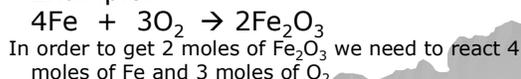


# Stoichiometry

Chapter 9, p. 275 - 294

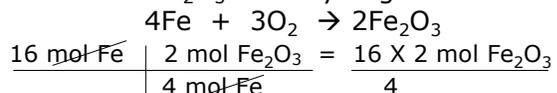
## Intro to Stoichiometry

- **Reaction Stoichiometry:**  
Involves the mass relationships between reactants and products in a chemical reaction
- Coefficients in a chemical reaction represent the mole ratios of each substance that react together
- Example:



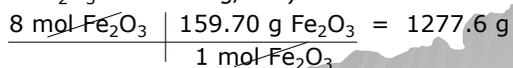
## Example Calculation

If you had 16 moles of Fe, how many moles of  $\text{Fe}_2\text{O}_3$  would you get?



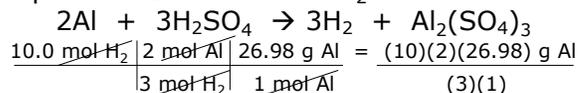
8 mol  $\text{Fe}_2\text{O}_3$

How many grams would that be? (molar mass of  $\text{Fe}_2\text{O}_3$  is 159.70 g/mol)



## Another Example

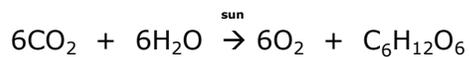
How many grams of Al are needed to produce 10.0 moles of  $\text{H}_2$ ?



179 g Al

## You Try It!

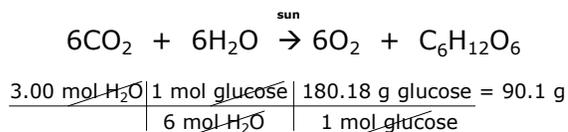
What mass of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) is produced when 3.00 mol of water react with the  $\text{CO}_2$  during photosynthesis?



Hint: go from mol water, to mol of glucose, to grams of glucose

## Solution!

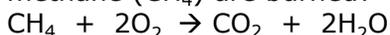
What mass of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) is produced when 3.00 mol of water react with the  $\text{CO}_2$  during photosynthesis?



### Stoichiometry: Mass-to-Mass

- **Conversions from moles to grams (and grams to moles) are necessary to determine how many grams are required or obtained in any chemical reaction**

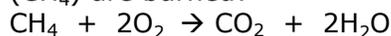
- Example: how many grams of water are produced when 120.0 grams of methane (CH<sub>4</sub>) are burned?



Hint: convert grams methane to moles methane to moles water to grams water

### Stoichiometry: Mass-to-Mass

Example: how many grams of water are produced when 120.0 grams of methane (CH<sub>4</sub>) are burned?

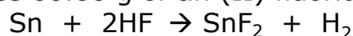


$$120.00 \text{ g CH}_4 \left| \frac{1 \text{ mol CH}_4}{16.05 \text{ g CH}_4} \right| \left| \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \right| \left| \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right| =$$

$$\frac{(120.00)(1)(2)(18.02) \text{ g H}_2\text{O}}{(16.05)(1)(1)} = 269.5 \text{ g H}_2\text{O}$$

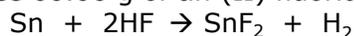
### Stoichiometry: Mass-to-Mass

Example: What mass of HF is required to produce 60.00 g of tin (II) fluoride?



### Stoichiometry: Mass-to-Mass

Example: What mass of HF is required to produce 60.00 g of tin (II) fluoride?



$$60.00 \text{ g SnF}_2 \left| \frac{1 \text{ mol SnF}_2}{156.71 \text{ g SnF}_2} \right| \left| \frac{2 \text{ mol HF}}{1 \text{ mol SnF}_2} \right| \left| \frac{20.01 \text{ g HF}}{1 \text{ mol HF}} \right| =$$

$$\frac{(60.00)(1)(2)(20.01) \text{ g HF}}{(156.71)(1)(1)} = 15.32 \text{ g HF}$$

### Stoichiometry: Limiting Reactants

- **Because chemical reactions occur in fixed ratios, the amount of each product is limited by how much of each reactant is present**
- **When one of the reactants is used up before one or more of the others, we call this reactant the "limiting reactant" or "limiting reagent"**
- **The limiting reactant "limits" the amount of each product that can be obtained in the reaction**
- **The other reactant(s) may be "in excess"**

### Analogy: Brownies

#### Ingredients

½ cup butter  
1 cup white sugar  
2 eggs  
1 teaspoon vanilla extract  
1/3 cup unsweetened cocoa powder  
½ cup all-purpose flour  
¼ teaspoon salt  
¼ teaspoon baking powder

What if you wanted to make brownies but only had 1 egg, but enough of all the other ingredients?

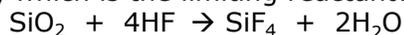
Which ingredient is the limiting ingredient? Excess? How much product would you have if you stick with the ratios in the recipe?

### Tips on Determining the Limiting Reactant

- Calculate how much product is obtained using the amount of one reactant
- Do the same calculation for the other reactant(s)
- Compare the numbers. Which product amount is less?
- The limiting reactant is the one that will provide the least amount of product.

### Example: Limiting Reactant Calculation with Moles

If 2.0 mol of HF are exposed to 4.5 mol of SiO<sub>2</sub>, which is the limiting reactant?



Given: 2.0 mol HF, 4.5 mol SiO<sub>2</sub>

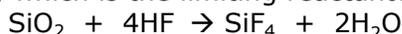
Unknown: limiting reactant

$$\frac{2.0 \text{ mol HF}}{4 \text{ mol HF}} \times \frac{1 \text{ mol SiO}_2}{1 \text{ mol SiO}_2} = 0.50 \text{ mol SiO}_2 \text{ required}$$

Since there are 4.5 mol SiO<sub>2</sub> available and only 0.50 mol is required, we would say that SiO<sub>2</sub> is the excess reactant and HF is the limiting reactant.

### Alternate Calculation for Same Problem

If 2.0 mol of HF are exposed to 4.5 mol of SiO<sub>2</sub>, which is the limiting reactant?



Given: 2.0 mol HF, 4.5 mol SiO<sub>2</sub>

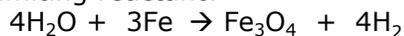
Unknown: limiting reactant

$$\frac{4.5 \text{ mol SiO}_2}{1 \text{ mol SiO}_2} \times \frac{4 \text{ mol HF}}{4 \text{ mol HF}} = 18 \text{ mol HF required}$$

Since there are only 2 mol of HF available but 18 mol of HF is required to react with all the 4.5 mol of SiO<sub>2</sub>, we would still arrive at the conclusion that SiO<sub>2</sub> is the excess reactant and HF is the limiting reactant.

### Example: Limiting Reactant Calculation with Mass

When 36.0 g of H<sub>2</sub>O is mixed with 167 g of Fe to produce magnetite (Fe<sub>3</sub>O<sub>4</sub>), which is the limiting reactant?



Given: 167 g Fe, 36.0 g H<sub>2</sub>O

Unknown: limiting reactant

Convert to moles first!!!

$$\frac{167.0 \text{ g Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}}{1 \text{ mol Fe}} = 2.99 \text{ mol Fe}$$

$$\frac{36.0 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 2.00 \text{ mol H}_2\text{O} = \text{limiting}$$

### Example: Limiting Reactant Calculation with Mass

What mass in grams of magnetite (Fe<sub>3</sub>O<sub>4</sub>) is produced?



Tip: Use the limiting reactant, H<sub>2</sub>O

$$\frac{2.00 \text{ mol H}_2\text{O}}{4 \text{ mol H}_2\text{O}} \times \frac{1 \text{ mol Fe}_3\text{O}_4}{1 \text{ mol Fe}_3\text{O}_4} \times \frac{231.55 \text{ g Fe}_3\text{O}_4}{1 \text{ mol Fe}_3\text{O}_4} = 116 \text{ g Fe}_3\text{O}_4$$

### Example: Limiting Reactant Calculation with Mass

What mass in grams of excess reactant remains when the reaction is completed?



Tip: Use the limiting reactant, H<sub>2</sub>O

$$\frac{2.00 \text{ mol H}_2\text{O}}{4 \text{ mol H}_2\text{O}} \times \frac{3 \text{ mol Fe}}{3 \text{ mol Fe}} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 83.8 \text{ g Fe used in rxn}$$

167 g of Fe was used in the original question. Only 83.8 g Fe is required. Therefore:

$$167 \text{ g Fe present} - 83.8 \text{ g Fe used} = 83 \text{ g Fe left over}$$

## Percent Yield

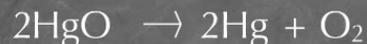
One of the things reaction stoichiometry allows us to do is determine the amount of product formed from a given amount of reactant. As an example, consider the reaction:



Suppose we heat up 1.0 gram of HgO. How many grams of oxygen would we expect to be produced?

## Reaction Stoichiometry

Amount of Product Formed and Percent Yield



$$1.0 \text{ g HgO} \frac{1 \text{ mol HgO}}{216.6 \text{ g HgO}} \frac{1 \text{ mol O}_2 \text{ produced}}{2 \text{ mol HgO consumed}} \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = 0.074 \text{ g O}_2$$

We call the 0.074 g O<sub>2</sub> the theoretical yield, and it is the maximum amount of product that can be formed

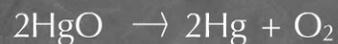
## Percent Yield

In many cases, the amount of product obtained is considerably smaller. For instance, there may be impurities present in the reactants, the reaction may not go to completion, or some of the product may be lost and not captured by the experimenter.

Let's say that we actually captured 0.069 grams of oxygen from 1.0 g of HgO. We call 0.069 g O<sub>2</sub> the actual yield.

## Reaction Stoichiometry

Amount of Product Formed and Percent Yield



Theoretical Yield: 0.074 g O<sub>2</sub>

Actual Yield: 0.069 g O<sub>2</sub>

Percent Yield:  $\frac{0.069 \text{ g O}_2}{0.074 \text{ g O}_2} = 93.2\% \text{ Yield}$

The percent yield is the ratio between the actual and theoretical yield, expressed as a percent. The theoretical yield is the max amount of product we can obtain, and the percent yield tells how much of this was actually obtained.