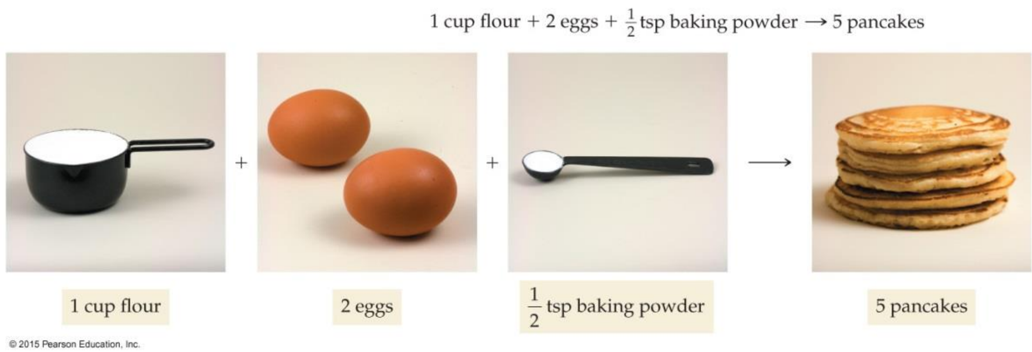
STOICHIOMETRY ANALOGY

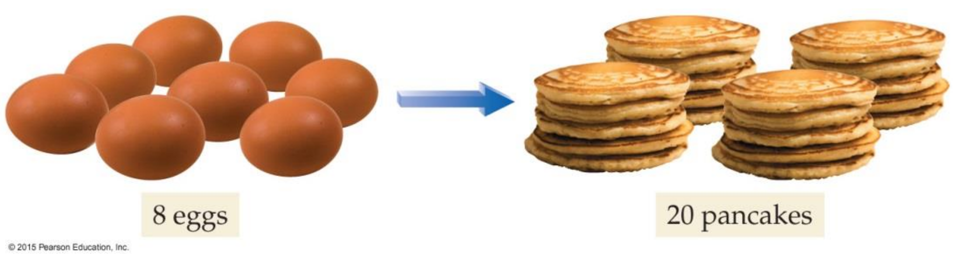
- ***Stoichiometry*** is the quantitative relationship between the ***reactants and products*** in a balanced ***chemical equation***. Stoichiometry allows chemists to predict how much of a reactant is necessary to form a given amount of product or how much of a reactant is required to completely react with another reactant.

- The concept of stoichiometry is analogous to the concept of a recipe such as the one shown below:



- Much like a chemical equation, the recipe above shows the numerical relationship between the ingredients (reactants) and the pancakes (products).

- For example, since 2 eggs are required to make 5 pancakes, it would follow that 8 eggs would be required to make 20 pancakes.



- Much like a chemical equation, the recipe contains numerical relationships between the pancake ingredients and the number of pancakes. Some other relationships are shown below:

1 cup flour = 5 pancakes

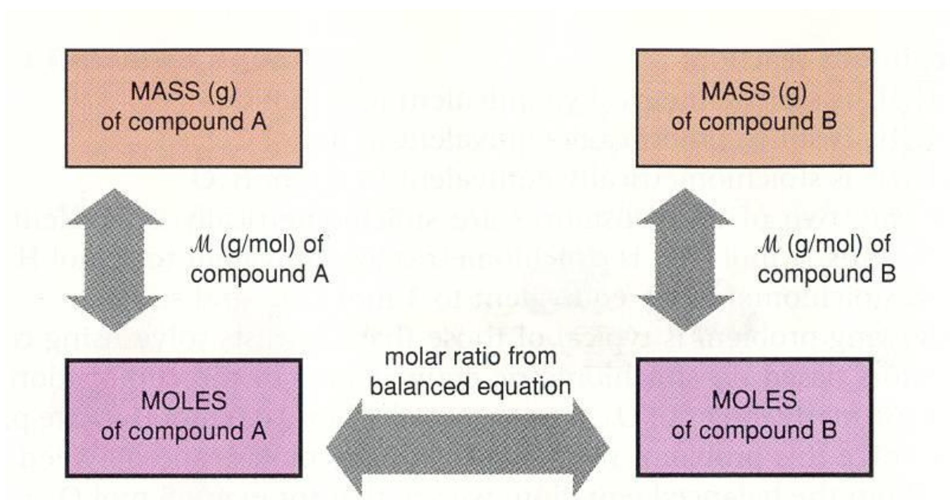
½ tsp baking powder = 5 pancakes

STOICHIOMETRY

- A balanced chemical equation provides several important information about the reactants and products in a chemical reaction. For example:

1 $N_2(g)$	**+**	**3 $H_2(g)$**	**⟶**	**2 $NH_3(g)$**
1 molecule		**3 molecules**		**2 molecules**
100 molecules		**300 molecules**		**200 molecules**
1 million molecules		**3 million molecules**		**2 million molecules**
1 mole N_2	**+**	**3 moles H_2**	**⟶**	**2 moles NH_3**

This is the **MOLE RATIO** between **REACTANTS** and **PRODUCTS**



Summary of Stoichiometric Calculations in Chemistry

<u>Examples:</u>

1. Determine each mole ration based on the reaction shown below:

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

a) $\dfrac{\text{mol } O_2}{\text{mol } CO_2} =$

b) $\dfrac{\text{mol } C_4H_{10}}{\text{mol } H_2O} =$

STOICHIOMETRIC CALCULATIONS

Mole-Mole Calculations:

- Relates moles of reactants and products in a balanced chemical equation

Examples:

1. How many moles of nitrogen will react with 2.4 moles of hydrogen to produce ammonia as shown in the reaction below?

$$1 \quad N_2(g) \quad + \quad 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

$$2.4 \text{ mol } H_2 \times \frac{1 \text{ mol } N_2}{3 \text{ mol } H_2} = 0.80 \text{ mol } N_2$$

Mole Ratio

2. How many moles of ammonia can be produced from 32 moles of hydrogen? (Assume excess nitrogen present)

$$32 \text{ mol } H_2 \times \underline{\qquad\qquad} = \qquad \text{mol } NH_3$$

3. In one experiment, 6.80 mol of ammonia are prepared. How many moles of hydrogen were used up in this experiment?

STOICHIOMETRIC CALCULATIONS

Mass-Mole Calculations:

- Relates moles and mass of reactants or products in a balanced chemical equation

Examples:

1. How many grams of ammonia can be produced from the reaction of 1.8 moles of nitrogen with excess hydrogen as shown below?

$$1\ N_2(g) \quad + \quad 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

$$1.8\text{ mol } N_2 \text{x} \frac{2\text{ mol } NH_3}{1\text{ mol } N_2} \text{x} \frac{17.04\text{ g}}{1\text{ mol}} = 61\text{ g } NH_3$$

Mole Ratio **Molar Mass**

2. How many moles of hydrogen gas are required to produce 75.0 g of ammonia?

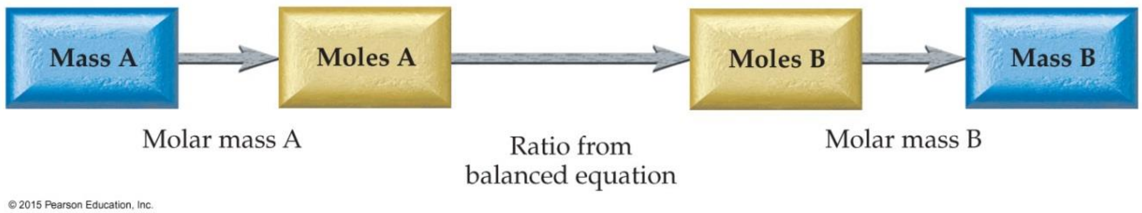
$$75.0\text{g } NH_3 \ \text{x} \ \underline{\hspace{3cm}} \ \text{x} \ \underline{\hspace{3cm}} = \qquad\qquad \text{mol } H_2$$

3. How many moles of ammonia can be produced from the reaction of 125 g of nitrogen as shown above?

STOICHIOMETRIC CALCULATIONS

<u>Mass-Mass Calculations:</u>

- Relates mass of reactants and products in a balanced chemical equation



<u>Examples:</u>

1. What mass of oxygen will be required to react completely with 96.1 g of propane, C_3H_8, according to the equation below?

$$1\ C_3H_8\ (g)\ +\ 5\ O_2\ (g) \longrightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

$$96.1\ g\ C_3H_8 \times \frac{1\ mol\ C_3H_8}{44.11\ g\ C_3H_8} \times \frac{5\ mol\ O_2}{1\ mol\ C_3H_8} \times \frac{32.00\ g\ O_2}{1\ mol\ O_2} = 349\ g\ O_2$$

Molar Mass | Mole Ratio | Molar Mass

2. What mass of carbon dioxide will be produced from the reaction of 175 g of propane, as shown above?

175 g C_3H_8 x ____________ x ____________ x ____________ = ____________ g CO_2

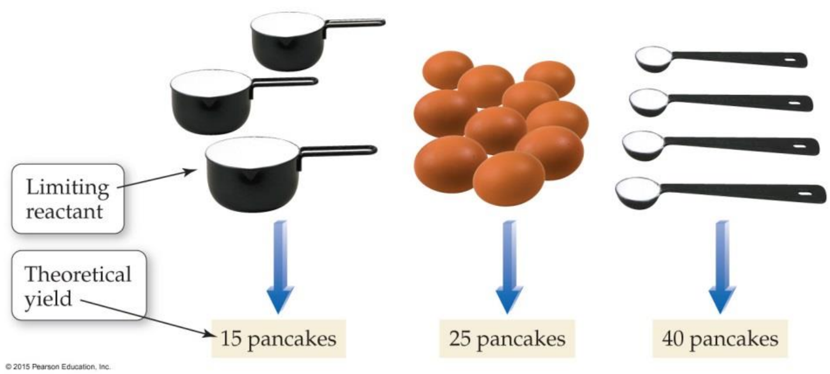
LIMITING REACTANT & THEORETICAL YIELD

- When ***2 or more reactants*** are combined in ***non-stoichiometric*** ratios, the amount of ***product*** produced is ***limited*** by the reactant that is ***not in excess*** (***limiting reactant***).

<u>Analogy:</u>

- Using the pancake analogy presented in the first part of this chapter, how many pancakes can be prepared from 3 cups, flour, 10 eggs and 4 tsp baking powder.

- Using the ratios of pancake with each ingredient, we can calculate that 15 pancakes can be made from using all the flour, 25 pancakes can be made from using all of the eggs, and 40 pancakes can be made from using all the baking powder.



- Consequently, unless there are more ingredients, only 15 pancakes can be made from the combination of the above ingredients. This is due to the flour being completely used up after making 15 pancakes.

- The flour is therefore called the ***limiting reactant***, which is the reactant that produces the smallest possible amount of products, called ***theoretical yield***.

<u>Summary:</u>

- **Limiting Reactant:** the reactant that is completely used during a chemical reaction.

- **Theoretical Yield**: the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.

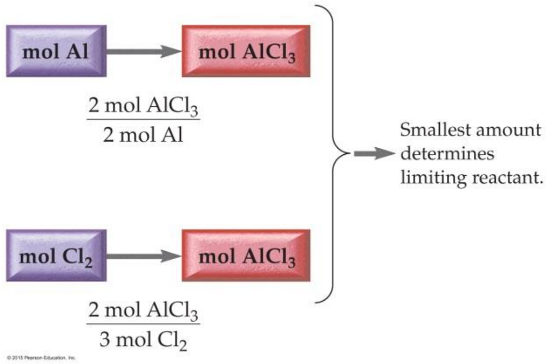
- **Actual Yield**: the amount of product actually produced in a chemical reaction.

LIMITING REACTANT & THEORETICAL YIELD

Examples

1. In the reaction below, if 0.552 moles of aluminum react with 0.887 moles of chlorine, what is the limiting reactant and theoretical yield of $AlCl_3$ in moles?

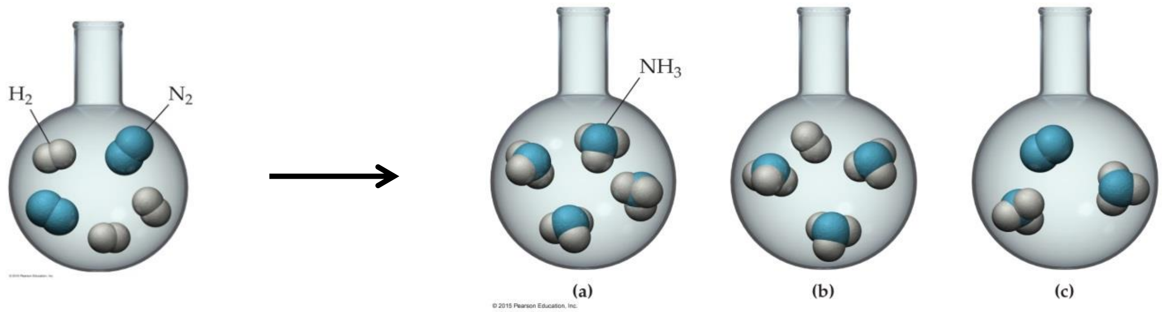
$$2\ Al\ (s) + 3\ Cl_2\ (g) \rightarrow 2\ AlCl_3\ (s)$$



2. Consider the reaction below:

$$N_2\ (g) + 3\ H_2\ (g) \rightarrow 2\ NH_3\ (g)$$

If the flask on the left represents the mixture before the reaction, which flask represents the products after the mixture has completely reacted?

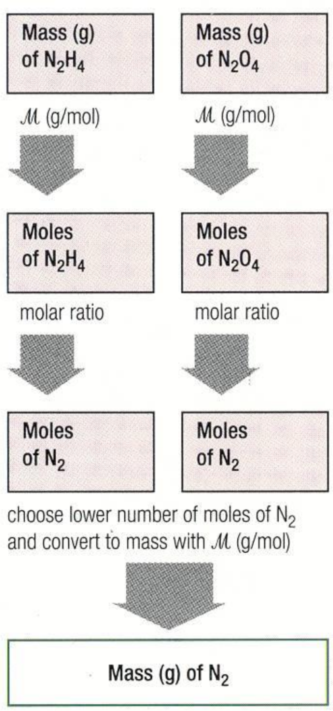


LIMITING REACTANT & THEORETICAL YIELD

3. A fuel mixture used in the early days of rocketry was a mixture of N_2H_4 and N_2O_4, as shown below. How many grams of N_2 gas is produced when 100 g of N_2H_4 and 200 g of N_2O_4 are mixed?

$$2\,N_2H_4\,(l) + N_2O_4\,(l) \rightarrow 3\,N_2\,(g) + 4\,H_2O\,(g)$$

Method 1



$$100\ \cancel{g\ N_2H_4} \times \frac{1\ \cancel{mol\ N_2H_4}}{32.06\ \cancel{g\ N_2H_4}} \times \frac{3\ mol\ N_2}{2\ \cancel{mol\ N_2H_4}} = 4.68\ mol\ N_2$$

Limiting Reactant

$$200\ \cancel{g\ N_2O_4} \times \frac{1\ \cancel{mol\ N_2O_4}}{92.02\ \cancel{g\ N_2O_4}} \times \frac{3\ mol\ N_2}{1\ \cancel{mol\ N_2O_4}} = 6.52\ mol\ N_2$$

$$4.68\ \cancel{mol\ N_2} \times \frac{28.02\ g\ N_2}{1\ \cancel{mol\ N_2}} = 131\ g\ N_2$$

4. In the previous problem, how many grams of excess reacant are left over after all the reaction has completed?

<u>**Step 1**</u>: calculate grams of N_2O_4 that reacted with limiting reactant (LR)

<u>**Step 2**</u>: calculate grams of N_2O_4 that remain unreacted.

PERCENT YIELD

- The ***percent yield*** of a reaction is obtained as follows:

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \text{x100} = \text{Percent yield}$$

Examples:

1. In an experiment forming ethanol, the theoretical yield is 50.5 g and the actual yield is 46.8 g. What is the percent yield for this reaction?

$$\%\text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \text{ x } 100 = \frac{46.8\text{ g}}{50.5\text{ g}} \text{ x } 100 = 92.7\ \%$$

2. Silicon carbide can be formed from the reaction of sand (SiO_2) with carbon as shown below:

$$SiO_2\ (s) + 3\ C\ (s) \longrightarrow SiC\ (s) + 2\ CO\ (g)$$

When 125 g of sand are processed, 68.4 of SiC is produced. What is the percent yield of SiC in this reaction?

$$125\text{ g }SiO_2 \text{ x } \underline{\hspace{3cm}} \text{ x } \underline{\hspace{3cm}} \text{ x } \underline{\hspace{3cm}} =$$

$$\%\text{ yield SiC} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \text{ x } 100 =$$

PERCENT YIELD

Examples:

3. In an experiment to prepare aspirin, the theoretical yield is 153.7 g and the actual yield is 124.0 g. What is the percent yield of this reaction?

$$\text{Percent yield} = \frac{\qquad\qquad}{\qquad\qquad} \times 100 =$$

4. Carbon tetrachloride (CCl_4) was prepared by reacting 100.0 g of Cl_2 with excess carbon disulfide (CS_2), as shown below. If 65.0 g was prepared in this reaction, calculate the percent yield.

$$CS_2 + 3\,Cl_2 \rightarrow CCl_4 + S_2Cl_2$$

Calculate the theoretical yield:

$$100.0 \text{ g } Cl_2 \times \frac{\qquad\qquad}{\qquad\qquad} \times \frac{\qquad\qquad}{\qquad\qquad} \times \frac{\qquad\qquad}{\qquad\qquad} = \qquad\qquad \text{g } CCl_4$$

Calculate percent yield:

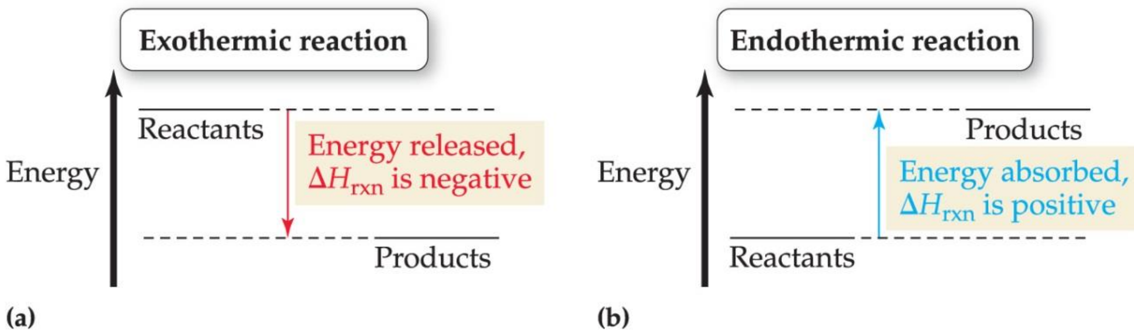
$$\text{Percent yield} = \frac{\qquad\qquad}{\qquad\qquad} \times 100 =$$

ENERGY CHANGES IN CHEMICAL REACTIONS

- ***Heat*** is thermal energy change that is emitted or absorbed when a chemical reaction takes place. In chemistry, thermal energy at constant pressure (a common situation for most chemical reactions) is quantified by a function called ***enthalpy***.

- ***Enthalpy of a reaction (ΔH_{rxn})*** is the amount of thermal energy (or heat) that flows when a reaction occurs at constant pressure.

- The ***direction of heat flow*** depends whether the products in a reaction have more or less energy than the reactants.



- When ***heat is released*** during a chemical reaction, it is said to be ***exothermic***. For exothermic reactions, ΔH is negative.

  $3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g)$ $\qquad \Delta H = -22.0$ kcal

- When ***heat is gained*** during a chemical reaction, it is said to be ***endothermic***. For endothermic reactions, ΔH is positive.

  $2\ H_2O\ (l) \rightarrow 2\ H_2\ (g) + O_2\ (g)$ $\qquad \Delta H = +137$ kcal

CALCULATING HEAT IN A REACTION

- The value of ΔH in a chemical reaction refers to the heat lost or gained for the number of moles of reactants and products in a balanced chemical equation.
  For example, based on the chemical equation shown below:

  $$2\ H_2O\ (l) \rightarrow 2\ H_2\ (g) + O_2\ (g) \qquad \Delta H = +137\ \text{kcal}$$

  The following conversion factors can be written:

  $$\frac{137\ \text{kcal}}{2\ \text{mol}\ H_2O} \quad \text{or} \quad \frac{137\ \text{kcal}}{2\ \text{mol}\ H_2} \quad \text{or} \quad \frac{137\ \text{kcal}}{1\ \text{mol}\ O_2}$$

- These conversion factors can be used to calculate the amount of heat associated with a particular chemical reaction, based on given amounts of reactants and products.

Examples:

1. Given the reaction shown below, how much heat, in kJ, is released when 50.0 g of NH_3 form?

$$3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g) \qquad \Delta H = -92.2\ \text{kJ}$$

2. How many kJ of heat are needed to react 25.0 g of HgO according to the reaction shown below:

$$2\ HgO\ (s) \rightarrow 2\ Hg\ (l) + O_2\ (g) \qquad \Delta H = 182\ \text{kJ}$$

3. Consider the generic reaction shown below:

$$2A + 3\ B \rightarrow 2\ C \qquad \Delta H = -100\ \text{kJ}$$

If a reaction mixture initially contains 5 mol of A and 6 mol of B, how much heat (in kJ) is evolved once the reaction has completed?