

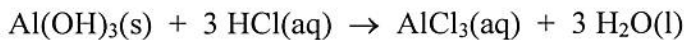
Key

# HC Stoichiometry, Limiting Reactant & % Yield

## SUPPLEMENTAL PRACTICE PROBLEMS

### General Stoichiometry

1. Several brands of antacid tablets use aluminum hydroxide to neutralize excess acid.



[Molar masses:            78.01            36.46            133.4            18.02]

If 0.750 g of  $\text{Al(OH)}_3$  is completely reacted:

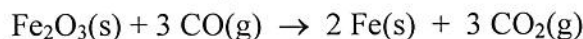
- a) What mass of HCl (aq) is required?

$$0.750 \text{ g Al(OH)}_3 \times \frac{1 \text{ mol Al(OH)}_3}{78.01 \text{ g Al(OH)}_3} \times \underbrace{\frac{3 \text{ mol HCl}}{1 \text{ mol Al(OH)}_3}}_{\text{mole ratio from equation}} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = \boxed{1.05 \text{ g HCl}}$$

- b) What mass of water is produced?

$$0.750 \text{ g Al(OH)}_3 \times \frac{1 \text{ mol Al(OH)}_3}{78.01 \text{ g Al(OH)}_3} \times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol Al(OH)}_3} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{0.520 \text{ g H}_2\text{O}}$$

2. The equation for one of the reactions in the process of reducing iron ore to the metal is



[Molar masses:            159.7            28.01            55.85            44.01]

- a) What is the maximum mass of iron, in grams, that can be obtained from 454 g of iron(III) oxide?

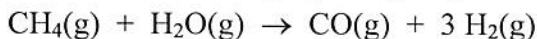
$$454 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = \boxed{318 \text{ g Fe}}$$

- b) What volume of  $\text{CO}_2$  (g) can be produced when 454 g of iron(III) oxide react completely?

$$454 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{3 \text{ mol CO}_2}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = \boxed{191 \text{ L CO}_2}$$

## Limiting Reactants

3. The reaction of methane and water is one way to prepare hydrogen:



[Molar masses: 16.04 18.02 28.01 2.02]

If you begin with 995 g of  $\text{CH}_4$  and 2510 g of water, what is the maximum mass of  $\text{H}_2$  that can be produced?

w/ 995g  $\text{CH}_4$ : g  $\text{CH}_4 \rightarrow$  mol  $\text{CH}_4 \rightarrow$  mol  $\text{H}_2 \rightarrow$  g  $\text{H}_2$

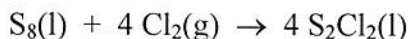
$$995 \text{ g } \text{CH}_4 \times \frac{1 \text{ mol } \text{CH}_4}{16.04 \text{ g } \text{CH}_4} \times \frac{3 \text{ mol } \text{H}_2}{1 \text{ mol } \text{CH}_4} \times \frac{2.02 \text{ g } \text{H}_2}{1 \text{ mol } \text{H}_2} = 376 \text{ g } \text{H}_2$$

Max. mass that can be produced  
 $\text{CH}_4$  is the limiting reactant

w/ 2510g  $\text{H}_2\text{O}$ : g  $\text{H}_2\text{O} \rightarrow$  mol  $\text{H}_2\text{O} \rightarrow$  mol  $\text{H}_2 \rightarrow$  g  $\text{H}_2$

$$2510 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} \times \frac{3 \text{ mol } \text{H}_2}{1 \text{ mol } \text{H}_2\text{O}} \times \frac{2.02 \text{ g } \text{H}_2}{1 \text{ mol } \text{H}_2} = 844 \text{ g } \text{H}_2$$

4. Disulfur dichloride,  $\text{S}_2\text{Cl}_2$ , is used to vulcanize rubber. It can be made by treating molten sulfur with gaseous chlorine:



[Molar masses: 256.6 70.91 135.0]

Starting with a mixture of 32.0 g of sulfur and 71.0 g of  $\text{Cl}_2$ , which is the limiting reactant? What is the maximum mass of  $\text{S}_2\text{Cl}_2$  that can be produced?

w/ 32.0g S: g  $\text{S}_8 \rightarrow$  mol  $\text{S}_8 \rightarrow$  mol  $\text{S}_2\text{Cl}_2 \rightarrow$  g  $\text{S}_2\text{Cl}_2$

$$32.0 \text{ g } \text{S}_8 \times \frac{1 \text{ mol } \text{S}_8}{256.6 \text{ g } \text{S}_8} \times \frac{4 \text{ mol } \text{S}_2\text{Cl}_2}{1 \text{ mol } \text{S}_8} \times \frac{135.0 \text{ g } \text{S}_2\text{Cl}_2}{1 \text{ mol } \text{S}_2\text{Cl}_2} = 67.3 \text{ g } \text{S}_2\text{Cl}_2$$

Max. mass of  $\text{S}_2\text{Cl}_2$   
 $\text{S}_8$  is LR.

w/ 71.0 g  $\text{Cl}_2$ : g  $\text{Cl}_2 \rightarrow$  mol  $\text{Cl}_2 \rightarrow$  mol  $\text{S}_2\text{Cl}_2 \rightarrow$  g  $\text{S}_2\text{Cl}_2$

$$71.0 \text{ g } \text{Cl}_2 \times \frac{1 \text{ mol } \text{Cl}_2}{70.91 \text{ g } \text{Cl}_2} \times \frac{4 \text{ mol } \text{S}_2\text{Cl}_2}{4 \text{ mol } \text{Cl}_2} \times \frac{135.0 \text{ g } \text{S}_2\text{Cl}_2}{1 \text{ mol } \text{S}_2\text{Cl}_2} = 135 \text{ g } \text{S}_2\text{Cl}_2$$

## Percent Yield

29. Diborane,  $B_2H_6$ , is a valuable compound in the synthesis of new organic compounds. One of several ways this boron compound can be made is by the reaction



[Molar masses: 37.84 253.8 27.67 149.9 2.02]

Suppose you use 1.203 g of  $NaBH_4$  with an excess of iodine and obtain 0.295 g of  $B_2H_6$ . What is the percent yield of  $B_2H_6$ ?

$$\% \text{ yield} = \frac{\text{actual mass}}{\text{theoretical mass}} \times 100\% = \frac{0.295 \text{ g}}{0.440 \text{ g}} \times 100\% = \boxed{67.0\%}$$

need to find theoretical mass  $B_2H_6$ :

$$1.203 \text{ g } NaBH_4 \times \frac{1 \text{ mol } NaBH_4}{37.84 \text{ g } NaBH_4} \times \frac{1 \text{ mol } B_2H_6}{2 \text{ mol } NaBH_4} \times \frac{27.67 \text{ g } B_2H_6}{1 \text{ mol } B_2H_6} = \boxed{\text{Theoretical: } 0.440 \text{ g}}$$

31. Disulfur dichloride, which has a revolting smell, can be prepared by directly combining  $S_8$  and  $Cl_2$ , but it can also be made by the following reaction:



[Molar masses: 103.0 41.99 108.1 135.0 58.46]

- a) Assume you begin with 5.23 g of  $SCl_2$  and excess  $NaF$ . What is the theoretical yield of  $S_2Cl_2$ ?

$g SCl_2 \rightarrow \text{mol } SCl_2 \rightarrow \text{mol } S_2Cl_2 \rightarrow g S_2Cl_2$

$$5.23 \text{ g } SCl_2 \times \frac{1 \text{ mol } SCl_2}{103.0 \text{ g } SCl_2} \times \frac{1 \text{ mol } S_2Cl_2}{3 \text{ mol } SCl_2} \times \frac{135.0 \text{ g } S_2Cl_2}{1 \text{ mol } S_2Cl_2} = \boxed{2.28 \text{ g } S_2Cl_2}$$

- b) If only 1.19 g of  $S_2Cl_2$  is obtained, what is the percent yield of the compound?

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% = \frac{1.19 \text{ g}}{2.28 \text{ g}} \times 100\% = \boxed{52.2\%}$$

