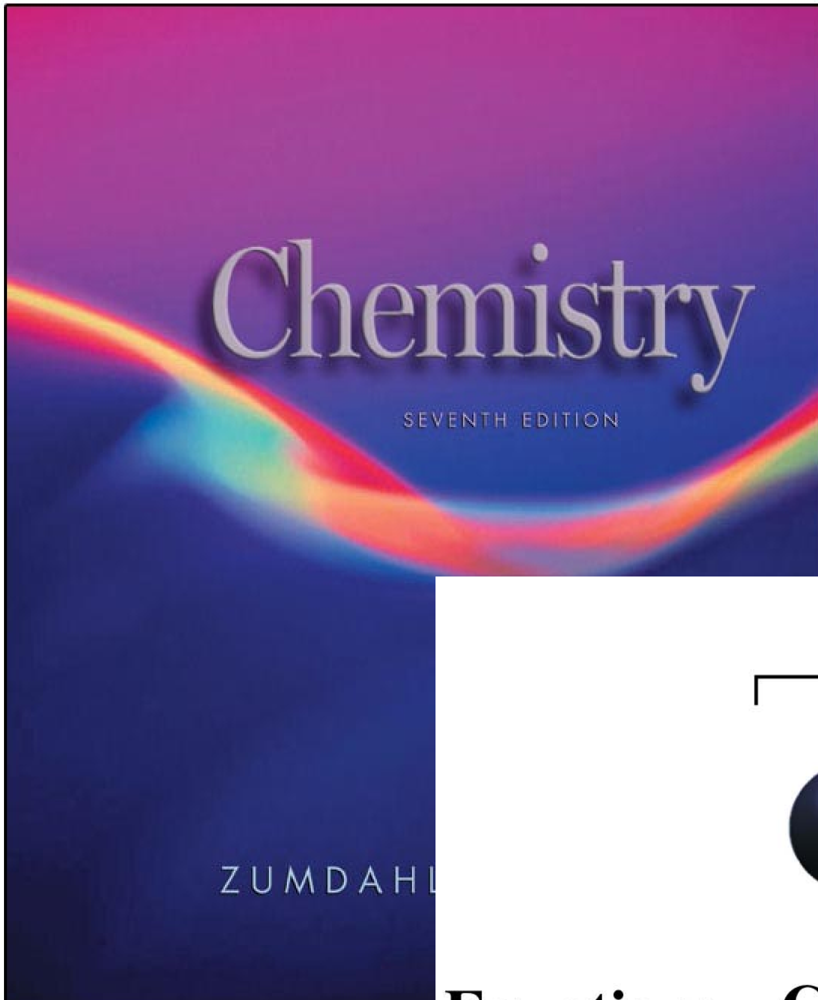
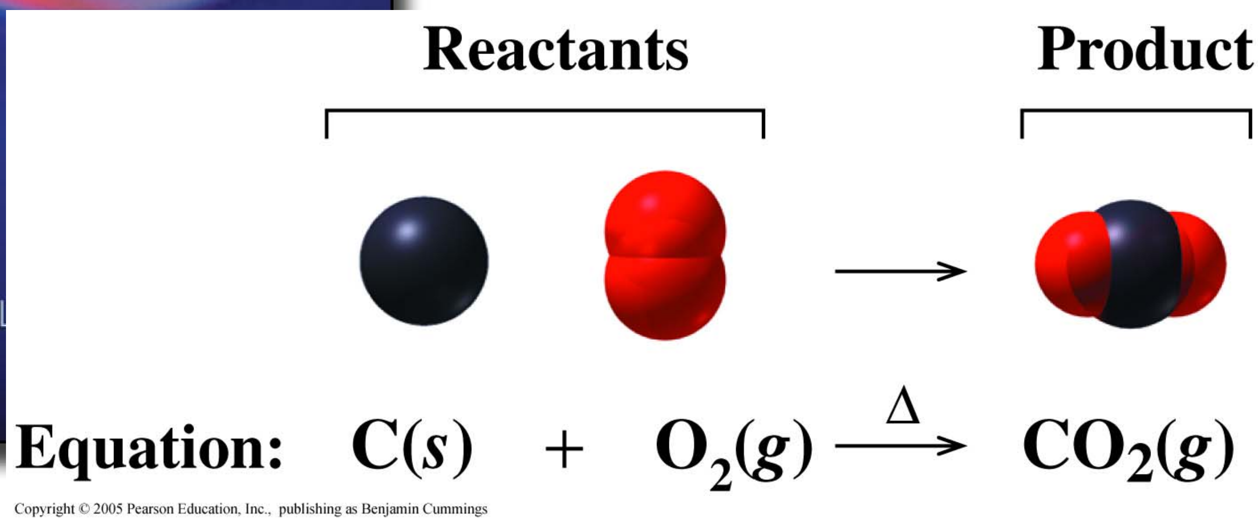


3.10 Calculations Involving a Limiting Reactant



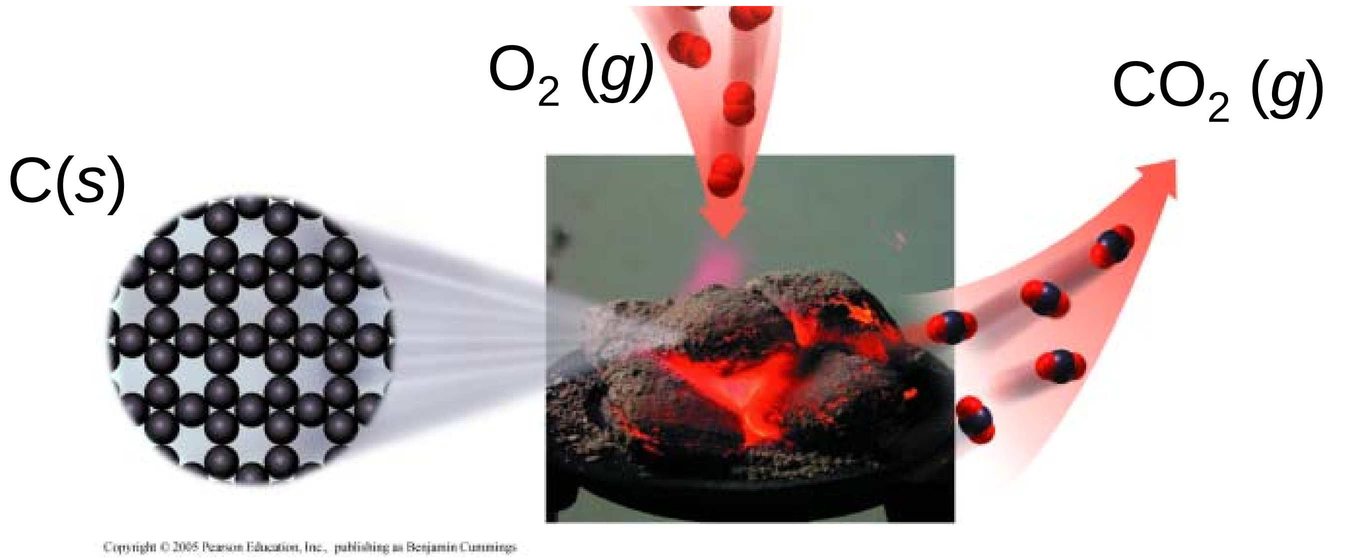
Definitions

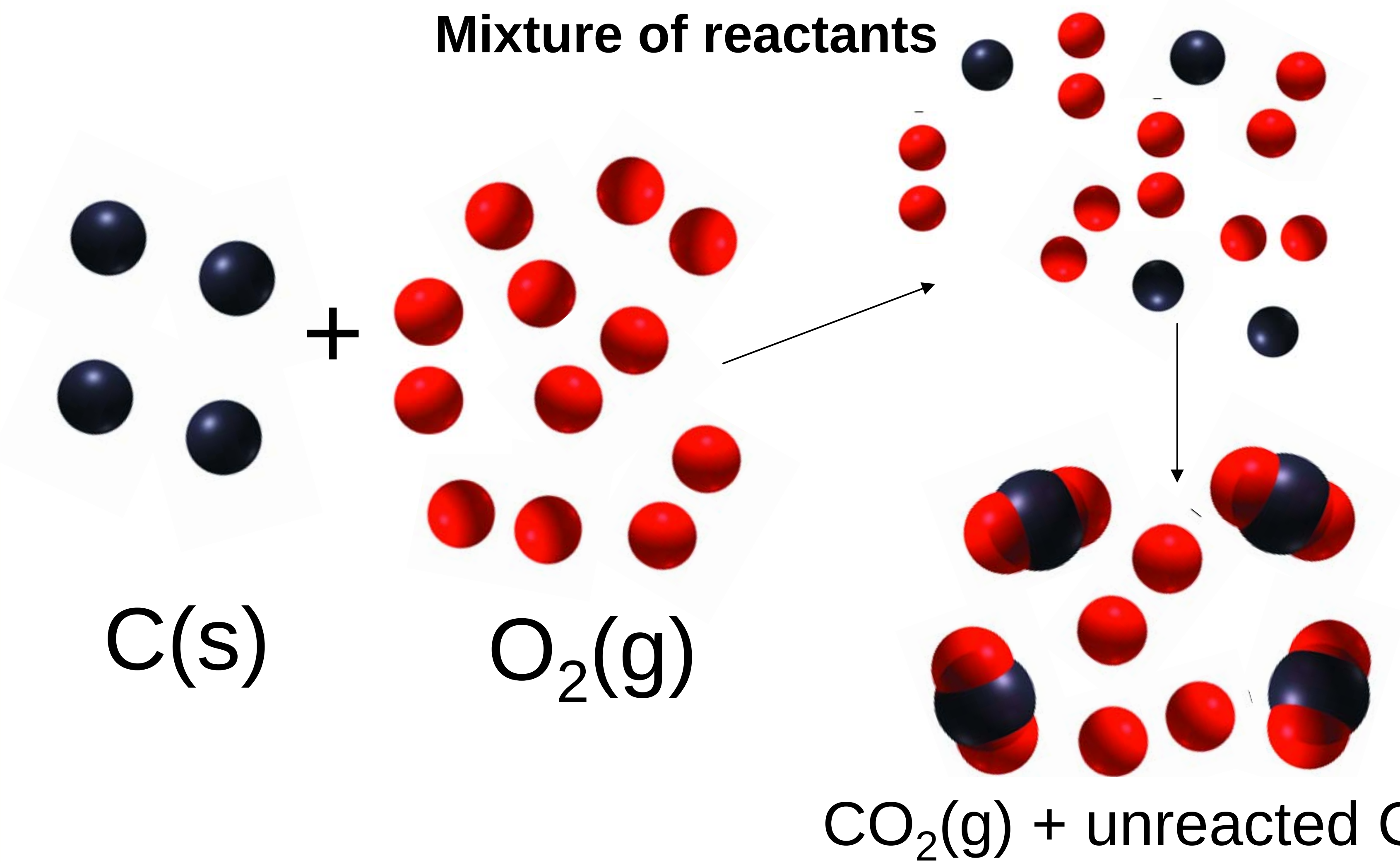
- Limiting-reactant principle – The maximum amount of product possible from a reaction is determined by the amount of reactant present in the least amount, based on its reaction coefficient and molecular weight.
- Limiting reactant – the reactant present in a reaction in the least amount, based on its reaction coefficients and molecular weight. It is the reactant that determines the maximum amount of product that can be formed.

Chemical Reaction for CO_2

Reactants **Product**

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$




Mixture of reactants
$C(s)$
+
$O_2(g)$
$CO_2(g)$ + unreacted O_2

Learning Check

For the balanced equation shown below, what would be the limiting reagent if 79.6 grams of NO were reacted with 59.5 grams of O_2?

$2NO + O_2 \Rightarrow 2NO_2$; NO or O_2

Solution

To answer this question, calculate the grams of NO_2 needed to react fully with 79.6 grams of NO and 59.5 grams of O_2, by using the balanced equation.

$$\frac{79.6\ \cancel{\text{g of NO}}}{} \times \frac{1\ \cancel{\text{mol of NO}}}{30.0\ \cancel{\text{g of NO}}} \times \frac{2\ \cancel{\text{mol of } NO_2}}{2\ \cancel{\text{mol of NO}}} \times \frac{46\ \text{grams of } NO_2}{1\ \cancel{\text{mol of } NO_2}}$$

$$= 122.05 \text{ grams of } NO_2$$

$$\frac{59.5\ \cancel{\text{g of } O_2}}{} \times \frac{1\ \cancel{\text{mol of } O_2}}{32.0\ \cancel{\text{g of } O_2}} \times \frac{2\ \cancel{\text{mol of } NO_2}}{1\ \cancel{\text{mol of } O_2}} \times \frac{46\ \text{grams of } NO_2}{1\ \cancel{\text{mol of } NO_2}}$$

$$= 171.06 \text{ grams of } NO_2$$

There is less NO_2 with NO than O_2, therefore, the limiting reactant is NO

Learning Check

For the balanced equation shown below, if 19.1 grams of CH_5N were reacted with 88.1 grams of O_2, how many grams of CO_2 would be produced, using the limited reactant to determine the quantity of a product that should be produced ?

$$4CH_5N + 11O_2 \Rightarrow 4CO_2 + 10H_2O + 4NO$$

Solution

To answer this question, calculate the grams of NO_2 needed to react fully with 19.1 grams of CH_5N and 88.1 grams of O_2, by using the balanced equation.

$$\cancel{19.1\ \text{g of}\ CH_5N} \times \frac{1\ \text{mol}\ \cancel{\text{of}\ CH_5N}}{31.0\ \cancel{\text{g of}\ CH_5N}} \times \frac{4\ \cancel{\text{mol of}\ CO_2}}{4\ \cancel{\text{mol of}}\ CH_5O} \times \frac{44\ \text{g of}\ CO_2}{1\ \cancel{\text{mol of}\ NO_2}}$$

$$= 21.1\ \text{grams of}\ CO_2$$

$$\cancel{88.1\ \text{g of}\ O_2} \times \frac{1\ \cancel{\text{mol of}\ O_2}}{32.0\ \cancel{\text{g of}\ O_2}} \times \frac{4\ \cancel{\text{mol of}\ CO_2}}{11\ \cancel{\text{mol of}\ O_2}} \times \frac{44\ \text{grams of}\ CO_2}{1\ \cancel{\text{mol of}\ CO_2}}$$

$$= 44.1\ \text{grams of}\ CO_2$$

There is 21.1 g of CO_2 produced with CH_5N than O_2, which is the limiting reactant

Percent Yield

- Percentage yield – the percentage of the theoretical amount of a product actually produced by a reaction.
- Actual yield – the mass product obtained in an experiment.
- Theoretical yield – the mass calculated to give the maximum amount of product.

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

Learning Check

A chemist wants to produce urea (N_2CH_4O) by reacting ammonia (NH_3) and carbon dioxide (CO_2). The balanced equation for the reaction is

$$2NH_3(g) + CO_2\,(g) \rightarrow N_2CH_4O(s) + H_2O(l)$$

The chemist reacts 5.11 g NH_3 with excess CO_2 and isolates 3.12 g of solid N_2CH_4O. Calculate the percentage yield of the experiment.

Solution

To answer this question, calculate first the theoretical yield of N_2CH_4O that should be made. Then use the actual yield to calculate the percentage yield.

$$5.11 \cancel{\text{ g of } NH_3} \times \frac{1 \cancel{\text{ mol of } NH_3}}{17.0 \cancel{\text{ g of } NH_3}} \times \frac{1 \cancel{\text{ mol of } N_2CH_4O}}{2 \cancel{\text{ mol of } NH_3}} \times \frac{60.1 \text{ g of } N_2CH_4O}{1 \cancel{\text{ mol of } N_2CH_4O}}$$

$= 9.03$ grams of N_2CH_4O is the theoretical yield

The actual yield was 3.12 g of N_2CH_4O, so the % yield is

$$\%\text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{3.12 \text{ g}}{9.03 \text{ g}} \times 100 = 34.6\%$$

Learning Check

- Methanol (CH_3OH), also called methyl alcohol, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combination of gaseous carbon monoxide and hydrogen. Suppose 68.5 kg CO(g) is reacted with 8.60 kg H_2(g). Calculate the theoretical yield of methanol. If $3.57\text{x}10^4$ g CH_3OH is actually produced, what is the percent yield of methanol?