

Chapter 8: Quantities in Chemical Reactions

Stoichiometry: the **numerical** relationship between chemical quantities in a balanced chemical equation.



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 .

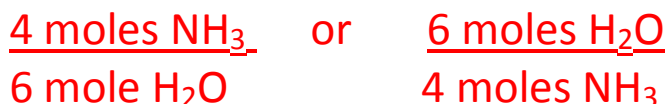
It also can be interpreted as: **4 moles of NH_3 reacts with 5 moles of O_2 to produce 4 moles of NO and 6 moles of H_2O .**

How many moles of water can be produced from 4 moles of NH_3 in the chemical reaction above? (assume excess O_2)

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

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

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



?? Moles $\text{H}_2\text{O} = 5.6$ moles NH_3

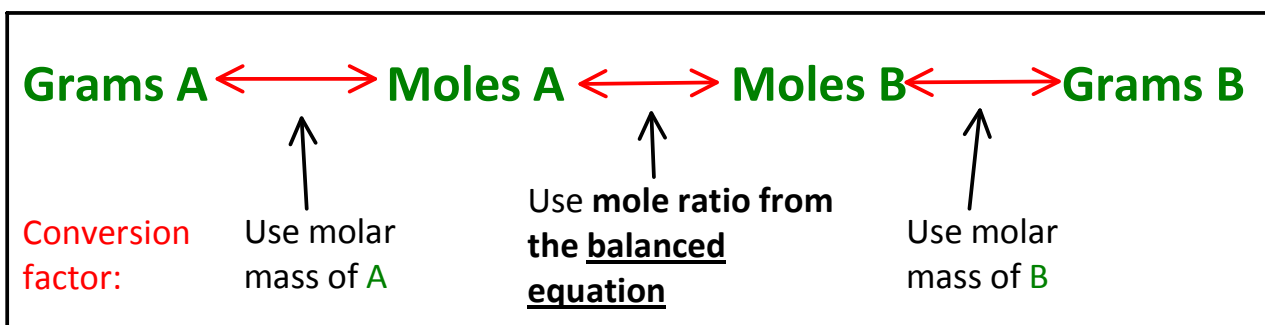
Mole to Mole Conversions



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

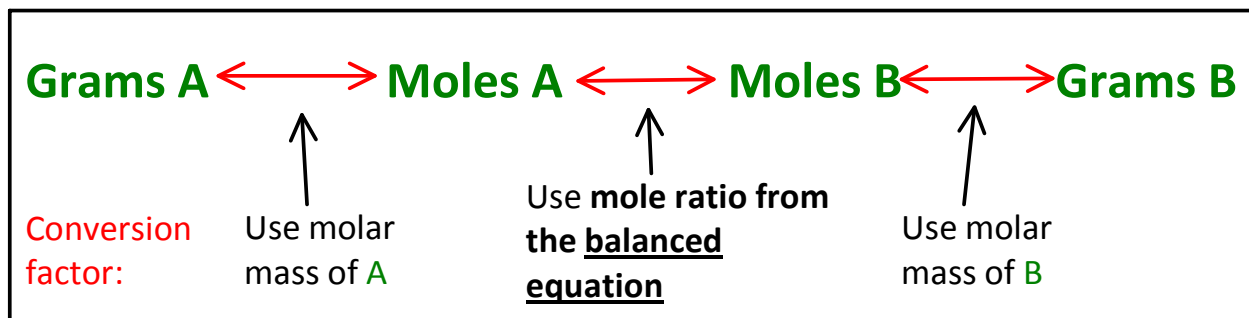


Ex. How many grams of H_2O_2 are needed to completely react with 50.0 g of N_2H_4 ?

(molar mass $\text{N}_2\text{H}_4 = 32.05\text{g/mol}$; $\text{H}_2\text{O}_2 = 34.02\text{ g/mol}$)



Mass-to-Mass Conversions



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?



Mass-to-Mass Conversions

How many moles of H₂O are needed to completely react with 5.00g of NO₂?



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?)

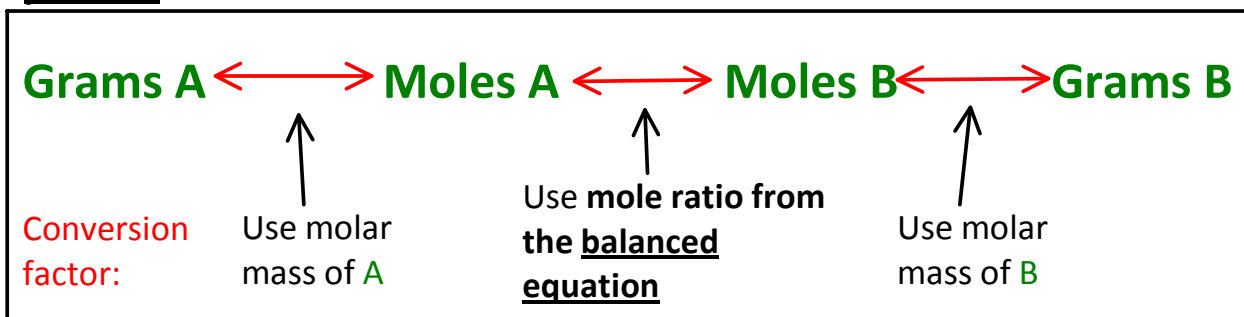
and/or

1) How much product can be formed?

Suppose you have 500.0 g Pb and 500.0 g H₂S. Which reactant will be used up first?

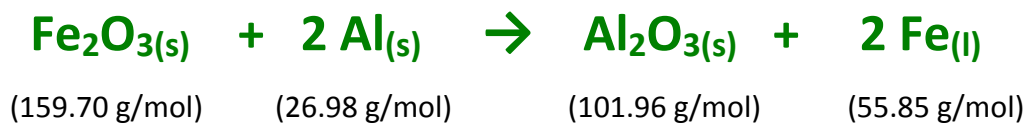


To solve this problem: Need to do stoichiometric calculations with **BOTH** given starting materials and convert to amount of **ONE chosen product** in both sets of calculations....



Limiting Reactant Calculations

If 100.0 g Fe₂O₃ is reacted with 45.00 g Al:



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

If 1.00 mole of Cr is reacted with 1.00 mole of O₂, how much Cr₂O₃ can be produced?



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.



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

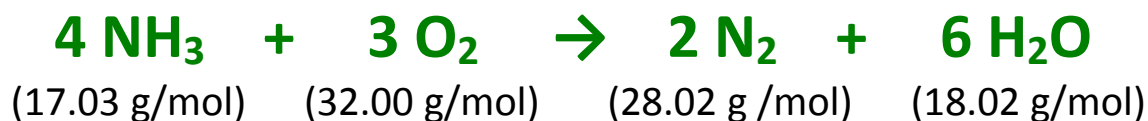
Actual Yield - the amount of product **actually** obtained
(The actual yield is always smaller than theoretical yield)

- 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)

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

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

Limiting Reactant/Percent Yield Practice



a) If 115.0 g NH₃ is reacted with 200.0 g of O₂, which is the limiting reactant?

b) What is the theoretical yield of N₂ (in grams)?

c) If only 67.0 g of N₂ is obtained, what is the percent yield?