Chapter 8

Quantities in Chemical Reactions

Based on slides provided with Introductory Chemistry, Fifth Edition
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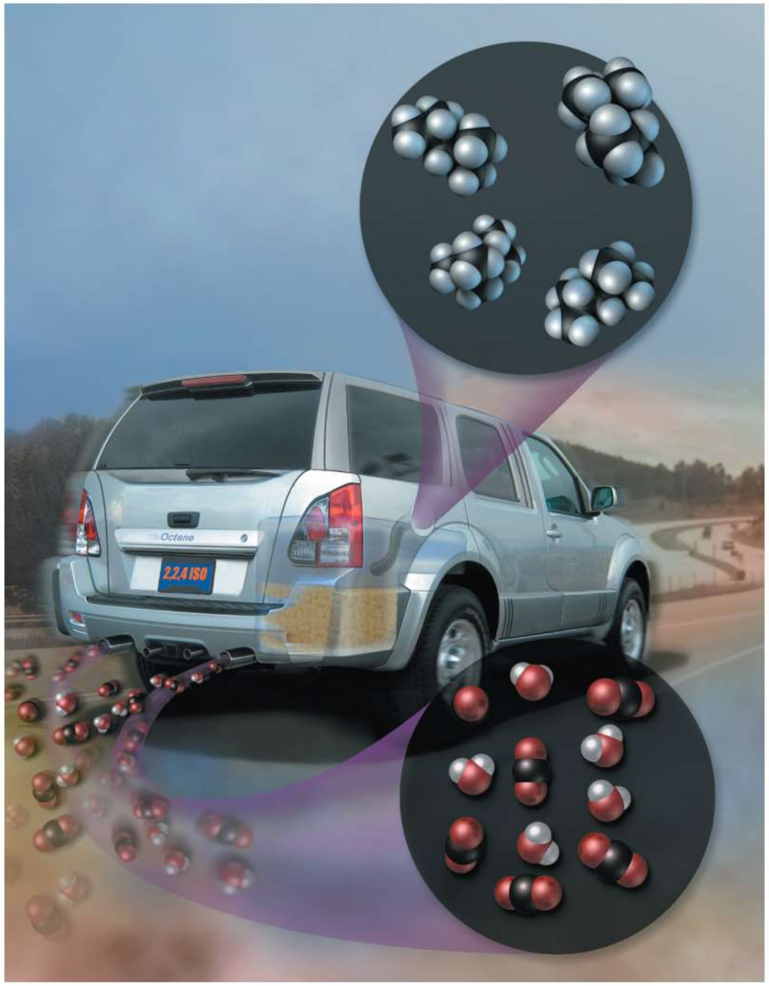
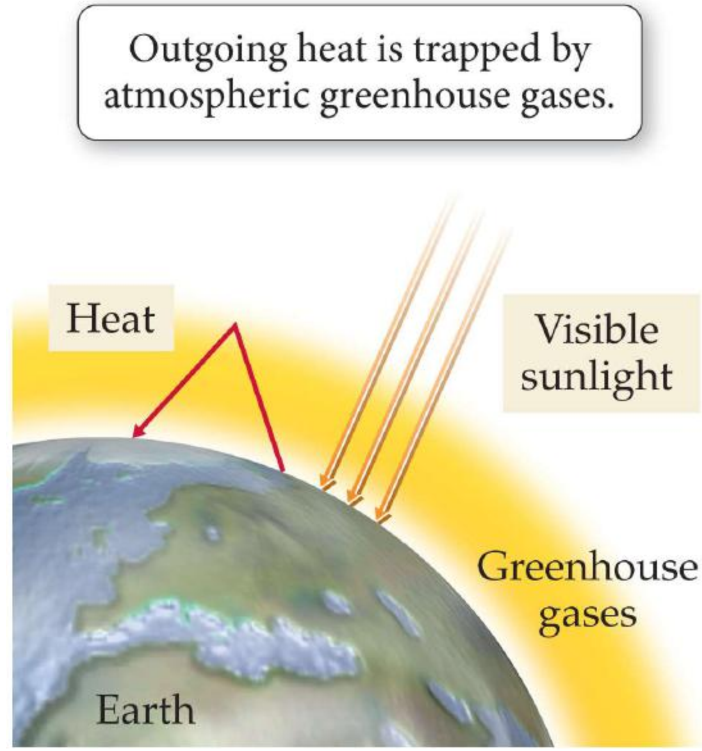
Global Warming: Too Much Carbon Dioxide

- The combustion of fossil fuels such as octane (shown here) produces water and carbon dioxide as products.
- Carbon dioxide is a greenhouse gas that is believed to be responsible for global warming.



The Greenhouse Effect

- Greenhouse gases act like glass in a greenhouse, allowing visible-light energy to enter the atmosphere but preventing heat energy from escaping.
- Outgoing heat is trapped by greenhouse gases such as CO_2.



Combustion of Fossil Fuels Produces CO_2

- Consider the combustion of octane (C_8H_{18}), a component of gasoline:

  $$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

- The balanced chemical equation shows that 16 mol of CO_2 are produced for every 2 mol of octane burned.

Combustion of Fossil Fuels Produces CO_2

- Since we know the world's annual fossil fuel consumption, we can estimate the world's annual CO_2 production using the balanced chemical equation.

- Calculation shows that the world's annual CO_2 production—from fossil fuel combustion—matches the measured annual atmospheric CO_2 increase, implying that fossil fuel combustion is indeed responsible for increased atmospheric CO_2 levels.

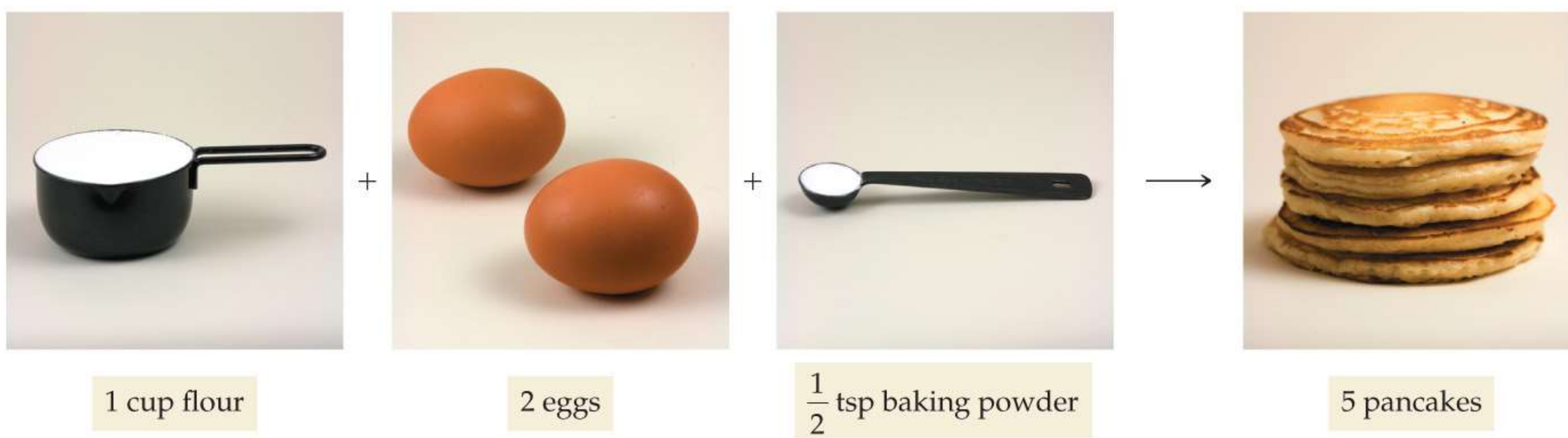
Stoichiometry: Relationships between Ingredients

- The numerical relationship between chemical quantities in a balanced chemical equation is called reaction **stoichiometry**.
- We can predict the amounts of products that form in a chemical reaction based on the amounts of reactants.
- We can predict how much of the reactants are necessary to form a given amount of product.
- We can predict how much of one reactant is required to completely react with another reactant.

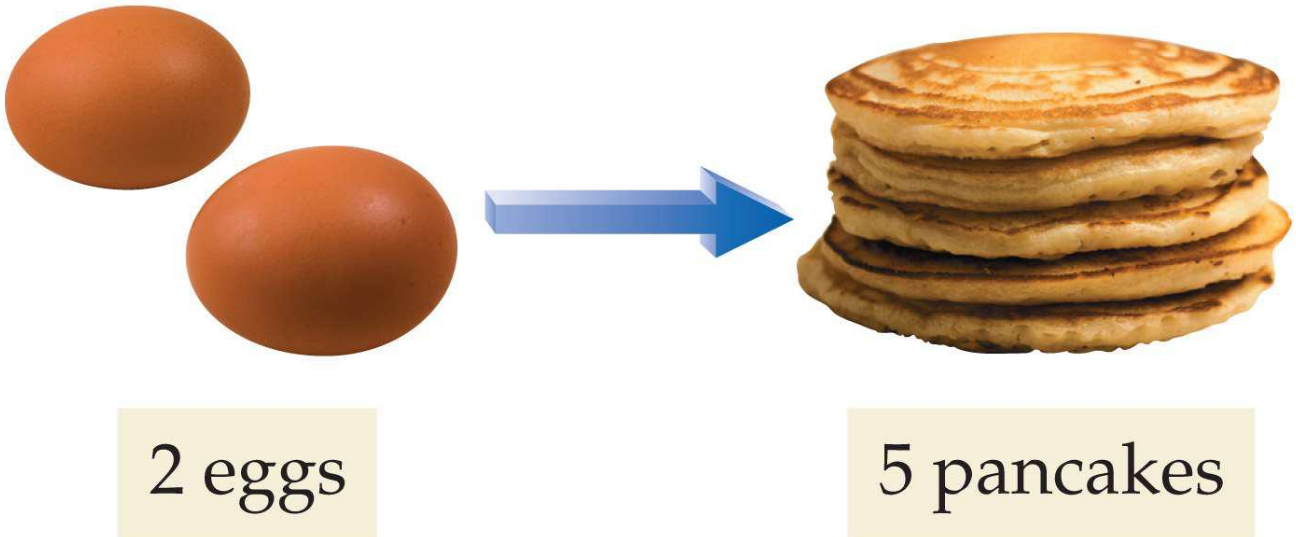
Making Pancakes: Relationships between Ingredients

- A recipe gives numerical relationships between the ingredients and the number of pancakes.

1 cup flour + 2 eggs + $\frac{1}{2}$ tsp baking powder $\rightarrow$ 5 pancakes

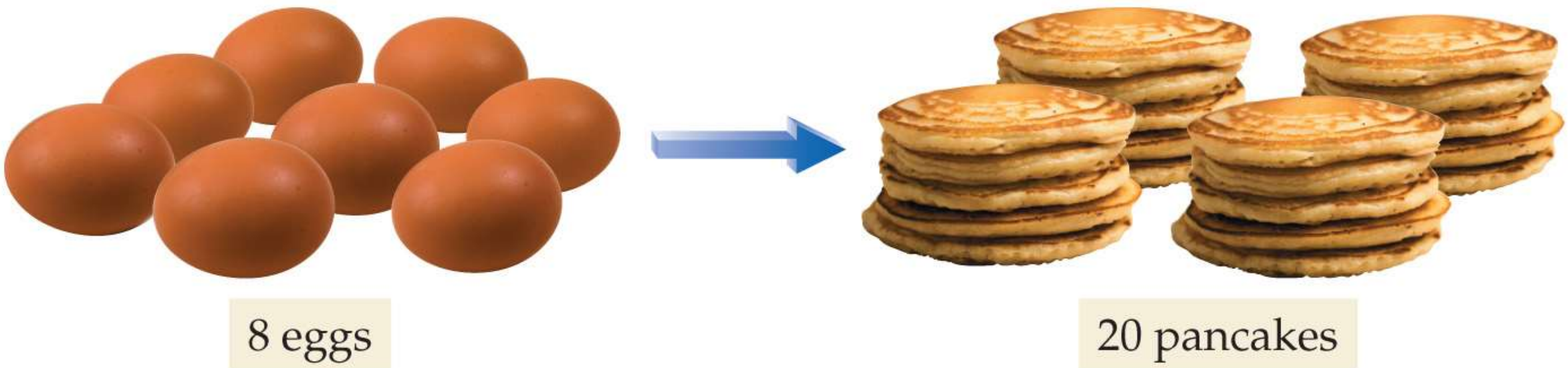


Making Pancakes: Relationships between Ingredients



- The recipe shows the numerical relationships between the pancake ingredients.
- If we have 2 eggs—and enough of everything else—we can make 5 pancakes.
- We can write this relationship as a ratio.
- 2 eggs:5 pancakes

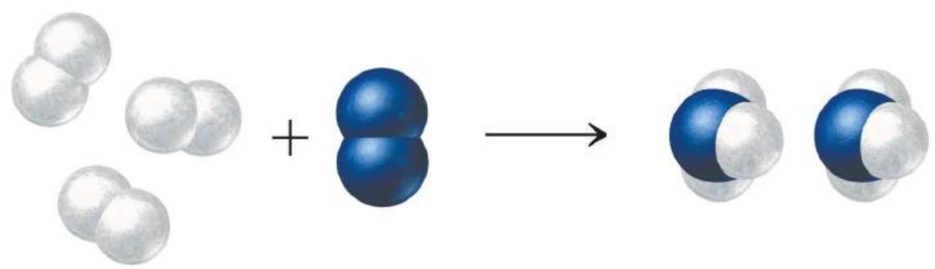
What if we have 8 eggs? Assuming that we have enough of everything else, how many pancakes can we make?



$$8\ \cancel{\text{eggs}} \times \frac{5\ \text{pancakes}}{2\ \cancel{\text{eggs}}} = 20\ \text{pancakes}$$

Making Molecules: Mole-to-Mole Conversions

- In a balanced chemical equation, we have a "recipe" for how reactants combine to form products.
- The following equation shows how hydrogen and nitrogen combine to form ammonia (NH_3).



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

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

The balanced equation shows that 3 H_2 molecules react with 1 N_2 molecule to form 2 NH_3 molecules.

We can express these relationships as ratios.

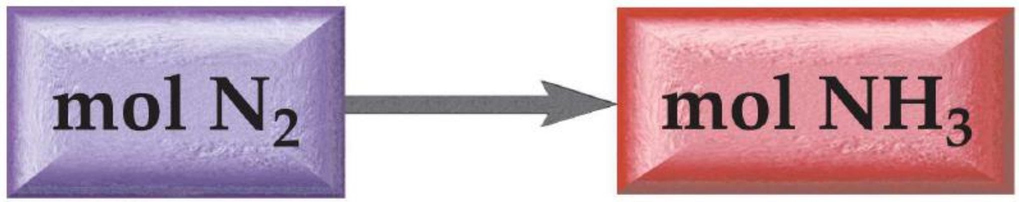
3 H_2 molecules : 1 N_2 molecule : 2 NH_3 molecules

We don’t normally deal with individual molecules, and we can express the same ratios in moles.

3 mol H_2 : 1 mol N_2 : 2 mol NH_3

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

- If we have 3 mol of N_2, and more than enough H_2, how much NH_3 can we make?



$$\frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2}$$

$$3\ \cancel{\text{mol } N_2} \times \frac{2 \text{ mol } NH_3}{1\ \cancel{\text{mol } N_2}} = 6 \text{ mol } NH_3$$

Stoichiometry in Action: Not Enough Oxygen When Burning Octane

- The balanced equation shows that 2 moles of octane require 25 moles of oxygen to burn completely:

  $2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$

- In the case of octane, a shortage of O_2 causes side reactions that result in pollutants such as carbon monoxide (CO) and ozone.
- The 1990 amendments to the Clean Air Act required oil companies to put additives in gasoline that increased its oxygen content.

Stoichiometry in Action: Controversy over Oxygenated Fuels

- MTBE (methyl tertiary butyl ether, $CH_3OC(CH_3)_3$) was the additive of choice by the oil companies.
- MTBE is a compound that does not biodegrade readily.
- MTBE made its way into drinking water through gasoline spills at gas stations, from boat motors, and from leaking underground storage tanks.
- Ethanol (C_2H_5OH), made from the fermentation of grains, is now used as a substitute for MTBE to increase oxygen content in motor fuel.
- Ethanol was not used originally because it was more expensive.

Practice:
Check your solution on next page

According to the following equation, how many moles of water are made in the combustion of 0.10 moles of glucose ($C_6H_{12}O_6$)?

$$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$$

Practice

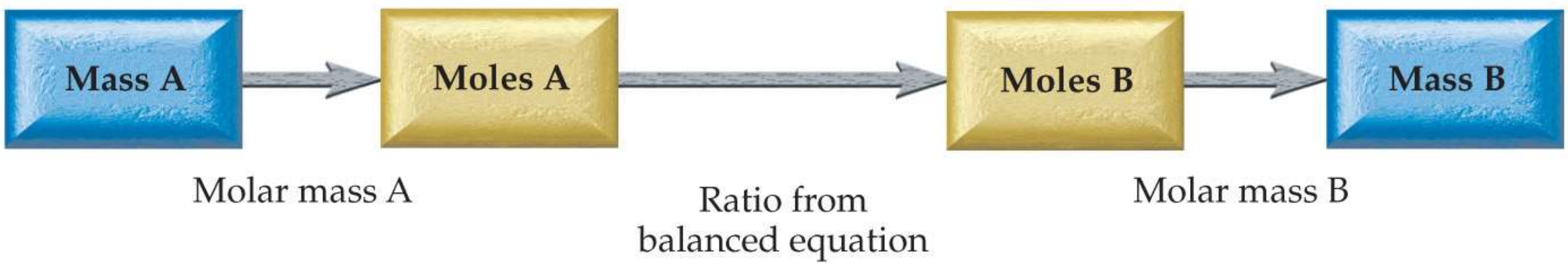
According to the following equation, how many moles of water are made in the combustion of 0.10 moles of glucose ($C_6H_{12}O_6$)?

$$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$$

0.1 mol $C_6H_{12}O_6 \times$(6mol H_2O/1mol $C_6H_{12}O_6$)=
0.6mol H_2O

Making Molecules: Mass-to-Mass Conversions

- A *chemical equation* contains conversion factors between *moles* of reactants and *moles* of products.
- We are often interested in relationships between *mass* of reactants and *mass* of products.
- The general outline for this type of calculation is:



$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$

- What mass of carbon dioxide is emitted by an automobile per 5.0×10^2 g pure octane used?
- The balanced chemical equation gives us a relationship between moles of C_8H_{18} and moles of CO_2.
- Before using that relationship, we must convert from grams to moles.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

SOLUTION MAP:

$$\text{g } C_8H_{18} \xrightarrow{\dfrac{1 \text{ mol } C_8H_{18}}{114.3 \text{ g } C_8H_{18}}} \text{mol } C_8H_{18} \xrightarrow{\dfrac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}}} \text{mol } CO_2 \xrightarrow{\dfrac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}} \text{g } CO_2$$

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

SOLUTION:

$$5.0 \times 10^2\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.3\ \cancel{\text{g } C_8H_{18}}} \times \frac{16\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{44.01\ \text{g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 1.5 \times 10^3\ \text{g } CO_2$$

Practice:
Check your solution on next page

How many grams of glucose can be synthesized from 37.8 g of CO_2 in photosynthesis?

$6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6\ O_2$

Practice: How many grams of glucose can be synthesized from 37.8 g of CO_2 in photosynthesis?

Given:	37.8 g CO_2, $6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6\ O_2$
Find:	g $C_6H_{12}O_6$
Conceptual Plan:	g CO_2 → mol CO_2 → mol $C_6H_{12}O_6$ → g $C_6H_{12}O_6$ $\frac{1\,\text{mol}}{44.01\,\text{g}}$ $\frac{1\,\text{mol}\ C_6H_{12}O_6}{6\,\text{mol}\ CO_2}$ $\frac{180.2\,\text{g}}{1\,\text{mol}}$
Relationships:	1 mol $C_6H_{12}O_6$ = 180.2g, 1 mol CO_2 = 44.01g, 1 mol $C_6H_{12}O_6$: 6 mol CO_2

Solution:

$$37.8\ \cancel{\text{g}\ CO_2} \times \frac{1\ \cancel{\text{mol}\ CO_2}}{44.01\ \cancel{\text{g}\ CO_2}} \times \frac{1\ \cancel{\text{mol}\ C_6}H_{12}O_6}{6\ \cancel{\text{mol}\ CO_2}} \times \frac{180.2\ \text{g}\ C_6H_{12}O_6}{1\ \cancel{\text{mol}\ C_6}H_{12}O_6}$$
$$= 25.8\ \text{g}\ C_6H_{12}O_6$$

Check: because 6x moles of CO_2 as $C_6H_{12}O_6$, but the molar mass of $C_6H_{12}O_6$ is 4x CO_2, the number makes sense

Practice:

Check your solution on next page

How many grams of O_2 can be made from the decomposition of 100.0 g of PbO_2?

$2\ PbO_2(s) \rightarrow 2\ PbO(s) + O_2(g)$

(PbO_2 = 239.2, O_2 = 32.00)

Practice — How many grams of O_2 can be made from the decomposition of 100.0 g of PbO_2?

$$2\,PbO_2(s) \rightarrow 2\,PbO(s) + O_2(g)$$

Given: **Find:**	100.0 g PbO_2, $2\,PbO_2 \rightarrow 2\,PbO + O_2$ g O_2
Conceptual Plan:	g PbO_2 $\xrightarrow{\frac{1\text{ mol}}{239.2\text{ g}}}$ mol PbO_2 $\xrightarrow{\frac{1\text{ mol }O_2}{2\text{ mol }PbO_2}}$ mol O_2 $\xrightarrow{\frac{32.00\text{ g}}{1\text{ mol}}}$ g O_2
Relationships:	1 mol O_2 = 32.00g, 1 mol PbO_2 = 239.2g, 1 mol O_2 : 2 mol PbO_2
Solution:	$100.0\ \cancel{\text{g }PbO_2} \times \frac{1\ \cancel{\text{mol }PbO_2}}{239.2\ \cancel{\text{g }PbO_2}} \times \frac{1\ \cancel{\text{mol }O_2}}{2\ \cancel{\text{mol }PbO_2}} \times \frac{32.00\text{ g }O_2}{1\ \cancel{\text{mol }O_2}}$ $= 6.689\text{ g }O_2$
Check:	because ½ moles of O_2 as PbO_2, and the molar mass of PbO_2 is 7x O_2, the number makes sense

Limiting Reactant, Theoretical Yield, and Percent Yield

More pancakes …

Recall the original equation:

$$1 \text{ cup flour} + 2 \text{ eggs} + \tfrac{1}{2} \text{ tsp baking powder} \longrightarrow 5 \text{ pancakes}$$

Limiting Reactant, Theoretical Yield, and Percent Yield

$$1 \text{ cup flour} + 2 \text{ eggs} + \tfrac{1}{2} \text{ tsp baking powder} \longrightarrow 5 \text{ pancakes}$$

- Suppose we have 3 cups flour, 10 eggs, and 4 tsp baking powder.
- How many pancakes can we make?

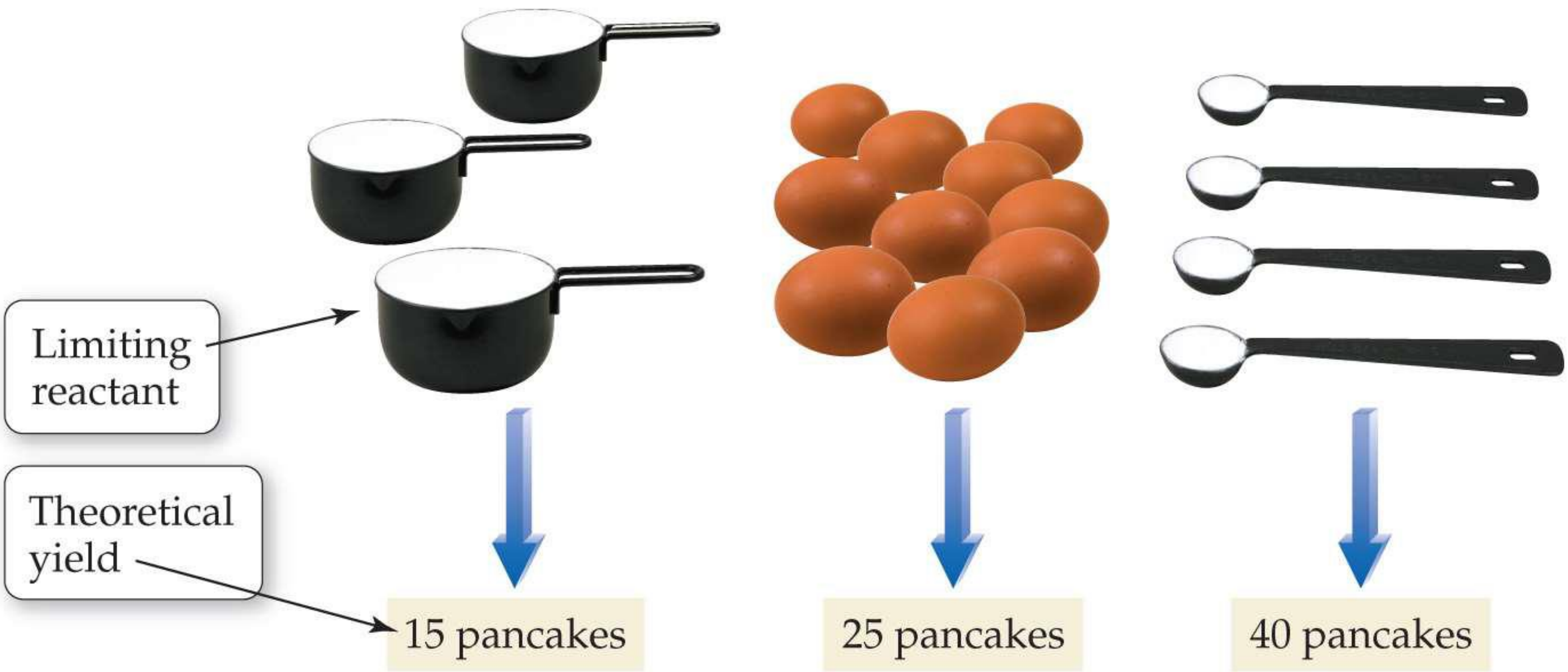
$$3 \cancel{\text{ cups flour}} \times \frac{5 \text{ pancakes}}{1 \cancel{\text{ cup flour}}} = 15 \text{ pancakes}$$

$$10 \cancel{\text{ eggs}} \times \frac{5 \text{ pancakes}}{2 \cancel{\text{ eggs}}} = 25 \text{ pancakes}$$

$$4 \cancel{\text{ tsp baking powder}} \times \frac{5 \text{ pancakes}}{\frac{1}{2} \cancel{\text{ tsp baking powder}}} = 40 \text{ pancakes}$$

We have enough flour for 15 pancakes, enough eggs for 25 pancakes, and enough baking powder for 40 pancakes.

Limiting Reactant, Theoretical Yield, and Percent Yield



If this were a chemical reaction, the flour would be the limiting reactant and 15 pancakes would be the theoretical yield.

Limiting Reactant, Theoretical Yield, and Percent Yield

- Suppose we cook our pancakes. We accidentally burn 3 of them and 1 falls on the floor.
- So even though we had enough flour for 15 pancakes, we finished with only 11 pancakes.
- If this were a chemical reaction, the 11 pancakes would be our **actual yield**, the amount of product actually produced by a chemical reaction.

Limiting Reactant, Theoretical Yield, and Percent Yield

- Our **percent yield**, the percentage of the theoretical yield that was actually attained, is:

$$\text{Percent yield} = \frac{11 \text{ pancakes}}{15 \text{ pancakes}} \times 100\% = 73\%$$

Since 4 of the pancakes were ruined,
we got only 73% of our theoretical yield.

Actual Yield and Percent Yield

- The actual yield of a chemical reaction must be determined experimentally and depends on the reaction conditions.
- The actual yield is almost always less than 100%.
- Some of the product does not form.
- Product is lost in the process of recovering it.

Limiting Reactant, Theoretical Yield, Actual Yield, and Percent Yield

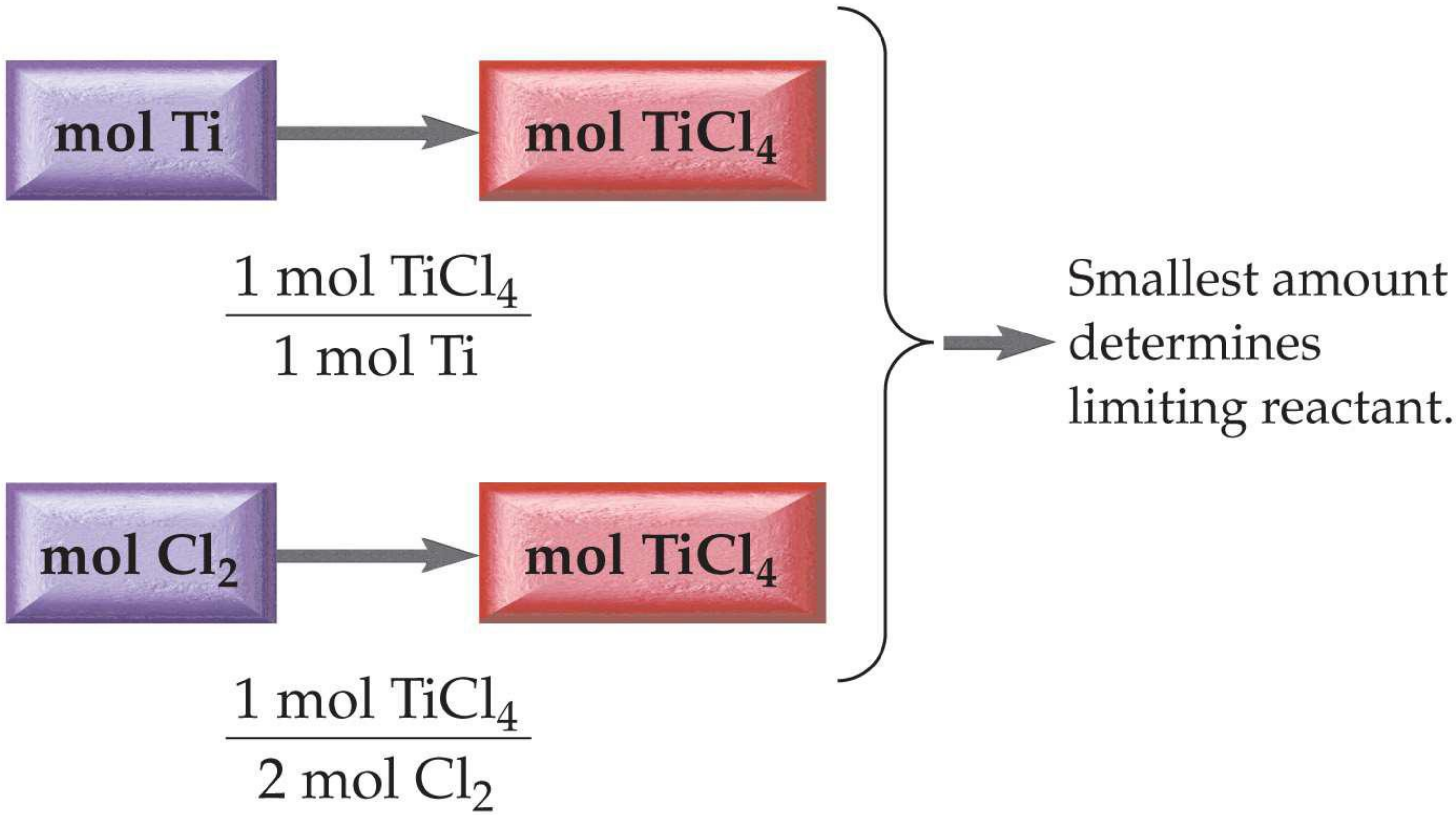
To summarize:

- **Limiting reactant (or limiting reagent)**—the reactant that is completely consumed in 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 by a chemical reaction.
- **Percent yield**—(actual yield/theoretical yield)×100%

Limiting Reactant and Percent Yield: Mole to Mole

Example: $Ti(s) + 2\ Cl_2(g) \rightarrow TiCl_4(s)$
Given (***moles***): 1.8 mol Ti and 3.2 mol Cl_2
Find: limiting reactant and theoretical yield

SOLUTION MAP:



Limiting Reactant and Percent Yield: Mole to Mole

Example: $Ti(s) + 2\ Cl_2(g) \rightarrow TiCl_4(s)$
Given (***moles***): 1.8 mol Ti and 3.2 mol Cl_2
Find: limiting reactant and theoretical yield
SOLUTION:

$$1.8\ \cancel{\text{mol Ti}} \times \frac{1\ \text{mol TiCl}_4}{1\ \cancel{\text{mol Ti}}} = 1.8\ \text{mol TiCl}_4$$

$$3.2\ \cancel{\text{mol Cl}_2} \times \frac{1\ \text{mol TiCl}_4}{2\ \cancel{\text{mol Cl}_2}} = \boxed{1.6\ \text{mol TiCl}_4}$$

Limiting reactant

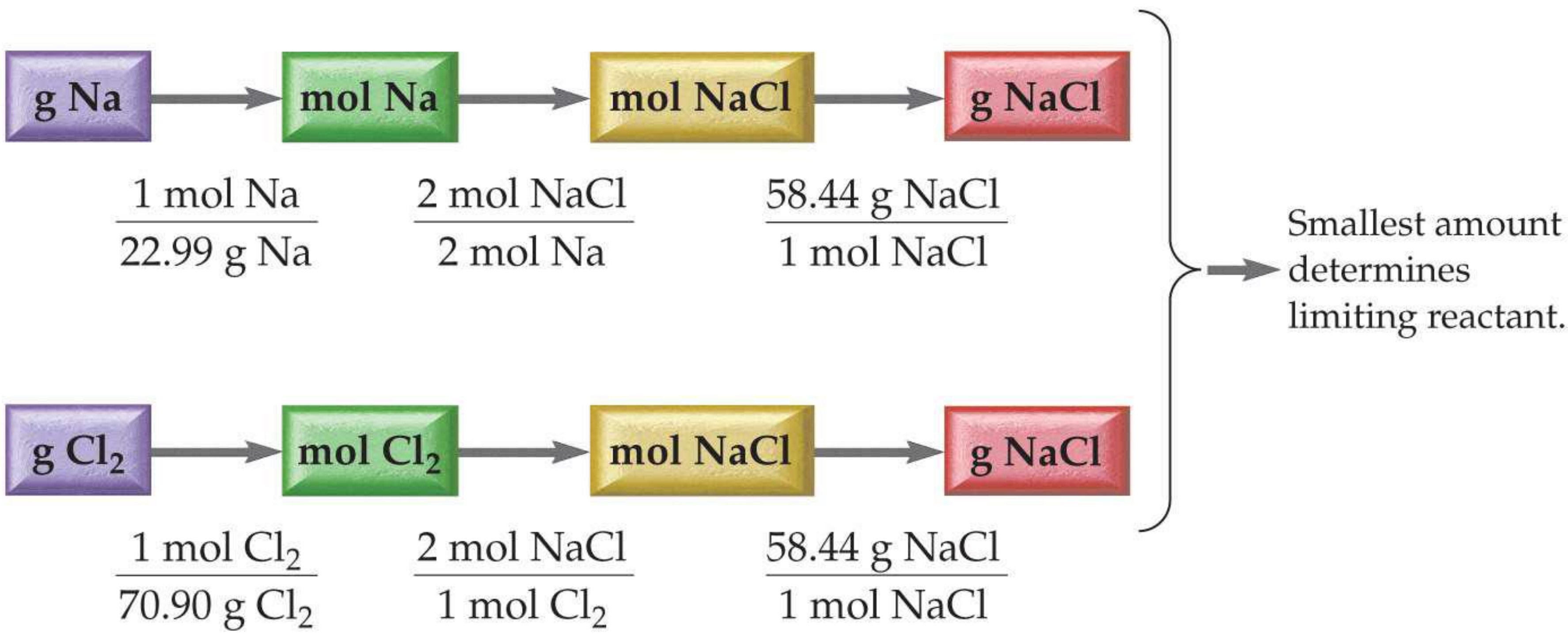
Least amount of product

Limiting Reactant, Theoretical Yield, Actual Yield, and Percent Yield

- In many industrial applications, the more costly reactant or the reactant that is most difficult to remove from the product mixture is chosen to be the limiting reactant.
- When working in the laboratory, we measure the amounts of reactants in ***grams***.
- To find limiting reactants and theoretical yields from initial masses, we must add two steps to our calculations.

Limiting Reactant and Percent Yield: Gram to Gram

Example: 2 Na(s) + Cl_2(g) → 2 NaCl(s)
Given (***grams***): 53.2 g Na and 65.8 g Cl_2
Find: limiting reactant and theoretical yield
SOLUTION MAP:



Limiting Reactant and Percent Yield: Gram to Gram

Example: $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

Given (***grams***): 53.2 g Na and 65.8 g Cl_2

Find: limiting reactant and theoretical yield

SOLUTION:

$$53.2\ \text{g}\ \cancel{\text{Na}} \times \frac{1\ \cancel{\text{mol Na}}}{22.99\ \cancel{\text{g Na}}} \times \frac{2\ \cancel{\text{mol NaCl}}}{2\ \cancel{\text{mol Na}}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = 135\ \text{g NaCl}$$

$$65.8\ \text{g}\ \boxed{\cancel{\text{Cl}_2}} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{2\ \cancel{\text{mol NaCl}}}{1\ \cancel{\text{mol Cl}_2}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = \boxed{108\ \text{g NaCl}}$$

Limiting reactant

Least amount of product

Theoretical Yield and Percent Yield

Example: $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$
Given (***grams***): actual yield 86.4 g NaCl
Find: percent yield

- The actual yield is usually less than the theoretical yield because at least a small amount of product is lost or does not form during a reaction.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \frac{86.4\ \text{g}}{108\ \text{g}} \times 100\% = 80.0\%$$

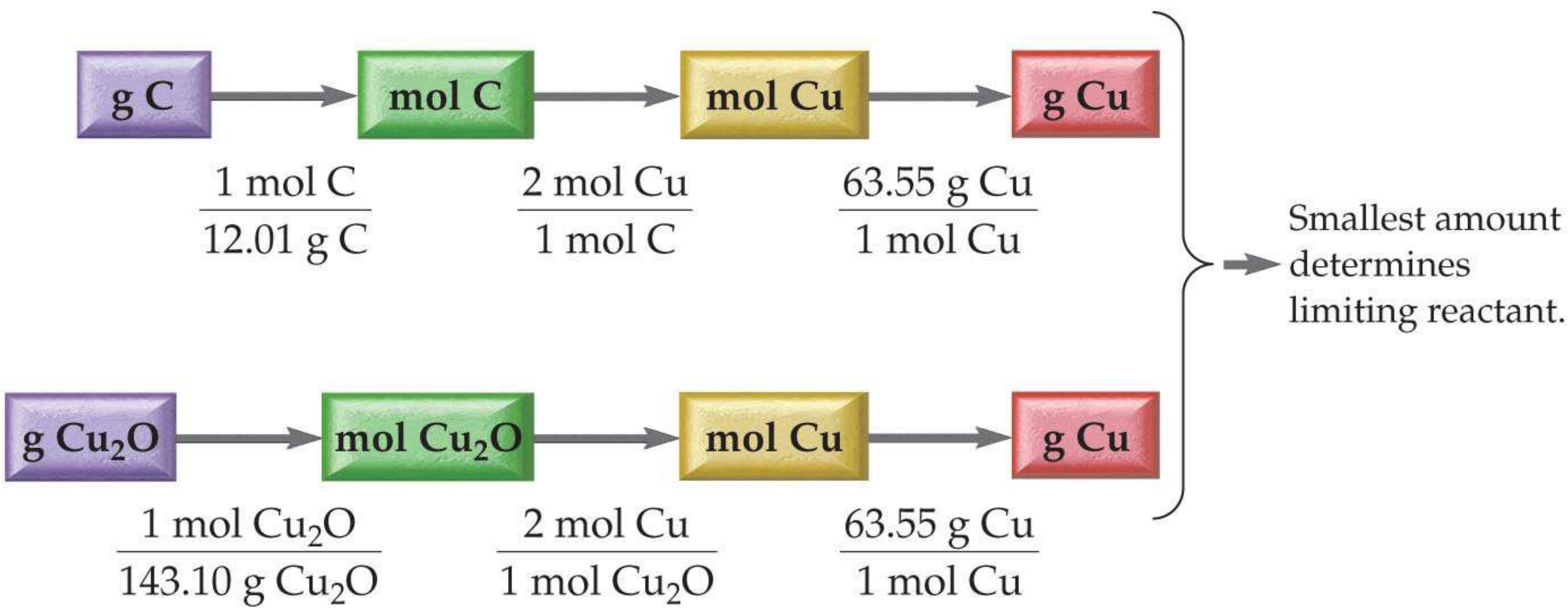
Limiting Reactant and Percent Yield: Gram to Gram

Example 8.6: $Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Given (***grams***): 11.5 g C and 114.5 g Cu_2O

Find: limiting reactant and theoretical yield

SOLUTION MAP:



Relationships Used

- The main conversion factors are the stoichiometric relationships between moles of each reactant and moles of copper.
- The other conversion factors are the molar masses of copper(I) oxide, carbon, and copper.

1 mol Cu_2O : 2 mol Cu

1 mol C : 2 mol Cu

Molar mass Cu_2O = 143.10 g/mol

Molar mass C = 12.01 g/mol

Molar mass Cu = 63.55 g/mol

Limiting Reactant and Percent Yield: Gram to Gram

Example 8.6: $Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Given (***grams***): 11.5 g C and 114.5 g Cu_2O

Find: limiting reactant and theoretical yield

SOLUTION:

$$11.5\ \cancel{\text{g C}} \times \frac{1\ \cancel{\text{mol C}}}{12.01\ \cancel{\text{g C}}} \times \frac{2\ \cancel{\text{mol Cu}}}{1\ \cancel{\text{mol C}}} \times \frac{63.55\ \text{g Cu}}{1\ \cancel{\text{mol Cu}}} = 122\ \text{g Cu}$$

$$\underbrace{114.5\ \text{g}\ \cancel{\text{Cu}_2\text{O}}}_{\text{Limiting reactant}} \times \frac{1\ \cancel{\text{mol Cu}_2\text{O}}}{143.10\ \cancel{\text{g Cu}_2\text{O}}} \times \frac{2\ \cancel{\text{mol Cu}}}{1\ \cancel{\text{mol Cu}_2\text{O}}} \times \frac{63.55\ \text{g Cu}}{1\ \cancel{\text{mol Cu}}} = \underbrace{101.7\ \text{g Cu}}_{\text{Least amount of product}}$$

Actual Yield and Percent Yield

Example 8.6: $Cu_2O(s) + C(s) \rightarrow 2\,Cu(s) + CO(g)$

Given (***grams***): actual yield 87.4 g Cu

Find: percent yield

SOLUTION:

$$\text{Theoretical yield} = 101.7 \text{ g Cu}$$

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

$$= \frac{87.4 \text{ g}}{101.7 \text{ g}} \times 100\% = 85.9\%$$

A shortcut for finding the limiting reactant

- The amount of product calculated for different reactants is affected in exactly the same way by the coefficient of the product in the reaction equation

- What makes a difference is the <u>coefficient of each reactant</u>

$Ti(s) + 2\ Cl_2(g) \rightarrow TiCl_4(s)$
Given 1.8 mol Ti and 3.2 mol Cl_2

$$1.8\ \cancel{\text{mol Ti}} \times \frac{\boxed{\boxed{1}\ \text{mol TiCl}_4}}{1\ \cancel{\text{mol Ti}}} = 1.8\ \text{mol TiCl}_4$$

same for all reactants

$$3.2\ \cancel{\text{mol Cl}_2} \times \frac{\boxed{\boxed{1}\ \text{mol TiCl}_4}}{2\ \cancel{\text{mol Cl}_2}} = 1.6\ \text{mol TiCl}_4$$

Limiting reactant

Least amount of product

Reactant coefficient divides the reactant amount

A shortcut for finding the limiting reactant

Limiting reactant is the one for which the following ratio is the smallest:

$$\frac{(\textbf{moles}\text{ of reactant available})}{(\textbf{coefficient}\text{ in the reaction equation})}$$

Practice:
Check your solution on next page

How many moles of Si_3N_4 can be made from 1.20 moles of Si and 1.00 moles of N_2 in the reaction $3\ Si + 2\ N_2 \rightarrow Si_3N_4$?

Practice — How many moles of Si_3N_4 can be made from 1.20 moles of Si and 1.00 moles of N_2 in the reaction $3\ Si + 2\ N_2 \rightarrow Si_3N_4$?

Given:	1.20 mol Si, 1.00 mol N_2
Find:	mol Si_3N_4
Conceptual Plan:	mol Si $\xrightarrow{\frac{1\ mol\ Si_3N_4}{3\ mol\ Si}}$ mol Si_3N_4; mol N_2 $\xrightarrow{\frac{1\ mol\ Si_3N_4}{2\ mol\ N_2}}$ mol Si_3N_4; Pick least amount → **Limiting reactant and theoretical yield**
Relationships:	2 mol N_2 : 1 Si_3N_4; 3 mol Si : 1 Si_3N_4

Solution:

$$1.20\ \cancel{mol\ Si} \times \frac{1\ mol\ Si_3N_4}{3\ \cancel{mol\ Si}} = 0.400\ mol\ Si_3N_4$$

$$1.00\ \cancel{mol\ N_2} \times \frac{1\ mol\ Si_3N_4}{2\ \cancel{mol\ N_2}} = 0.500\ mol\ Si_3N_4$$

smaller amount

Theoretical yield

Practice:

Check your solution on next page

- When 28.6 kg of C are allowed to react with 88.2 kg of TiO_2 in the reaction below, 42.8 kg of Ti are obtained. Find the limiting reactant, theoretical yield, and percent yield.

$$TiO_2(s) + 2\,C(s) \rightarrow Ti(s) + 2\,CO(g)$$

Practice:
When 28.6 kg of C reacts with 88.2 kg of TiO_2, 42.8 kg of Ti are obtained. Find the limiting reactant, theoretical yield, and percent yield
$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

- Write down the given quantity and its units

Given: 28.6 kg C

88.2 kg TiO_2

42.8 kg Ti produced

Practice:
Find the limiting reactant, theoretical yield, and percent yield
$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information
Given: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti

- Write down the quantity to find and/or its units

  Find: limiting reactant

  theoretical yield

  percent yield

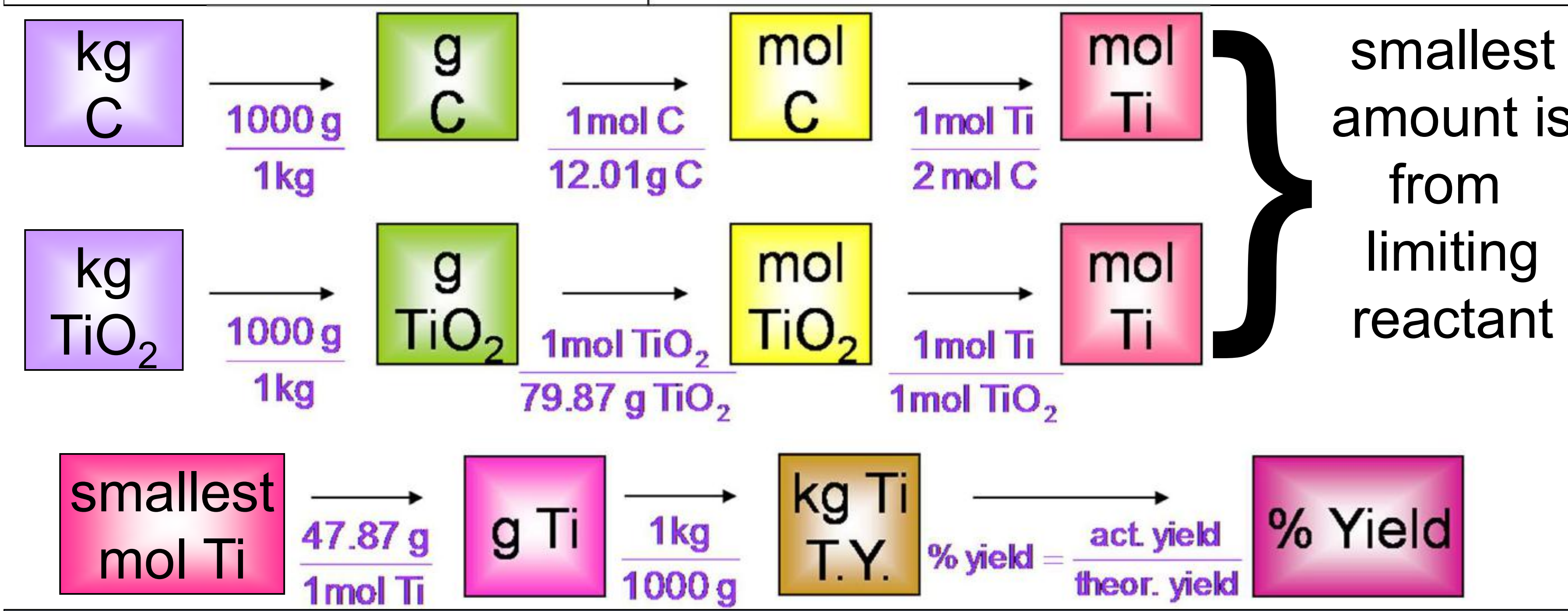
Practice:
Find the limiting reactant, theoretical yield, and percent yield

$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information

Given: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti

Find: lim. rct., theor. yld., % yld.



Or we can directly start with the limiting reactant found via the shortcut

Practice:
Find the limiting reactant, theoretical yield, and percent yield

$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information

Given: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti

Find: lim. rct., theor. yld., % yld.

Plan: kg rct → g rct → mol rct → mol Ti

pick smallest mol Ti → Th.Y kg Ti → %Y Ti

(or start with the lim.rct. given by shortcut)

- Collect needed relationships

1000 g = 1 kg

Molar Mass TiO_2 = 79.87 g/mol

Molar Mass Ti = 47.87 g/mol

Molar Mass C = 12.01 g/mol

1 mole TiO_2 : 1 mol Ti (from the chem. equation)

2 mole C : 1 mol Ti (from the chem. equation)

Practice:
Find the limiting reactant, theoretical yield, and percent yield
$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information
Given: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti
Find: lim. rct., theor. yld., % yld.
CP: kg rct $\rightarrow$ g rct $\rightarrow$ mol rct $\rightarrow$ mol Ti
pick smallest mol Ti $\rightarrow$ TY kg Ti $\rightarrow$ %Y Ti
Rel: 1 mol C=12.01g; 1 mol Ti =47.87g;
1 mol TiO_2 = 79.87g; 1000g = 1 kg;
1 mol TiO_2 : 1 mol Ti; 2 mol C : 1 mol Ti

- Apply the conceptual plan

$$28.6\,\cancel{\text{kg}}\,\text{C} \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{\cancel{1\,\text{mole C}}}{12.01\,\cancel{\text{g}}\,\text{C}} \times \frac{1\,\text{mol Ti}}{\cancel{2\,\text{mol C}}} = 1.1\underline{9}07\times10^{3}\ \text{mol Ti}$$

$$88.2\,\cancel{\text{kg}}\,\text{TiO}_2 \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{\cancel{1\,\text{mole TiO}_2}}{79.87\,\cancel{\text{g}}\,\text{TiO}_2} \times \frac{1\,\text{mol Ti}}{\cancel{1\,\text{mol TiO}_2}} = 1.1\underline{0}43\times10^{3}\ \text{mol Ti}$$

limiting reactant

smallest moles of Ti

Or start with the limiting reactant found via the shortcut

Practice:
Find the limiting reactant, theoretical yield, and percent yield
$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information
Given: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti
Find: lim. rct., theor. yld., % yld.
CP: kg rct → g rct → mol rct → mol Ti
mol Ti from lim.rct.→ TY kg Ti → %Y Ti
Rel: 1 mol C=12.01g; 1 mol Ti =47.87g;
1 mol TiO_2 = 79.87g; 1000g = 1 kg;
1 mol TiO_2 : 1 mol Ti; 2 mol C : 1 mol Ti

- Apply the conceptual plan

$$1.1\underline{0}43 \times 10^3\ \cancel{\text{mol}}\ \text{Ti} \times \frac{47.87\ \cancel{\text{g}}\ \text{Ti}}{1\ \cancel{\text{mol}}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 52.9\ \text{kg Ti}$$

theoretical yield

Practice:
Find the limiting reactant, theoretical yield, and percent yield

$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information

Given: 28.6 kg C, 88.2 kg TiO_2, **42.8 kg Ti**

Find: lim. rct., theor. yld., % yld.

CP: kg rct → g rct → mol rct → mol Ti
pick smallest mol Ti → TY kg Ti → %Y Ti

Rel: 1 mol C=12.01g; 1 mol Ti =47.87g;
1 mol TiO_2 = 79.87g; 1000g = 1 kg;
1 mol TiO_2 : 1 mol Ti; 2 mol C : 1 mol Ti

- Apply the conceptual plan

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

$$\frac{42.8\ \cancel{\text{kg}}\ \text{Ti}}{52.9\ \cancel{\text{kg}}\ \text{Ti}} \times 100\% = 80.9\%$$

Practice:
Find the limiting reactant, theoretical yield, and percent yield
$TiO_2(s) + 2\ C(s) \rightarrow Ti(s) + 2\ CO(g)$

Information
Given: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti
Find: lim. rct., theor. yld., % yld.
CP: kg rct → g rct → mol rct → mol Ti
pick smallest mol Ti → TY kg Ti → %Y Ti
Rel: 1 mol C=12.01g; 1 mol Ti =47.87g;
1 mol TiO_2 = 79.87g; 1000g = 1 kg;
1 mol TiO_2 : 1 mol Ti; 2 mol C : 1 mol Ti

- Check the solutions

limiting reactant = TiO_2
theoretical yield = 52.9 kg
percent yield = 80.9%

Because Ti has lower molar mass than TiO_2, the T.Y. makes sense and the percent yield makes sense as it is less than 100%

Practice:

Check your solution on next page

How many grams of $N_2(g)$ can be made from 9.05 g of NH_3 reacting with 45.2 g of CuO?

$$2\ NH_3(g) + 3\ CuO(s) \rightarrow N_2(g) + 3\ Cu(s) + 3\ H_2O(l)$$

If 4.61 g of N_2 are made, what is the percent yield?

Practice How many grams of $N_2(g)$ can be made from 9.05 g of NH_3 reacting with 45.2 g of CuO? $2\ NH_3(g) + 3\ CuO(s) \rightarrow N_2(g) + 3\ Cu(s) + 3\ H_2O(l)$
If 4.61 g of N_2 are made, what is the percent yield?

Given: 9.05 g NH_3, 45.2 g CuO

Find: g N_2

Conceptual Plan:

$$\text{g } NH_3 \xrightarrow{\frac{1\text{ mol}}{17.03\text{ g}}} \text{mol } NH_3 \xrightarrow{\frac{1\text{ mol } N_2}{2\text{ mol } NH_3}} \text{mol } N_2$$

$$\text{g } CuO \xrightarrow{\frac{1\text{ mol}}{79.55\text{ g}}} \text{mol } CuO \xrightarrow{\frac{1\text{ mol } N_2}{3\text{ mol } CuO}} \text{mol } N_2$$

Choose smallest

$$\text{mol } N_2 \xrightarrow{\frac{28.02\text{ g}}{1\text{ mol}}} \text{g } N_2$$

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

Relationships: 1 mol NH_3 = 17.03g, 1 mol CuO = 79.55g, 1 mol N_2 = 28.02 g
2 mol NH_3 : 1 mol N_2, 3 mol CuO : 1 mol N_2

Practice — How many grams of $N_2(g)$ can be made from 9.05 g of NH_3 reacting with 45.2 g of CuO? $2\ NH_3(g) + 3\ CuO(s) \rightarrow N_2(g) + 3\ Cu(s) + 3\ H_2O(l)$ If 4.61 g of N_2 are made, what is the percent yield?

Solution:

$$9.05\ \cancel{g\ NH_3} \times \frac{1\ \cancel{mol\ NH_3}}{17.03\ \cancel{g\ NH_3}} \times \frac{1\ mol\ N_2}{2\ \cancel{mol\ NH_3}} = 0.2657\ mol\ N_2$$

$$45.2\ \cancel{g\ CuO} \times \frac{1\ \cancel{mol\ CuO}}{79.55\ \cancel{g\ CuO}} \times \frac{1\ mol\ N_2}{3\ \cancel{mol\ CuO}} = 0.1894\ mol\ N_2$$

$$0.1894\ \cancel{mol\ N_2} \times \frac{28.02\ g\ N_2}{1\ \cancel{mol\ N_2}} = \boxed{5.31\ g\ N_2}$$

Theoretical yield

$$\text{Percent Yield} = \frac{4.61\ \cancel{g\ N_2}}{5.31\ \cancel{g\ N_2}} \times 100\% = 86.8\%\ \text{Yield}$$

Check: because the percent yield is less than 100, the answer makes sense

Enthalpy Change: A Measure of the Heat Evolved or Absorbed in a Reaction

- Chemical reactions can be *exothermic* (they *emit* thermal energy when they occur).
- Chemical reactions can be *endothermic* (they *absorb* thermal energy when they occur).
- The *amount* of thermal energy emitted or absorbed by a chemical reaction, under conditions of constant pressure (which are common for most everyday reactions), can be quantified with a function called **enthalpy**.

Enthalpy Change: A Measure of the Heat Evolved or Absorbed in a Reaction

We define the **enthalpy of reaction**, ΔH_{rxn}, as the amount of thermal energy (or heat) that flows when a reaction occurs at constant pressure.

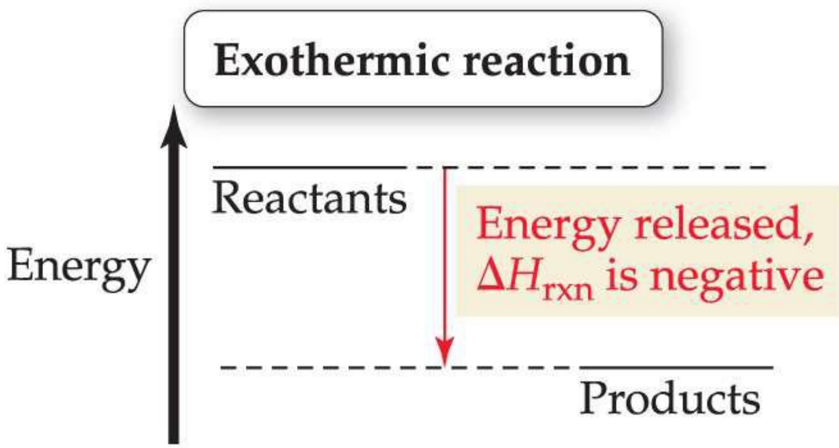
- "Enthalpy of reaction" is actually "<u>change</u> in enthalpy due to reaction"; that's why we have Δ
- Enthalpy is a form of energy content
- Its loss is revealed as released heat
- Its gain is revealed as absorbed heat

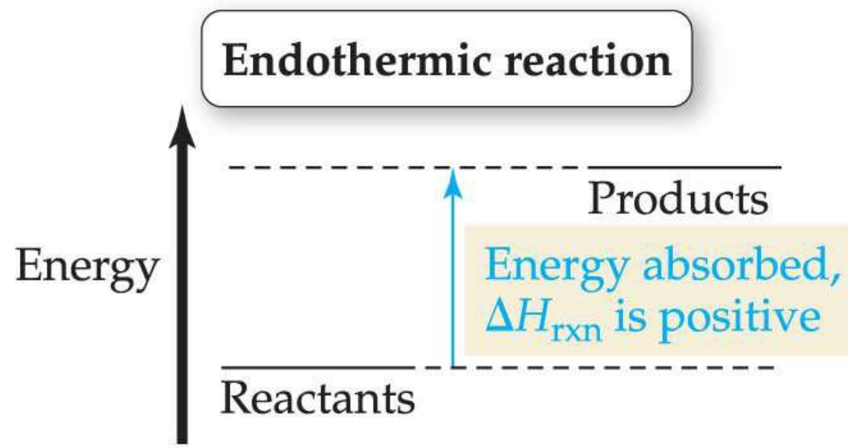
Sign of ΔH_{rxn}

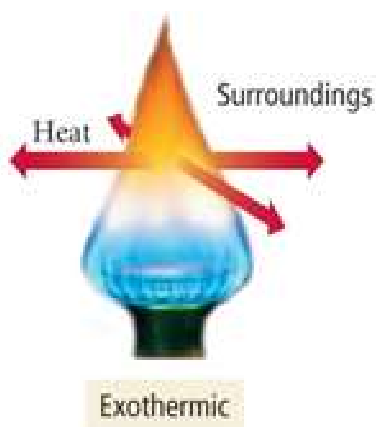
- The *sign* of ΔH_{rxn} (positive or negative) depends on the *direction* in which thermal energy flows when the reaction occurs.
- Energy flowing *out* of the chemical system is like a withdrawal and carries a negative sign.
- Energy flowing *into* the system is like a deposit and carries a positive sign.

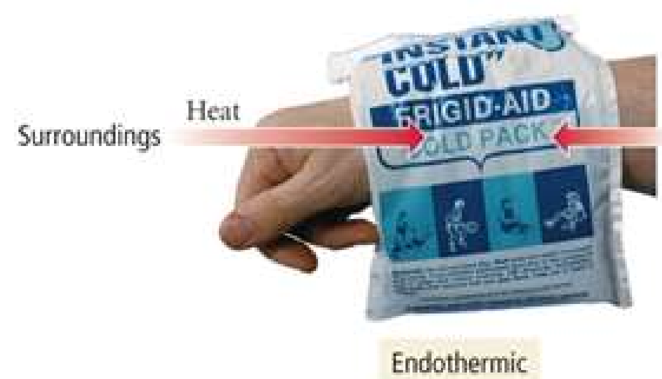
Exothermic and Endothermic Reactions

- In an exothermic reaction, energy is released into the surroundings.
- In an endothermic reaction, energy is absorbed from the surroundings.









Sign of ΔH_{rxn}

When thermal energy flows out of the reaction and into the surroundings (as in an exothermic reaction), then ΔH_{rxn} is negative.

The enthalpy of reaction for the combustion of CH_4, the main component in natural gas, is as follows:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \quad \Delta H_{rxn} = -802.3 \text{ kJ}$$

This reaction is exothermic and therefore has a negative enthalpy of reaction.

The magnitude of ΔH_{rxn} tells us that 802.3 kJ of heat are emitted when 1 mol CH_4 reacts with 2 mol O_2.

Sign of ΔH_{rxn}

When thermal energy flows into the reaction and out of the surroundings (as in an endothermic reaction), then ΔH_{rxn} is positive.

For example, the reaction between nitrogen and oxygen gas to form nitrogen monoxide

$$N_2(g) + O_2(g) \rightarrow 2\ NO(g) \qquad \Delta H_{rxn} = +182.6\ \text{kJ}$$

is endothermic and therefore has a positive enthalpy of reaction.

The magnitude of ΔH_{rxn} tells us that 182.6 kJ of heat is absorbed from the surroundings when 1 mol N_2 reacts with 1 mol O_2.

Stoichiometry of ΔH_{rxn}

- The amount of heat emitted or absorbed when a chemical reaction occurs depends on the *amounts* of reactants that actually react.
- We usually specify ΔH_{rxn} in combination with the balanced chemical equation for the reaction.
- The magnitude of ΔH_{rxn} is for the stoichiometric amounts of reactants and products for the reaction <u>*as written*</u>.
- So, when we specify ΔH_{rxn}, the coefficients (normally indicating <u>ratios</u> only) in the reaction are <u>actual moles</u> corresponding to <u>that</u> value of ΔH_{rxn}
  -- they still indicate stoichiometric ratios, of course

Stoichiometry of ΔH_{rxn}

For example, the balanced equation and ΔH_{rxn} for the combustion of propane (the fuel used in LP gas) is as follows:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H_{rxn} = -2044 \text{ kJ}$$

When 1 mole of C_3H_8 reacts with 5 moles of O_2 to form 3 moles of CO_2 and 4 moles of H_2O, 2044 kJ of heat is emitted.

These ratios can be used to construct conversion factors between amounts of reactants or products and the quantity of heat exchanged.

Stoichiometry of ΔH_{rxn}

- To find out how much heat is emitted upon the combustion of a certain mass in grams of propane C_3H_8, we can use the following solution map:

$$\text{g } C_3H_8 \xrightarrow{\frac{1 \text{ mol } C_3H_8}{44.11 \text{ g } C_3H_8}} \text{mol } C_3H_8 \xrightarrow{\frac{-2044 \text{ kJ}}{1 \text{ mol } C_3H_8}} \text{kJ}$$

Example 8.7: Stoichiometry Involving ΔH_{rxn}

- An LP gas tank in a home barbecue contains 11.8×10^3 g of propane (C_3H_8).
- Calculate the heat (in kJ) associated with the complete combustion of all of the propane in the tank.

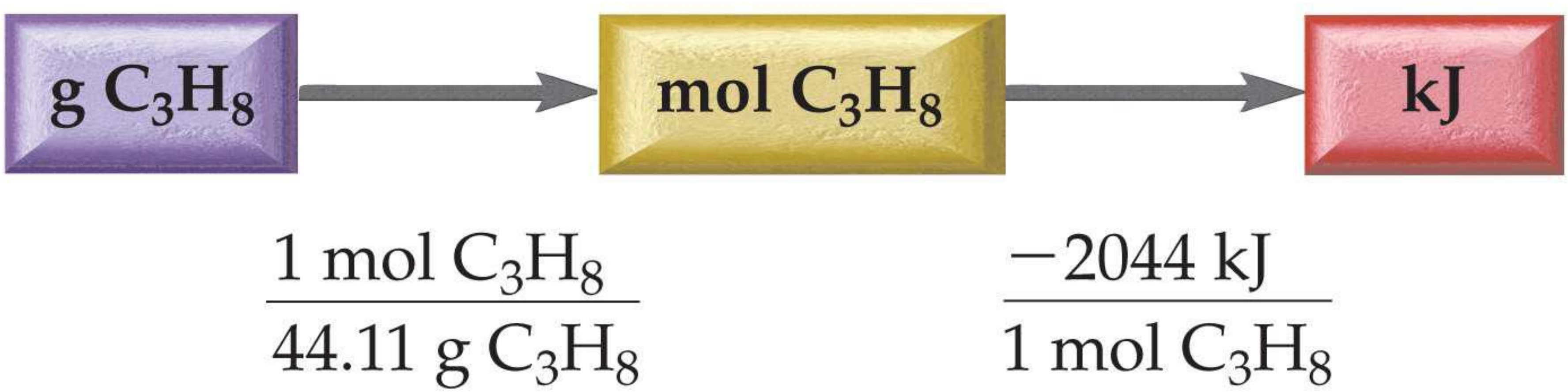
$$C_3H_8(g) + 5\,O_2(g) \rightarrow 3\,CO_2(g) + 4\,H_2O(g)$$

$$\Delta H_{rxn} = -2044\text{ kJ}$$

Stoichiometry Involving ΔH_{rxn}

Example: Complete combustion of 11.8×10^3 g of propane (C_3H_8)

SOLUTION MAP:



RELATIONSHIPS USED:

1 mol C_3H_8 : –2044 kJ (from balanced equation)

Molar mass C_3H_8 = 44.11 g/mol

Stoichiometry Involving ΔH_{rxn}

Example: Complete combustion of 11.8×10^3 g of propane (C_3H_8)

SOLUTION:

$$11.8 \times 10^3\ \cancel{\mathrm{g\,C_3H_8}} \times \frac{1\ \cancel{\mathrm{mol\,C_3H_8}}}{44.11\ \cancel{\mathrm{g\,C_3H_8}}} \times \frac{-2044\ \mathrm{kJ}}{1\ \cancel{\mathrm{mol\,C_3H_8}}} = -5.47 \times 10^5\ \mathrm{kJ}$$

Often in the questions, the absolute value of heat (q), $|q|$, is requested and words are used to convey the sign of the heat absorbed or given off in the reaction.

q = "heat" = heat absorbed

heat absorbed = q $\qquad$ $q = \Delta H$

heat released = $-q$ $\qquad$ $-q = -\Delta H$

Practice:
Check your solution on next page

How much heat is evolved (released) when a 0.483 g diamond (a form of carbon) is burned?
($\Delta H_{combustion}$ = −395.4 kJ/mol C)

Practice – How much heat is evolved when a 0.483 g diamond is burned?

Given:	0.483 g C, ΔH = −395.4 kJ/mol C
Find:	q, kJ/mol
Concept Plan:	g → mol → kJ $\frac{1\,\text{mol C}}{12.01\,\text{g}}$ $\frac{-395.4\,\text{kJ}}{1\,\text{mol C}}$ 1 mol C = −395.4 kJ, Molar Mass = 12.01 g/mol
Solution:	$0.483\,\text{g}\,\cancel{\text{C}} \times \frac{1\,\cancel{\text{mol C}}}{12.01\,\text{g}\,\cancel{\text{C}}} \times \frac{-395.4\,\text{kJ}}{1\,\cancel{\text{mol C}}} = -15.9\,\text{kJ}$
Check:	the sign is correct and the number is reasonable because the amount of diamond is less than 1 mole

One more thing:

- If we consider the reverse of a reaction, the ΔH_{rxn} sign is reversed (positive becomes negative, negative becomes positive)

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g) \quad \Delta H_{rxn} = -2044$ kJ

$3\ CO_2(g) + 4\ H_2O(g) \rightarrow C_3H_8(g) + 5\ O_2(g) \quad \Delta H_{rxn} = +2044$ kJ

We would have to <u>supply</u> 2044 kJ of energy to combine 3 moles of $CO_2(g)$ and 4 moles of H_2O to obtain 1 mole of $C_3H_8(g)$ and 5 moles of $O_2(g)$

Everyday Chemistry Bunsen Burners

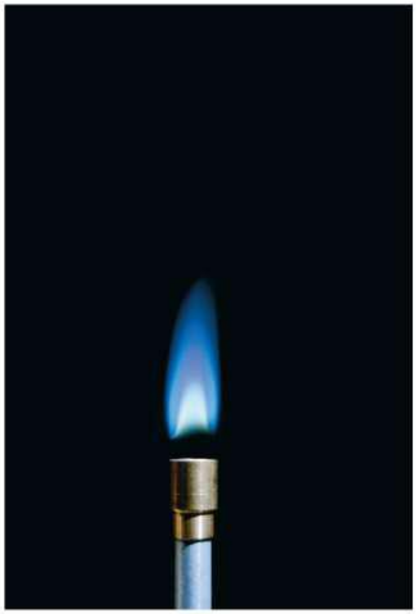
- Most Bunsen burners have a mechanism to adjust the amount of air (and therefore of oxygen) that is mixed with the methane.
- If you light the burner with the air completely closed off, you get a yellow, smoky flame that is not very hot.
- As you increase the amount of air going into the burner, the flame becomes bluer, less smoky, and hotter.
- When you reach the optimum adjustment, the flame has a sharp, inner blue triangle, gives off no smoke, and is hot enough to melt glass easily.
- Continuing to increase the air beyond this point causes the flame to become cooler again and may actually extinguish it.

A Bunsen Burner at Various Stages of Air Intake Adjustment

(a) No air

(b) Small amount of air

(c) Optimum



(d) Too much air

Chapter 8 in Review

- **Stoichiometry:** A balanced chemical equation gives quantitative relationships between the amounts of reactants and products. The quantitative relationship between reactants and products in a chemical reaction is called reaction stoichiometry.

Chapter 8 in Review

- **Limiting Reactant, Theoretical Yield, and Percent Yield:**
- The limiting reactant in a chemical reaction is the reactant that limits the amount of product that can be made.
- The theoretical yield in a chemical reaction is the amount of product that can be made based on the amount of the limiting reactant.
- The actual yield in a chemical reaction is the amount of product actually produced.
- The percent yield in a chemical reaction is the actual yield divided by theoretical yield times 100%.

- **Enthalpy of Reaction:** The amount of heat released or absorbed by a chemical reaction under conditions of constant pressure is the enthalpy of reaction (ΔH_{rxn}).
- The magnitude of ΔH_{rxn} is associated with the stoichiometric amounts of reactants and products for the reaction *as written*.

Chemical Skills Learning Objectives

1. LO: Recognize the numerical relationship between chemical quantities in a balanced chemical equation.
2. LO: Carry out mole-to-mole conversions between reactants and products based on the numerical relationship between chemical quantities in a balanced chemical equation.
3. LO: Carry out mass-to-mass conversions between reactants and products based on the numerical relationship between chemical quantities in a balanced chemical equation and molar masses.
4. LO: Calculate limiting reactant, theoretical yield, and percent yield for a given amount of reactants in a balanced chemical equation.
5. LO: Calculate the amount of thermal energy emitted or absorbed by a chemical reaction.

Highlight Problem EOC 8.101

- Scientists have grown progressively more worried about the potential for global warming caused by increasing atmospheric carbon dioxide levels.
- The world burns the fossil fuel equivalent of approximately 9.0×10^{12} kg of petroleum per year.
- Assume that all of this petroleum is in the form of octane (C_8H_{18}) and calculate how much CO_2 in kilograms is produced by world fossil fuel combustion per year, and use the given rate of consumption to solve the question on the next page.

Highlight Problem EOC 8.101

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

- The balanced chemical equation shows that 16 mol of CO_2 are produced for every 2 mol of octane burned.
- If the atmosphere currently contains approximately 3.0×10^{15} kg of CO_2, how long will it take for the world's fossil fuel combustion to double the amount of atmospheric carbon dioxide?