Chapter 3
Stoichiometry – 2
Limiting, Excess Reagent and Percent Yields

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Limiting Reagent

- Any problem giving the starting amount for more than one reactant is a limiting reactant problem.
- The limiting reagent is entirely consumed when a reaction goes to completion.
- Once one reactant has been completely consumed, the reaction stops.
- **All amounts produced and reacted are determined by the limiting reactant.**
- How can we determine the limiting reactant?
    1. Use each given amount to calculate the amount of product produced.
    2. The limiting reactant will produce the *lesser* or *least* amount of product.

Magnesium metal is used to prepare zirconium metal, which is used to make the container for nuclear fuel (the nuclear fuel rods):

$$ZrCl_4(g) + 2Mg(s) \rightarrow 2MgCl_2(s) + Zr(s)$$

How many moles of zirconium metal can be produced from a reaction mixture containing 0.20 mol $ZrCl_4$ and 0.50 mol Mg?

$$0.20 \text{ mol } ZrCl_4 \frac{1 \text{ mol Zr}}{1 \text{ mol } ZrCl_4} = 0.20 \text{ mol Zr}$$

$$0.50 \text{ mol Mg} \frac{1 \text{ mol Zr}}{2 \text{ mol Mg}} = 0.25 \text{ mol Zr}$$

Since $ZrCl_4$ gives the lesser amount of Zr, $ZrCl_4$ is the limiting reactant.

0.20 mol Zr will be produced.

Urea, CH_4N_2O, is used as a nitrogen fertilizer. It is manufactured from ammonia and carbon dioxide at high pressure and high temperature:

$$2NH_3 + CO_2(g) \rightarrow CH_4N_2O + H_2O$$

In a laboratory experiment, 10.0 g NH_3 and 10.0 g CO_2 were added to a reaction vessel. What is the maximum quantity (in grams) of urea that can be obtained? How many grams of the excess reactant are left at the end of the reactions?

Molar masses

NH_3	1(14.01) + 3(1.008) = 17.02 g
CO_2	1(12.01) + 2(16.00) = 44.01 g
CH_4N_2O	1(12.01) + 4(1.008) + 2(14.01) + 1(16.00) = 60.06 g

$$10.0\ \text{g NH}_3 \frac{1\ \text{mol NH}_3}{17.024\ \text{g NH}_3} \frac{1\ \text{mol CH}_4\text{N}_2\text{O}}{2\ \text{mol NH}_3} \frac{60.06\ \text{g CH}_4\text{N}_2\text{O}}{1\ \text{mol CH}_4\text{N}_2\text{O}} = 17.6\ \text{g CH}_4\text{N}_2\text{O}$$

$$10.0\ \text{g CO}_2 \frac{1\ \text{mol CO}_2}{44.01\ \text{g CO}_2} \frac{1\ \text{mol CH}_4\text{N}_2\text{O}}{1\ \text{mol CO}_2} \frac{60.06\ \text{g CH}_4\text{N}_2\text{O}}{1\ \text{mol CH}_4\text{N}_2\text{O}} = 13.6\ \text{g CH}_4\text{N}_2\text{O}$$

CO_2 is the limiting reactant since it gives the lesser amount of CH_4N_2O.

13.6 g CH_4N_2O will be produced.

To find the excess NH_3, we need to find how much NH_3 reacted. We use the limiting reactant as our starting point.

$$10.0\text{ g } CO_2 \; \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \; \frac{2\text{ mol } NH_3}{1\text{ mol } CO_2} \; \frac{17.02\text{ g } NH_3}{1\text{ mol } NH_3}$$

$$= 7.734605771\text{ g } NH_3$$

$$= 7.73\text{ g } NH_3 \text{ reacted}$$

Now subtract the amount reacted from the starting amount:

10.0	at start
−7.73	reacted
2.27 g	remains

2.27 g NH_3 is left unreacted

Percent Yield

Theoretical Yield

Theoretical yield is the maximum amount of product that can be obtained by a reaction from given amounts of reactants.

This is a *calculated* amount.

Actual Yield

The amount of product that is actually obtained.

This is a *measured* amount.

Percentage Yield

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$2NH_3 + CO_2(g) \rightarrow CH_4N_2O + H_2O$

When 10.0 g NH_3 and 10.0 g CO_2 are added to a reaction vessel, the limiting reactant is CO_2. The theoretical yield is 13.6 of urea. When this reaction was carried out, 9.3 g of urea was obtained. What is the percent yield?

Theoretical yield = 13.6 g
Actual yield = 9.3 g

$$\text{Percent yield} = \frac{9.3\text{ g}}{13.6\text{ g}} \times 100\%$$

= 68% yield

Review

- Limiting reactant
- Excess Reagent
- % yield