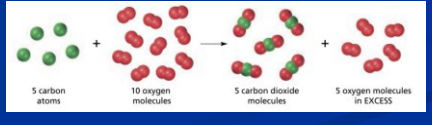
LIMITING REAGENTS AND PERCENT YIELD

Objectives

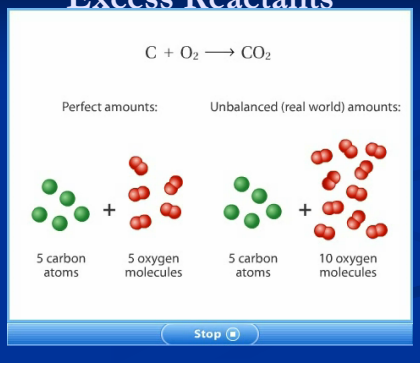
- **Describe** a method for determining which of two reactants is a limiting reactant.
- **Calculate** the amount in moles or mass in grams of a product, given the amounts in moles or masses in grams of two reactants, one of which is in excess.
- **Distinguish** between theoretical yield, actual yield, and percentage yield.
- **Calculate** percentage yield, given the actual yield and quantity of a reactant.

Limiting Reactants

- **The limiting reactant** is the reactant that limits the amount of the other reactant that can combine and the amount of product that can form in a chemical reaction.
- **The excess reactant** is the substance that is not used up completely in a reaction.



Limiting Reactants and Excess Reactants



Sample Problem

- Silicon dioxide (quartz) is usually quite unreactive but reacts readily with hydrogen fluoride according to the following equation.

- $SiO_2(s) + 4HF(g) \rightarrow SiF_4(g) + 2H_2O(l)$

- If 6.0 mol HF is added to 4.5 mol SiO_2, which is the limiting reactant?

Sample Problem Solution

$SiO_2(s) + 4HF(g) \rightarrow SiF_4(g) + 2H_2O(l)$

- **Given:** amount of HF = 6.0 mol
  amount of SiO_2 = 4.5 mol
- **Unknown:** limiting reactant
- **Solution:**

$$\text{mol HF} \times \frac{\text{mol } SiF_4}{\text{mol HF}} = \text{mol } SiF_4 \text{ produced}$$

mole ratio

$$\text{mol } SiO_2 \times \frac{\text{mol } SiF_4}{\text{mol } SiO_2} = \text{mol } SiF_4 \text{ produced}$$

Sample Problem Solution, *continued*

- $SiO_2(s) + 4HF(g) \rightarrow SiF_4(g) + 2H_2O(l)$

$$6.0 \text{ mol HF} \times \frac{1 \text{ mol } SiF_4}{4 \text{ mol HF}} = 1.5 \text{ mol } SiF_4 \text{ produced}$$

$$4.5 \text{ mol } SiO_2 \times \frac{1 \text{ mol } SiF_4}{1 \text{ mol } SiO_2} = 4.5 \text{ mol } SiF_4 \text{ produced}$$

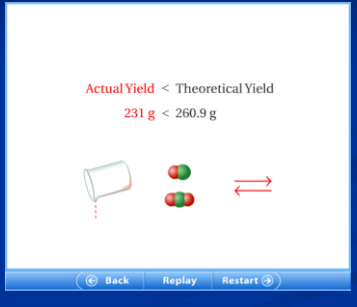
- **HF is the limiting reactant.**

Percentage Yield

- The **theoretical yield** is the maximum amount of product that can be produced from a given amount of reactant.
- The **actual yield** of a product is the measured amount of that product obtained from a reaction.
- The **percentage yield** is the ratio of the actual yield to the theoretical yield, multiplied by 100.

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

Comparing Actual and Theoretical Yield



Percent Yield

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

$C_7H_6O_3 + C_4H_6O_4 \longrightarrow C_9H_8O_4 + C_2H_4O_2$

- Actual yield: 231 g $C_9H_8O_4$
- Theoretical yield: 260.9 g $C_9H_8O_4$
- Percent yield: ? %

percent yield = 88.5%

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Percentage Yield

Sample Problem

- Chlorobenzene, C_6H_5Cl, is used in the production of many important chemicals, such as aspirin, dyes, and disinfectants. One industrial method of preparing chlorobenzene is to react benzene, C_6H_6, with chlorine, as represented by the following equation.
  - $C_6H_6(l) + Cl_2(g) \rightarrow C_6H_5Cl(l) + HCl(g)$
- When 36.8 g C_6H_6 react with an excess of Cl_2, the actual yield of C_6H_5Cl is 38.8 g.
- What is the percentage yield of C_6H_5Cl?

Percentage Yield

Sample Problem Solution

$C_6H_6(l) + Cl_2(g) \rightarrow C_6H_5Cl(l) + HCl(g)$

- **Given:** mass of C_6H_6 = 36.8 g
- mass of Cl_2 = excess
- actual yield of C_6H_5Cl = 38.8 g
- **Unknown:** percentage yield of C_6H_5Cl
- **Solution:**
- Theoretical yield

$$\text{g } C_6H_6 \times \underbrace{\frac{\text{mol } C_6H_6}{\text{g } C_6H_6}}_{\textit{molar mass factor}} \times \underbrace{\frac{\text{mol } C_6H_5Cl}{\text{mol } C_6H_6}}_{\textit{mol ratio}} \times \underbrace{\frac{\text{g } C_6H_5Cl}{\text{mol } C_6H_5Cl}}_{\textit{molar mass}} = \text{g } C_6H_5Cl$$

Percentage Yield, *continued*

Sample Problem H Solution, *continued*

$$C_6H_6(l) + Cl_2(g) \rightarrow C_6H_5Cl(l) + HCl(g)$$

- Theoretical yield

$$36.8\ g\ C_6H_6 \times \frac{1\ mol\ C_6H_6}{78.12\ g\ C_6H_6} \times \frac{1\ mol\ C_6H_5Cl}{1\ mol\ C_6H_6} \times \frac{112.56\ g\ C_6H_5Cl}{1\ mol\ C_6H_5Cl}$$

$$= 53.0\ g\ C_6H_5Cl$$

- Percentage yield

$$\text{percent yield } C_6H_5Cl = \frac{\text{actual yield}}{\text{theorectical yield}} \times 100$$

$$\text{percent yield} = \frac{38.8\ g}{53.0\ g} \times 100 = 73.2\%$$