Chapter 8: Quantities in Chemical Reactions

<u>Stoichiometry</u>: the **numerical** relationship between chemical quantities in a balanced chemical equation.

Ex. $\underline{4}\ NH_3 + \underline{5}\ O_2 \rightarrow \underline{4}\ NO + \underline{6}\ H_2O$

The reaction above can mean: 4 molecules of NH_3 reacts with 5 molecules of O_2 to produce 4 molecules of NO and 6 molecules of H_2O.

It also can be interpreted as: **<u>4 moles</u> of NH_3 reacts with <u>5 moles</u> of O_2 to produce <u>4 moles</u> of NO and <u>6 moles</u> of H_2O.**

How many moles of water can be produced from 4 moles of NH_3in the chemical reaction above? (assume excess O_2)

How many moles of water can be produced from **<u>5.6 moles of NH_3</u>** in the chemical reaction above?

To solve this problem we need to use a **<u>mole ratio</u>** derived from the balanced chemical equation (text uses term "equivalence")

Mole ratios relating amounts of NH_3 and H_2O in equation above:

$$\frac{4 \text{ moles } NH_3}{6 \text{ mole } H_2O} \quad \text{or} \quad \frac{6 \text{ moles } H_2O}{4 \text{ moles } NH_3}$$

?? Moles H_2O = 5.6 moles NH_3

Mole to Mole Conversions

$$4\,NH_3 \;+\; 5\,O_2 \;\rightarrow\; 4\,NO \;+\; 6\,H_2O$$

How many moles of O_2 are needed to produce 7.2 moles of NO? (assume there is excess NH_3)

How many moles of ammonia (NH_3) is needed for complete reaction with 0.850 moles of oxygen gas?

Mass-to-Mass Conversions

Grams A	$\leftrightarrow$	**Moles A**	$\leftrightarrow$	**Moles B**	$\leftrightarrow$	**Grams B**
Conversion factor:	Use molar mass of A		Use **mole ratio from the <u>balanced equation</u>**		Use molar mass of B	

Ex. How many grams of H_2O_2 are needed to completely react with 50.0 g of N_2H_4?

(molar mass N_2H_4 = 32.05g/mol; H_2O_2 = 34.02 g/mol)

$$N_2H_{4\,(l)} \;+\; 2\,H_2O_{2(l)} \;\rightarrow\; N_{2(g)} \;+\; 4\,H_2O_{(g)}$$

Mass-to-Mass Conversions

Grams A $\longleftrightarrow$ **Moles A** $\longleftrightarrow$ **Moles B** $\longleftrightarrow$ **Grams B**

Conversion factor:	Use molar mass of A	Use **mole ratio from the <u>balanced equation</u>**	Use molar mass of B

$$N_2H_{4\,(l)} \;+\; 2\,H_2O_{2(l)} \rightarrow N_{2(g)} \;+\; 4\,H_2O_{(g)}$$

How many grams of H_2O will be produced if 50.0 g of N_2H_4 completely react?

How many grams of $MgCl_2$ will be formed if 1.00 g of HCl is consumed?

$$Mg(OH)_{2\,(s)} \;+\; 2\,HCl_{(aq)} \rightarrow MgCl_{2(aq)} \;+\; 2\,H_2O_{(l)}$$

Mass-to-Mass Conversions

How many moles of H_2O are needed to completely react with 5.00g of NO_2?

$$3\ NO_{2\,(g)} \quad + \quad H_2O_{(l)} \rightarrow 2\ HNO_{3(l)} \quad + \quad NO_{(g)}$$

Limiting Reactant

Analogy: There are 3 assembly lines in a bicycle factory. On a single day, the 3 assembly lines produced the following quantities:

Handle bars - 34 Tires - 62 Frames - 37

How many bicycles can be made?

Limiting reactant - the reactant that is USED UP first (limits the amount of product that can be made).

Limiting Reactant Calculations

In LIMITING REACTANT problems the amounts of **TWO REACTANTS** are given. You are asked:

1) Which reactant is used up first? (which is the limiting reactant?)

and/or

1) How much product can be formed?

Suppose you have 500.0 g Pb and 500.0 g H_2S. Which reactant will be used up first?

$$\mathbf{Pb + H_2S \rightarrow PbS + H_2}$$

To solve this problem: Need to do stoichiometric calculations with **BOTH** given starting materials and convert to amount of **ONE chosen product** in both sets of calculations….

$$\textbf{Grams A} \longleftrightarrow \textbf{Moles A} \longleftrightarrow \textbf{Moles B} \longleftrightarrow \textbf{Grams B}$$

Conversion factor:	Use molar mass of A	Use **mole ratio from the balanced equation**	Use molar mass of B

Limiting Reactant Calculations

If 100.0 g Fe_2O_3 is reacted with 45.00 g Al:

$$Fe_2O_{3(s)} \quad + \quad 2\,Al_{(s)} \quad \rightarrow \quad Al_2O_{3(s)} \quad + \quad 2\,Fe_{(l)}$$

(159.70 g/mol) (26.98 g/mol) (101.96 g/mol) (55.85 g/mol)

a) Which reactant is used up first (is the limiting reactant)?
b) How many grams of iron will be formed?

If 1.00 mole of Cr is reacted with 1.00 mole of O_2, how much Cr_2O_3 can be produced?

$$4\ Cr_{(s)} \quad + \quad 3\,O_{2(g)} \quad \rightarrow \quad 2\ Cr_2O_{3(s)}$$

Theoretical Yield & Percent Yield

Theoretical Yield - the CALCULATED amount of a product that can be obtained from given amounts of reactants in a chemical reaction.

- It is the **maximum** amount of product that can be formed under ideal conditions.

Ex. $4\ NH_3 + 3\ O_2 \rightarrow 2\ N_2 + 6\ H_2O$

- Theoretically, starting with 4 moles of NH_3 a maximum of 2 moles N_2 can be produced.

Actual Yield - the amount of product **actually** obtained (The actual yield is always smaller than theoretical yield)

- Starting with 4 moles of NH_3, only 1 mole of N_2 was actually obtained **when the experiment was carried out. (The actual yield will be given - it is NOT calculated)**

$$\textbf{\% Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

What is the percent yield of N_2 formed in the reaction above?

Limiting Reactant/Percent Yield Practice

$4\ NH_3$	+	$3\ O_2$	$\rightarrow$	$2\ N_2$	+	$6\ H_2O$
(17.03 g/mol)		(32.00 g/mol)		(28.02 g /mol)		(18.02 g/mol)

a) If 115.0 g NH_3 is reacted with 200.0 g of O_2, which is the limiting reactant?

b) What is the theoretical yield of N_2 (in grams)?

c) If only 67.0 g of N_2 is obtained, what is the percent yield?