe}} \times \frac{231.55\text{ g Fe}_3\text{O}_4}{1\text{ mol Fe}_3\text{O}_4} = 48.99\text{ g Fe}_3\text{O}_4$$

$$\frac{12.21\text{ g Fe}}{55.85\text{ g Fe}} \times \frac{1\text{ mol Fe}}{3\text{ mol Fe}} \times \frac{1\text{ mol Fe}_3\text{O}_4}{2\text{ mol Fe}} \times \frac{231.55\text{ g Fe}_3\text{O}_4}{1\text{ mol Fe}_3\text{O}_4} = 16.87\text{ g Fe}_3\text{O}_4$$

Fe is limiting

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

$$\boxed{16.9\text{ g}}$$

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

$$\frac{16.87\text{ g Fe}_3\text{O}_4}{231.55\text{ g Fe}_3\text{O}_4} \times \frac{1\text{ mol Fe}_3\text{O}_4}{2\text{ mol O}_2} \times \frac{2\text{ mol O}_2}{1\text{ mol Fe}_3\text{O}_4} \times \frac{32.00\text{ g O}_2}{1\text{ mol O}_2} = 4.66\text{ g used}$$

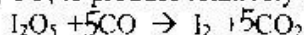
$$\frac{13.54\text{ g (start)}}{4.66\text{ g (used)}} = \boxed{8.88\text{ g left over}}$$

d. Kelly performed completed this reaction in a lab and made 15.88 g of Fe₃O₄, what was her percent yield?

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

key

3. Diiodine 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:



- 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 I₂ was produced. Calculate the percent yield of the reaction.

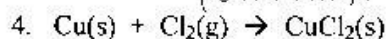
$$\frac{2.00\text{g CO}}{28.0\text{g CO}} \times \frac{1\text{mol CO}}{1\text{mol CO}} \times \frac{1\text{mol I}_2}{5\text{mol CO}} \times \frac{253.8\text{g I}_2}{1\text{mol I}_2} = 3.63\text{g I}_2$$

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

- 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.

$$\frac{3.17\text{g I}_2}{253.8\text{g I}_2} \times \frac{1\text{mol I}_2}{1\text{mol I}_2} \times \frac{5\text{mol CO}}{1\text{mol I}_2} \times \frac{28.0\text{g CO}}{1\text{mol CO}} = 1.75\text{g CO Reacted}$$

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



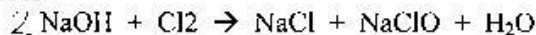
- a. If 12.5 g Cu reacts with excess chlorine, calculate the theoretical yield of CuCl₂.

$$\frac{12.5\text{g Cu}}{63.55\text{g}} \times \frac{1\text{mol Cu}}{1\text{mol Cu}} \times \frac{1\text{mol CuCl}_2}{1\text{mol Cu}} \times \frac{134.45\text{g CuCl}_2}{1\text{mol CuCl}_2} = \boxed{26.4\text{g}}$$

- b. If only 25.4g of CuCl₂ 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.



- a. What is the percent yield of the reaction if 1.2kg of Cl₂ reacts to form .90kg of NaClO?

$$\frac{1.2\text{kg Cl}_2}{1\text{kg}} \times \frac{1000\text{g}}{70.9\text{g Cl}_2} \times \frac{1\text{mol Cl}_2}{1\text{mol Cl}_2} \times \frac{1\text{mol NaClO}}{1\text{mol Cl}_2} \times \frac{74.44\text{g NaClO}}{1\text{mol NaClO}} = 1259.9\text{g NaClO}$$

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

6. The percent yield for the reaction $\text{PCl}_3 + \text{Cl}_2 \rightarrow \text{PCl}_5$ is 83.2%. What mass of PCl₅ is expected from the reaction of 73.7 g of PCl₃ with excess chlorine?

$$\frac{73.7\text{g PCl}_3}{137.32\text{g PCl}_3} \times \frac{1\text{mol PCl}_3}{1\text{mol PCl}_3} \times \frac{1\text{mol PCl}_5}{1\text{mol PCl}_3} \times \frac{208.22\text{g PCl}_5}{1\text{mol PCl}_5} = 111.75\text{g PCl}_5 \times 0.832 = \boxed{93.0\text{g PCl}_5}$$

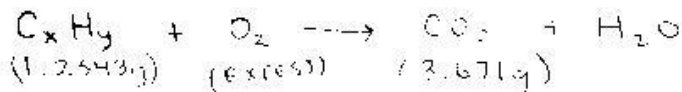
Conservation of Mass Worksheet

key

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 $\text{CO}_2(\text{g})$ 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.



- 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?

$$\frac{3.671\text{g CO}_2}{144\text{g CO}_2} \times \frac{1\text{mol CO}_2}{1\text{mol CO}_2} \times \frac{1\text{mol C}}{1\text{mol CO}_2} \times \frac{12\text{g C}}{1\text{mol C}} = 1.00\text{g C}$$

you could also use percent comp.

All the carbon came from the hydrocarbon

There was 1.00g C in the original unknown compound.

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

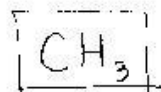
$$\text{C} = \frac{1.00\text{g C}}{1.2543\text{g C}_x\text{H}_y} \times 100 = 79.7\% \text{ Carbon!}$$

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

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

$$\frac{79.7\text{g C}}{12\text{g C}} \times \frac{1\text{mol C}}{1\text{mol C}} = 6.64\text{mol C}$$

$$\frac{20.3\text{g H}}{1\text{g H}} \times \frac{1\text{mol H}}{1\text{mol H}} = 20.3\text{mol H}$$



- 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?

$$\text{Empirical weight} = 15.04\text{g}$$

$$\frac{30.08\text{g/mol}}{15.04\text{g/mol}} \times 2 (\text{H}) = \boxed{\text{C}_2\text{H}_6}$$

ONE

A compound containing carbon and hydrogen is analyzed. Combustion of a 16.81 g sample of this compound produces 38.91 g of $\text{CO}_2(\text{g})$. 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

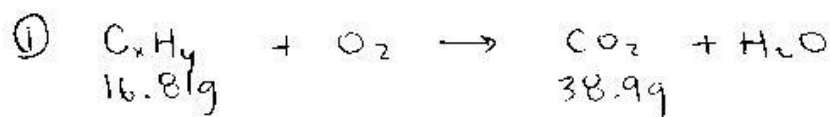
A compound containing carbon and hydrogen is analyzed. Combustion of a 0.213 g sample of this compound yields 0.2132 g of water. What is the empirical formula of this compound? What is the molecular formula if the molar mass of the compound is 81.15 g/mol?

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 $\text{CO}_2(\text{g})$ 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...



$$\frac{38.9\text{g CO}_2}{44\text{g CO}_2} \times \frac{1\text{mol CO}_2}{1\text{mol CO}_2} \times \frac{1\text{mol C}}{1\text{mol C}} \times \frac{12\text{g C}}{1\text{mol C}} = 10.62\text{g C}$$

$$\text{C} = \frac{10.62\text{g C}}{16.81\text{g C}_x\text{H}_y} \times 100 = 63.2\% \text{ C}$$

$$\text{H} = 100\% - 63.2\% = 36.8\% \text{ H}$$

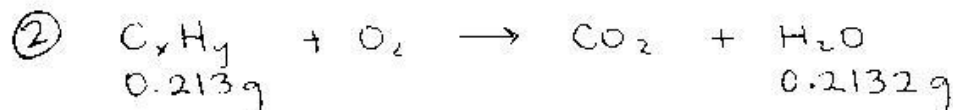
$$\frac{63.2\text{g C}}{12\text{g C}} \times \frac{1\text{mol C}}{1\text{mol C}} = \frac{5.27\text{mol C}}{5.27\text{mol}} = 1\text{mol C}$$

$$\frac{36.8\text{g H}}{1.01\text{g H}} \times \frac{1\text{mol H}}{1\text{mol H}} = \frac{36.43\text{mol H}}{5.27\text{mol}} = 7\text{mol H}$$

Empirical Form
 CH_7

$$\text{CH}_7 = 19.08\text{g}$$

$$\frac{38.03\text{g/mol}}{19.08\text{g/mol}} = 2 (\text{CH}_7) = \boxed{\text{C}_2\text{H}_{14}}$$



$$\frac{0.2132\text{g H}_2\text{O}}{18\text{g H}_2\text{O}} \times \frac{1\text{mol H}_2\text{O}}{1\text{mol H}_2\text{O}} \times \frac{2\text{mol H}}{1\text{mol H}_2\text{O}} \times \frac{1.01\text{g H}}{1\text{mol H}} = 0.0239\text{g H}$$

$$\frac{0.0239\text{g H}}{0.213\text{g C}_x\text{H}_y} \times 100 = 11.2\% \text{ H} \quad \frac{11.2\text{g H}}{1.01\text{g H}} \times \frac{1\text{mol H}}{1\text{mol H}} = \frac{11.1\text{mol H}}{7.39\text{mol}} = 1.5$$

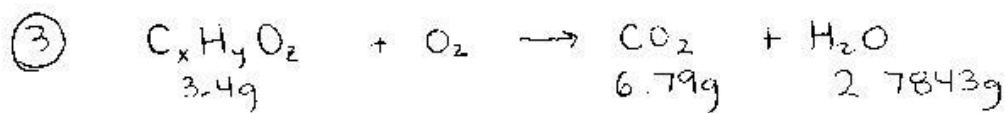
$$100\% - 11.2\% = 88.8\% \text{ C} \quad \frac{88.8\text{g C}}{12.01\text{g C}} \times \frac{1\text{mol C}}{1\text{mol C}} = \frac{7.39\text{mol C}}{7.39\text{mol}} = 1$$

$$\text{H} = 1.5\text{mol} \times 2 = 3\text{mol H}$$
$$\text{C} = 1\text{mol} \times 2 = 2\text{mol C}$$

$\boxed{\text{C}_2\text{H}_3}$ empirical

$$\text{C}_2\text{H}_3 = 27.05\text{g} \quad \frac{81.15\text{g}}{27.05\text{g}} = 3 (\text{C}_2\text{H}_3)$$

molecular Formula
 $\boxed{\text{C}_6\text{H}_9}$



$$\frac{6.79\text{g CO}_2}{44\text{g CO}_2} \times \frac{1\text{mol CO}_2}{1\text{mol CO}_2} \times \frac{1\text{mol C}}{1\text{mol CO}_2} \times \frac{12\text{g C}}{1\text{mol C}} = 1.85\text{g C}$$

$$\frac{2.7843\text{g H}_2\text{O}}{18\text{g H}_2\text{O}} \times \frac{1\text{mol H}_2\text{O}}{1\text{mol H}_2\text{O}} \times \frac{2\text{mol H}}{1\text{mol H}_2\text{O}} \times \frac{1.01\text{g H}}{1\text{mol H}} = 0.312\text{g H}$$

$$\text{C} : \frac{1.85\text{g C}}{3.4\text{g C}_x\text{H}_y\text{O}_z} \times 100 = 54.4\% \text{ C}$$

$$\text{H} : \frac{0.312\text{g H}}{3.4\text{g C}_x\text{H}_y\text{O}_z} \times 100 = 9.2\% \text{ H}$$

$$\text{O} : 100\% - 54.4\% - 9.2\% = 36.4\% \text{ O}$$

$$\frac{54.4\text{g C}}{12.01\text{g C}} \times \frac{1\text{mol C}}{1\text{mol C}} = \frac{4.53\text{mol C}}{2.28} = 2\text{mol C}$$

Empirical
 $\text{C}_2\text{H}_4\text{O}$

$$\frac{9.2\text{g H}}{1.01\text{g H}} \times \frac{1\text{mol H}}{1\text{mol H}} = \frac{9.11\text{mol H}}{2.28} = 4\text{mol H}$$

$$\frac{36.4\text{g O}}{16\text{g O}} \times \frac{1\text{mol O}}{1\text{mol O}} = \frac{2.28\text{mol O}}{2.28} = 1\text{mol O}$$

$\text{C}_2\text{H}_4\text{O} = 44.06\text{g/mol} \leftarrow$ molecular & Empirical formulas are the same.

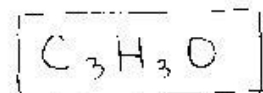
Percent Composition and Molecular Formula Worksheet

- 1) What's the empirical formula of a molecule containing 65.5% carbon, 5.5% hydrogen, and 29.0% oxygen?

$$\frac{65.5 \text{ g C}}{12.01 \text{ g/mol}} = \frac{5.45 \text{ mol C}}{1.81 \text{ mol}} = 3 \text{ mol C}$$

$$\frac{5.5 \text{ g H}}{1.01 \text{ g/mol}} = \frac{5.45 \text{ mol H}}{1.81 \text{ mol}} = 3 \text{ mol H}$$

$$\frac{29.0 \text{ g O}}{16.00 \text{ g/mol}} = \frac{1.81 \text{ mol O}}{1.81 \text{ mol}} = 1 \text{ mol O}$$



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

$$\text{C}_3\text{H}_3\text{O} = 55.06 \text{ g}$$

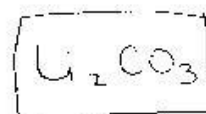
$$\frac{110 \text{ g/mol}}{55.06 \text{ g/mol}} = 2 (\text{C}_3\text{H}_3\text{O}) \Rightarrow \boxed{\text{C}_6\text{H}_6\text{O}_2}$$

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

$$\frac{18.7 \text{ g Li}}{6.94 \text{ g/mol}} = \frac{2.69 \text{ mol Li}}{1.36 \text{ mol}} = 2 \text{ mol Li}$$

$$\frac{16.3 \text{ g C}}{12.01 \text{ g/mol}} = \frac{1.36 \text{ mol C}}{1.36 \text{ mol}} = 1 \text{ mol C}$$

$$\frac{65.0 \text{ g O}}{16.00 \text{ g/mol}} = \frac{4.06 \text{ mol O}}{1.36 \text{ mol}} = 3 \text{ mol O}$$



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

$$\text{Li}_2\text{CO}_3 = 73.89 \text{ g}$$

[Molecular and Empirical
are the same
 Li_2CO_3]