Stoichiometry

Chapter 9, p. 275 - 294

Intro to Stoichiometry

- **<u>Reaction Stoichiometry</u>**: 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:
  $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

In order to get 2 moles of Fe_2O_3 we need to react 4 moles of Fe and 3 moles of O_2

Example Calculation

If you had 16 moles of Fe, how many moles of Fe_2O_3 would you get?

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$

$$\frac{16\ \cancel{\text{mol Fe}}}{} \cdot \frac{2\ \text{mol } Fe_2O_3}{4\ \cancel{\text{mol Fe}}} = \frac{16 \times 2\ \text{mol } Fe_2O_3}{4}$$

8 mol Fe_2O_3

How many grams would that be? (molar mass of Fe_2O_3 is 159.70 g/mol)

$$\frac{8\ \cancel{\text{mol } Fe_2O_3}}{} \cdot \frac{159.70\ \text{g } Fe_2O_3}{1\ \cancel{\text{mol } Fe_2O_3}} = 1277.6\ \text{g}$$

Another Example

How many grams of Al are needed to produce 10.0 moles of H_2?

$$2Al + 3H_2SO_4 \rightarrow 3H_2 + Al_2(SO_4)_3$$

$$\frac{10.0\ \cancel{\text{mol } H_2}}{} \cdot \frac{2\ \cancel{\text{mol Al}}}{3\ \cancel{\text{mol } H_2}} \cdot \frac{26.98\ \text{g Al}}{1\ \cancel{\text{mol Al}}} = \frac{(10)(2)(26.98)\ \text{g Al}}{(3)(1)}$$

179 g Al

You Try It!

What mass of glucose ($C_6H_{12}O_6$) is produced when 3.00 mol of water react with the CO_2 during photosynthesis?

$$6CO_2 + 6H_2O \xrightarrow{\text{sun}} 6O_2 + C_6H_{12}O_6$$

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

Solution!

What mass of glucose ($C_6H_{12}O_6$) is produced when 3.00 mol of water react with the CO_2 during photosynthesis?

$$6CO_2 + 6H_2O \xrightarrow{\text{sun}} 6O_2 + C_6H_{12}O_6$$

$$\frac{3.00\ \cancel{\text{mol } H_2O}}{} \cdot \frac{1\ \cancel{\text{mol glucose}}}{6\ \cancel{\text{mol } H_2O}} \cdot \frac{180.18\ \text{g glucose}}{1\ \cancel{\text{mol glucose}}} = 90.1\ \text{g}$$

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_4) are burned?

  $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

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_4) are burned?

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$$\frac{120.00\ \cancel{g\ CH_4}}{} \left| \frac{1\ \cancel{mol\ CH_4}}{16.05\ \cancel{g\ CH_4}} \right| \frac{2\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ CH_4}} \left| \frac{18.02\ g\ H_2O}{1\ \cancel{mol\ H_2O}} \right. =$$

$$\frac{(120.00)(1)(2)(18.02)\ g\ H_2O}{(16.05)(1)(1)} = 269.5\ g\ H_2O$$

Stoichiometry: Mass-to-Mass

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

$Sn + 2HF \rightarrow SnF_2 + H_2$

Stoichiometry: Mass-to-Mass

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

$Sn + 2HF \rightarrow SnF_2 + H_2$

$$\frac{60.00\ \cancel{g\ SnF_2}}{} \left| \frac{1\ \cancel{mol\ SnF_2}}{156.71\ \cancel{g\ SnF_2}} \right| \frac{2\ \cancel{mol\ HF}}{1\ \cancel{mol\ SnF_2}} \left| \frac{20.01\ g\ HF}{1\ \cancel{mol\ HF}} \right. =$$

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

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_2, which is the limiting reactant?

$$SiO_2 + 4HF \rightarrow SiF_4 + 2H_2O$$

Given: 2.0 mol HF, 4.5 mol SiO_2

Unknown: limiting reactant

$$\frac{2.0\ \cancel{\text{mol HF}}}{} \left| \frac{1\ \text{mol } SiO_2}{4\ \cancel{\text{mol HF}}} \right. = 0.50\ \text{mol } SiO_2 \text{ required}$$

Since there are 4.5 mol SiO_2 available and only 0.50 mol is required, we would say that SiO_2 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_2, which is the limiting reactant?

$$SiO_2 + 4HF \rightarrow SiF_4 + 2H_2O$$

Given: 2.0 mol HF, 4.5 mol SiO_2

Unknown: limiting reactant

$$\frac{4.5\ \cancel{\text{mol } SiO_2}}{} \left| \frac{4\ \text{mol HF}}{1\ \cancel{\text{mol } SiO_2}} \right. = 18\ \text{mol HF required}$$

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_2, we would still arrive at the conclusion that SiO_2 is the excess reactant and HF is the limiting reactant.

Example: Limiting Reactant Calculation with Mass

When 36.0 g of H_2O is mixed with 167 g of Fe to produce magnetite (Fe_3O_4), which is the limiting reactant?

$$4H_2O + 3Fe \rightarrow Fe_3O_4 + 4H_2$$

Given: 167 g Fe, 36.0 g H_2O

Unknown: limiting reactant

Convert to moles first!!!

$$\frac{167.0\ \cancel{\text{g Fe}}}{} \left| \frac{1\ \text{mol Fe}}{55.85\ \cancel{\text{g Fe}}} \right. = 2.99\ \text{mol Fe}$$

$$\frac{36.0\ \cancel{\text{g } H_2O}}{} \left| \frac{1\ \text{mol } H_2O}{18.02\ \cancel{\text{g } H_2O}} \right. = 2.00\ \text{mol } H_2O = \text{limiting}$$

Example: Limiting Reactant Calculation with Mass

What mass in grams of magnetite (Fe_3O_4) is produced?

$$4H_2O + 3Fe \rightarrow Fe_3O_4 + 4H_2$$

Tip: Use the limiting reactant, H_2O

$$\frac{2.00\ \cancel{\text{mol } H_2O}}{} \left| \frac{1\ \cancel{\text{mol } Fe_3O_4}}{4\ \cancel{\text{mol } H_2O}} \right| \frac{231.55\ \text{g } Fe_3O_4}{1\ \cancel{\text{mol } Fe_3O_4}} = 116\ \text{g } Fe_3O_4$$

Example: Limiting Reactant Calculation with Mass

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

$$4H_2O + 3Fe \rightarrow Fe_3O_4 + 4H_2$$

Tip: Use the limiting reactant, H_2O

$$\frac{2.00\ \cancel{\text{mol } H_2O}}{} \left| \frac{3\ \cancel{\text{mol Fe}}}{4\ \cancel{\text{mol } H_2O}} \right| \frac{55.85\ \text{g Fe}}{1\ \cancel{\text{mol Fe}}} = 83.8\ \text{g Fe used in rxn}$$

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

167 g Fe present – 83.8 g Fe used = 83 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:

$$2\ HgO \rightarrow 2\ Hg + O_2$$

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

$$2HgO \rightarrow 2Hg + O_2$$

$$1.0g\ HgO \rightarrow ?g\ O_2$$

$$1.0\ g\ HgO \frac{1\ mol\ HgO}{216.6\ g\ HgO} \frac{1\ mol\ O_2\ \text{produced}}{2\ mol\ HgO\ \text{consumed}} \frac{32.0\ g\ O_2}{1\ mol\ O_2}$$

$$= 0.074\ g\ O_2$$

We call the 0.074 g O_2 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_2 the actual yield.

Reaction Stoichiometry

Amount of Product Formed and Percent Yield

$$2HgO \rightarrow 2Hg + O_2$$

Theoretical Yield: 0.074 g O_2

Actual Yield: 0.069 g O_2

Percent Yield: $\frac{0.069\ g\ O_2}{0.074\ g\ O_2} = 93.2\%$ 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.