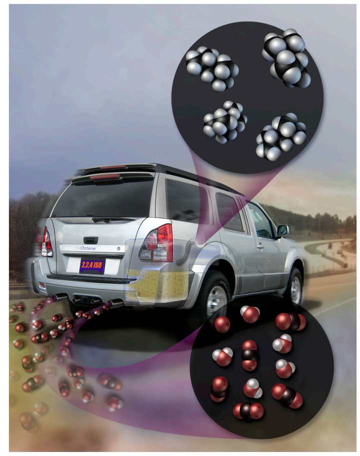
Introductory Chemistry, 3rd Edition
Nivaldo Tro

Chapter 8 Quantities in Chemical Reactions



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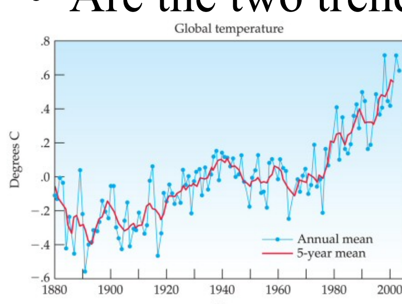
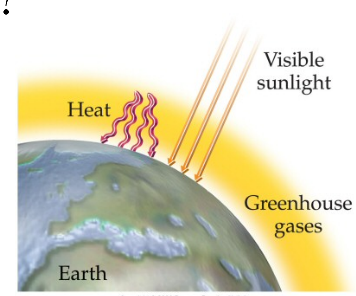
2009, Prentice Hall

Global Warming

- Scientists have measured an average 0.6 °C rise in atmospheric temperature since 1860.
- During the same period atmospheric CO_2 levels have risen 25%.
- Are the two trends causal?

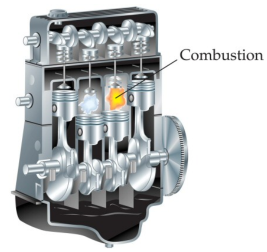




The Source of Increased CO_2

- The primary source of the increased CO_2 levels are combustion reactions of fossil fuels we use to get energy.
  - ✓1860 corresponds to the beginning of the Industrial Revolution in the U.S. and Europe.

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$ $2\,C_8H_{18}(l) + 25\,O_2(g) \rightarrow 16\,CO_2(g) + 18\,H_2O(g)$

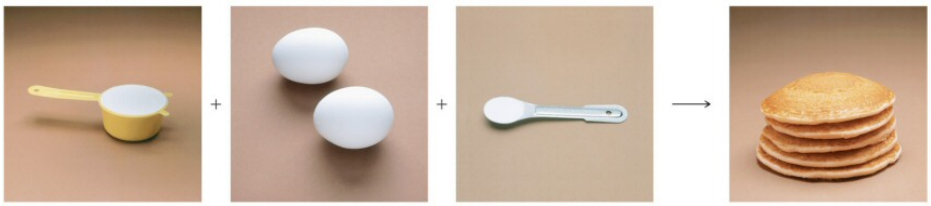


Quantities in Chemical Reactions

- The amount of every substance used and made in a chemical reaction is related to the amounts of all the other substances in the reaction.
  - ✓Law of Conservation of Mass.
  - ✓Balancing equations by balancing atoms.
- The study of the numerical relationship between chemical quantities in a chemical reaction is called **stoichiometry**.

Making Pancakes

- The number of pancakes you can make depends on the amount of the ingredients you use.



1 cup flour + 2 eggs + ½ tsp baking powder $\rightarrow$ 5 pancakes

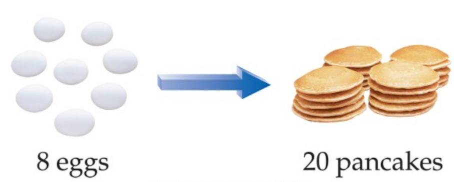
- This relationship can be expressed mathematically.

1 cup flour $\equiv$ 2 eggs $\equiv$ ½ tsp baking powder $\equiv$ 5 pancakes

Making Pancakes, Continued

- If you want to make more or less than 5 pancakes, you can use the number of eggs you have to determine the number of pancakes you can make.
  - ✓ Assuming you have enough flour and baking powder.

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



Making Molecules
Mole-to-Mole Conversions

- The balanced equation is the "recipe" for a chemical reaction.
- The equation $3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$ tells us that 3 molecules of H_2 react with exactly 1 molecule of N_2 and make exactly 2 molecules of NH_3 or:

  3 molecules $H_2 \equiv$ 1 molecule $N_2 \equiv$ 2 molecules NH_3
- Since we count molecules by moles:

$$3 \text{ moles } H_2 \equiv 1 \text{ mole } N_2 \equiv 2 \text{ moles } NH_3$$

Example 8.1—How Many Moles of NaCl Result from the Complete Reaction of 3.4 Mol of Cl_2? $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl$

Given: **Find:**	3.4 mol Cl_2 mol NaCl
Solution Map: **Relationships:**	$\text{mol } Cl_2 \xrightarrow{\frac{2 \text{ mol NaCl}}{1 \text{ mol } Cl_2}} \text{mol NaCl}$ 1 mol $Cl_2 \equiv$ 2 NaCl
Solution:	$3.4 \text{ mol } \cancel{Cl_2} \times \dfrac{2 \text{ mol NaCl}}{1 \text{ mol } \cancel{Cl_2}}$ $= 6.8 \text{ mol NaCl}$
Check:	Since the reaction makes 2 molecules of NaCl for every 1 mole of Cl_2, the number makes sense.

Example 8.1:

- Sodium chloride, NaCl, forms by the following reaction between sodium and chlorine. How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2? Assume there is more than enough Na.

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

Example:
How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2 in the reaction below?
$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

- Write down the given quantity and its units.

Given: 3.4 mol Cl_2

Example: How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2 in the reaction below? $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$	Information: Given: 3.4 mol Cl_2

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

  Find: ? moles NaCl

Example: How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2 in the reaction below? $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$	Information: Given: 3.4 mol Cl_2 Find: ? moles NaCl

- Collect needed conversion factors:

According to the equation:

1 mole $Cl_2 \equiv$ 2 moles NaCl

Example:
How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2 in the reaction below?
$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

Information:
Given: 3.4 mol Cl_2
Find: ? moles NaCl
Conversion Factor:
1 mol Cl_2 ≡ 2 mol NaCl

- Write a solution map for converting the units:

$$\text{mol Cl}_2 \xrightarrow[\frac{2\ \text{mol NaCl}}{1\ \text{mol Cl}_2}]{} \text{mol NaCl}$$

Example:
How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2 in the reaction below?
$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

Information:
Given: 3.4 mol Cl_2
Find: ? moles NaCl
Conversion Factor:
1 mol Cl_2 ≡ 2 mol NaCl
Solution Map: mol Cl_2 → mol NaCl

- Apply the solution map:

$$3.4\ \cancel{\text{mol}}\,\text{Cl}_2 \times \frac{2\ \text{mol NaCl}}{1\ \cancel{\text{mol}}\,\text{Cl}_2} = \text{moles NaCl}$$

$$= 6.8\ \text{mol NaCl}$$

- Significant figures and round:

$$= 6.8\ \text{moles NaCl}$$

Example: How many moles of NaCl result from the complete reaction of 3.4 mol of Cl_2 in the reaction below? $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$	Information: Given: 3.4 mol Cl_2 Find: ? moles NaCl Conversion Factor: 1 mol $Cl_2 \equiv$ 2 mol NaCl Solution Map: mol $Cl_2 \rightarrow$ mol NaCl

- Check the solution:

$$3.4 \text{ mol } Cl_2 \equiv 6.8 \text{ mol NaCl}$$

The units of the answer, moles NaCl, are correct.
The magnitude of the answer makes sense because the equation tells us you make twice as many moles of NaCl as the moles of Cl_2 .

Practice

- According to the following equation, how many moles of water are made in the combustion of 0.10 moles of glucose?

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

How Many Moles of Water Are Made in the Combustion of 0.10 Moles of Glucose?

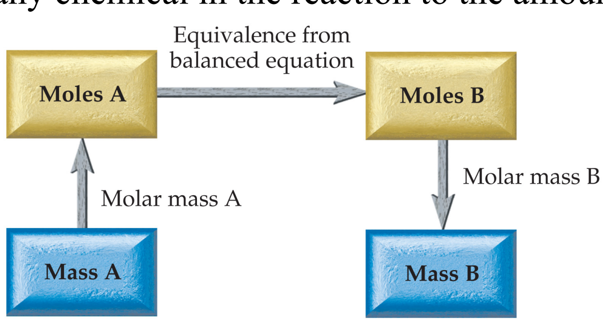
Given: **Find:**	0.10 moles $C_6H_{12}O_6$ moles H_2O
Solution Map:	$\text{mol } C_6H_{12}O_6 \xrightarrow{\frac{6 \text{ mol } H_2O}{1 \text{ mol } C_6H_{12}O_6}} \text{mol } H_2O$
Relationships:	$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O \quad \therefore\ 1 \text{ mol } C_6H_{12}O_6 \equiv 6 \text{ mol } H_2O$
Solution:	$0.10\ \cancel{\text{mol } C_6H_{12}O_6} \times \frac{6 \text{ mol } H_2O}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 0.6 \text{ mol } H_2O$
Check:	**0.6 mol H_2O = 0.60 mol H_2O** Since 6x moles of H_2O as $C_6H_{12}O_6$, the number makes sense.

Making Molecules Mass-to-Mass Conversions

- We know there is a relationship between the mass and number of moles of a chemical.

  1 mole = Molar Mass in grams.

- The molar mass of the chemicals in the reaction and the balanced chemical equation allow us to convert from the amount of any chemical in the reaction to the amount of any other.



Example 8.2—How Many Grams of Glucose Can Be Synthesized from 58.5 g of CO_2 in Photosynthesis?

- Photosynthesis:

$$6\ CO_2(g) + 6\ H_2O(g) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$$

- The equation for the reaction gives the mole relationship between amount of $C_6H_{12}O_6$ and CO_2, but we need to know the mass relationship, so the solution map will be:

g CO_2 → mol CO_2 → mol $C_6H_{12}O_6$ → g $C_6H_{12}O_6$

Example 8.2—How Many Grams of Glucose Can Be Synthesized from 58.5 g of CO_2 in Photosynthesis?, Continued

Given:	58.5 g CO_2
Find:	g $C_6H_{12}O_6$
Solution Map:	g CO_2 $\xrightarrow{\frac{1\ \text{mol}}{44.01\ \text{g}}}$ mol CO_2 $\xrightarrow{\frac{1\ \text{mol}\ C_6H_{12}O_6}{6\ \text{mol}\ CO_2}}$ mol $C_6H_{12}O_6$ $\xrightarrow{\frac{180.2\ \text{g}}{1\ \text{mol}}}$ g $C_6H_{12}O_6$
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:	$58.5\ \cancel{g\ CO_2} \times \frac{1\ \cancel{mol\ CO_2}}{44.01\ \cancel{g\ CO_2}} \times \frac{1\ \cancel{mol\ C_6H_{12}O_6}}{6\ \cancel{mol\ CO_2}} \times \frac{180.2\ g\ C_6H_{12}O_6}{1\ \cancel{mol\ C_6H_{12}O_6}}$ $= 39.9\ g\ C_6H_{12}O_6$
Check:	Since 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.

Example 8.2:

- In photosynthesis, plants convert carbon dioxide and water into glucose ($C_6H_{12}O_6$), according to the following reaction. How many grams of glucose can be synthesized from 58.5 g of CO_2? Assume there is more than enough water to react with all the CO_2.

$$6\,CO_2(g) + 6\,H_2O(l) \xrightarrow{\text{sunlight}} 6\,O_2(g) + C_6H_{12}O_6(aq)$$

Example:
How many grams of glucose can be synthesized from 58.5 g of CO_2 in the reaction?
$6\,CO_2(g) + 6\,H_2O(l) \rightarrow$
$6\,O_2(g) + C_6H_{12}O_6(aq)$

- Write down the given quantity and its units.

Given: 58.5 g CO_2

Example: How many grams of glucose can be synthesized from 58.5 g of CO_2 in the reaction? $6\ CO_2(g) + 6\ H_2O(l) \rightarrow 6\ O_2(g) + C_6H_{12}O_6(aq)$	Information: Given: 55.4 g CO_2

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

  Find: ? g $C_6H_{12}O_6$

Example: How many grams of glucose can be synthesized from 58.5 g of CO_2 in the reaction? $6\ CO_2(g) + 6\ H_2O(l) \rightarrow 6\ O_2(g) + C_6H_{12}O_6(aq)$	Information: Given: 55.4 g CO_2 Find: g $C_6H_{12}O_6$

- Collect needed conversion factors:

  Molar mass $C_6H_{12}O_6$ = 6(mass C) + 12(mass H) + 6(mass O)
  = 6(12.01) + 12(1.01) + 6(16.00) = 180.2 g/mol

  Molar mass CO_2 = 1(mass C) + 2(mass O)
  = 1(12.01) + 2(16.00) = 44.01 g/mol

  1 mole CO_2 = 44.01 g CO_2

  1 mole $C_6H_{12}O_6$ = 180.2 g $C_6H_{12}O_6$

  1 mole $C_6H_{12}O_6$ ≡ 6 mol CO_2 (from the chem. equation)

Example:
How many grams of glucose can be synthesized from 58.5 g of CO_2 in the reaction?
$6\ CO_2(g) + 6\ H_2O(l) \rightarrow 6\ O_2(g) + C_6H_{12}O_6(aq)$

Information:
Given: 58.5 g CO_2
Find: g $C_6H_{12}O_6$
Conversion Factors:
1 mol $C_6H_{12}O_6$ = 180.2 g
1 mol CO_2 = 44.01 g
1 mol $C_6H_{12}O_6 \equiv$ 6 mol CO_2

- Write a solution map:

$$\text{g } CO_2 \xrightarrow{\frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2}} \text{mol } CO_2 \xrightarrow{\frac{1\text{ mol } C_6H_{12}O_6}{6\text{ mol } CO_2}} \text{mol } C_6H_{12}O_6 \xrightarrow{\frac{180.2\text{ g } C_6H_{12}O_6}{1\text{ mol } C_6H_{12}O_6}} \text{g } C_6H_{12}O_6$$

Example:
How many grams of glucose can be synthesized from 58.5 g of CO_2 in the reaction?
$6\ CO_2(g) + 6\ H_2O(l) \rightarrow 6\ O_2(g) + C_6H_{12}O_6(aq)$

Information:
Given: 58.5 g CO_2
Find: g $C_6H_{12}O_6$
Conversion Factors:
1 mol $C_6H_{12}O_6$ = 180.2 g
1 mol CO_2 = 44.01 g
1 mol $C_6H_{12}O_6 \equiv$ 6 mol CO_2
Solution Map: g $CO_2 \rightarrow$ mol $CO_2 \rightarrow$ mol $C_6H_{12}O_6 \rightarrow$ g $C_6H_{12}O_6$

- Apply the solution map:

$$58.5\ \cancel{\text{g } CO_2} \times \frac{1\ \cancel{\text{mole}}\ CO_2}{44.01\ \cancel{\text{g } CO_2}} \times \frac{1\text{ mol } C_6\cancel{H_{12}O_6}}{6\text{ mol } \cancel{CO_2}} \times \frac{180.2\text{ g } C_6H_{12}O_6}{1\ \cancel{\text{mole } C_6H_{12}O_6}}$$

$$= 39.9216\text{ g } C_6H_{12}O_6$$

- Significant figures and round:

$$= 39.9\text{ g } C_6H_{12}O_6$$

Example:
How many grams of glucose can be synthesized from 58.5 g of CO_2 in the reaction?
$6\ CO_2(g) + 6\ H_2O(l) \rightarrow$
$6\ O_2(g) + C_6H_{12}O_6(aq)$

Information:
Given: 58.5 g CO_2
Find: g $C_6H_{12}O_6$
Conversion Factors:
1 mol $C_6H_{12}O_6$ = 180.2 g
1 mol CO_2 = 44.01 g
1 mol $C_6H_{12}O_6 \equiv$ 6 mol CO_2
Solution Map: g $CO_2 \rightarrow$ mol $CO_2 \rightarrow$ mol $C_6H_{12}O_6 \rightarrow$ g $C_6H_{12}O_6$

- Check the solution:

$$58.5\ g\ CO_2 = 39.9\ g\ C_6H_{12}O_6$$

The units of the answer, g $C_6H_{12}O_6$, are correct.
It is hard to judge the magnitude.

Practice—How Many Grams of O_2 Can Be Made from the Decomposition of 100.0 g of PbO_2?

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

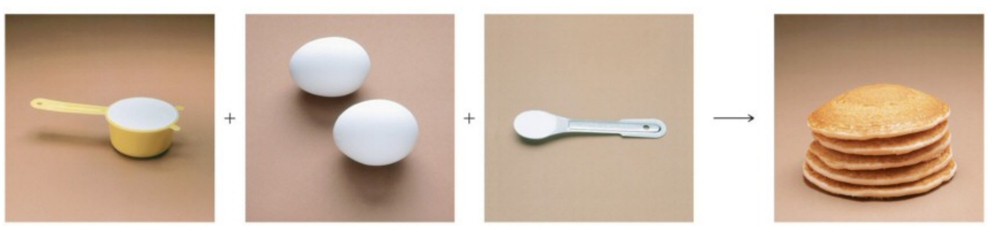
$(PbO_2 = 239.2,\ O_2 = 32.00)$

Practice—How Many Grams of O_2 Can Be Made from the Decomposition of 100.0 g of PbO_2? $2\ PbO_2(s) \rightarrow 2\ PbO(s) + O_2(g)$, Continued

Given:	100.0 g PbO_2, $2\ PbO_2 \rightarrow 2\ PbO + O_2$
Find:	g O_2
Solution Map:	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:	Since ½ moles of O_2 as PbO_2, and the molar mass of PbO_2 is 7x O_2, the number makes sense.

More Making Pancakes

- We know that:



1 cup flour + 2 eggs + ½ tsp baking powder → 5 pancakes

- But what would happen if we had 3 cups of flour, 10 eggs, and 4 tsp of baking powder?

More Making Pancakes, Continued

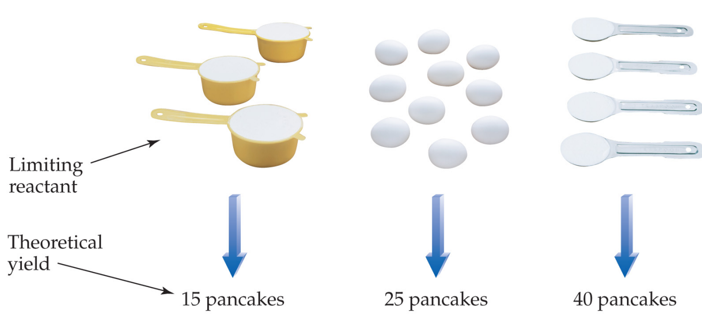
$$3 \text{ cups flour} \times \frac{5 \text{ pancakes}}{1 \text{ cups flour}} = 15 \text{ pancakes}$$

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

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

More Making Pancakes, Continued

- Each ingredient could potentially make a different number of pancakes.
- But all the ingredients have to work together!
- We only have enough flour to make 15 pancakes, so once we make 15 pancakes, the flour runs out no matter how much of the other ingredients we have.



More Making Pancakes, Continued

- The flour limits the amount of pancakes we can make. In chemical reactions we call this the **limiting reactant**.
  - ✓ Also known as limiting reagent.
- The maximum number of pancakes we can make depends on this ingredient. In chemical reactions, we call this the **theoretical yield**.
  - ✓ It also determines the amounts of the other ingredients we will use!

Example 8.4—What Is the Limiting Reactant and Theoretical Yield When 0.552 Mol of Al React with 0.887 Mol of Cl_2?

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

Given: 0.552 mol Al, 0.887 mol Cl_2

Find: mol $AlCl_3$

Solution Map:

$$\text{mol Al} \xrightarrow{\frac{2\ \text{mol}\ AlCl_3}{2\ \text{mol Al}}} \text{mol}\ AlCl_3$$

$$\text{mol}\ Cl_2 \xrightarrow{\frac{2\ \text{mol}\ AlCl_3}{3\ \text{mol}\ Cl_2}} \text{mol}\ AlCl_3$$

Pick least amount → Limiting reactant and theoretical yield

Relationships: 3 mol $Cl_2 \equiv 2\ AlCl_3$; 2 mol Al $\equiv$ 2 mol $AlCl_3$

Solution:

$$0.552\ \cancel{\text{mol}}\ \text{Al} \times \frac{2\ \text{mol}\ AlCl_3}{2\ \cancel{\text{mol}}\ \text{Al}} = 0.552\ \text{mol}\ AlCl_3$$

Limiting Reactant

$$0.877\ \cancel{\text{mol}}\ Cl_2 \times \frac{2\ \text{mol}\ AlCl_3}{3\ \cancel{\text{mol}}\ Cl_2} = 0.5847\ \text{mol}\ AlCl_3$$

Theoretical Yield

Example 8.4:

- What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2?
  $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$

Example:
What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2?
$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$

- Write down the given quantity and its units.

Given: 0.552 mol Al

0.877 mol Cl_2

Example: What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2? $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$	Information: Given: 0.552 mol Al, 0.877 mol Cl_2

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

Find: limiting reactant
theoretical yield

Example: What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2? $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$	Information: Given: 0.552 mol Al, 0.877 mol Cl_2 Find: limiting reactant, theor. yield

- Collect needed conversion factors:

$2\ \text{mol}\ AlCl_3 \equiv 2\ \text{mol Al}$ (from the chem. equation)

$2\ \text{mol}\ AlCl_3 \equiv 3\ \text{mol}\ Cl_2$ (from the chem. equation)

Example:
What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2?
$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$

Information:
Given: 0.552 mol Al, 0.877 mol Cl_2
Find: limiting reactant, theor. yield
Conversion Factors:
$2\text{ mol }AlCl_3 \equiv 2\text{ mol Al}$,
$2\text{ mol }AlCl_3 \equiv 3\text{ mol }Cl_2$

- Write a solution map:

$$\left.\begin{array}{l} \text{mol Al} \xrightarrow{\frac{2\text{ mol AlCl}_3}{2\text{ mol Al}}} \text{mol AlCl}_3 \\ \\ \text{mol Cl}_2 \xrightarrow{\frac{2\text{ mol AlCl}_3}{3\text{ mol Cl}_2}} \text{mol AlCl}_3 \end{array}\right\}$$

Smallest amount is theoretical yield

Example:
What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2?
$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$

Information:
Given: 0.552 mol Al, 0.877 mol Cl_2
Find: limiting reactant, theor. yield
Conversion Factors:
$2\text{ mol }AlCl_3 \equiv 2\text{ mol Al}$,
$2\text{ mol }AlCl_3 \equiv 3\text{ mol }Cl_2$
Solution Map:
mol each reactant $\rightarrow$ mol $AlCl_3$

- Apply the solution map:

Limiting reactant = Al

$$0.552\ \cancel{\text{mol Al}} \times \frac{2\text{ mol AlCl}_3}{2\ \cancel{\text{mol Al}}} = 0.552\text{ mol AlCl}_3$$

Smallest amount

$$0.877\ \cancel{\text{mol Cl}_2} \times \frac{2\text{ mol AlCl}_3}{3\ \cancel{\text{mol Cl}_2}} = 0.58\underline{4}7\text{ mol AlCl}_3$$

Theoretical yield = 0.552 mol $AlCl_3$

Example:
What is the limiting reactant and theoretical yield when 0.552 mol of Al react with 0.887 mol of Cl_2?
$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3$

Information:
Given: 0.552 mol Al, 0.877 mol Cl_2
Find: limiting reactant, theor. yield
Conversion Factors:
2 mol $AlCl_3 \equiv$ 2 mol Al,
2 mol $AlCl_3 \equiv$ 3 mol Cl_2
Solution Map:
mol each reactant → mol $AlCl_3$

- Check the solution:

Limiting reactant = Al
Theoretical yield = 0.552 mol $AlCl_3$

Usually hard to judge as there are multiple factors, but because Al resulted in smallest amount of $AlCl_3$, the answer makes sense.

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 $3\ Si + 2\ N_2 \rightarrow Si_3N_4$?

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 $3\ Si + 2\ N_2 \rightarrow Si_3N_4$?, Continued

Given:	1.20 mol Si, 1.00 mol N_2
Find:	mol Si_3N_4
Solution Map:	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:	$1.20\ \cancel{mol}\ Si \times \frac{1\ mol\ Si_3N_4}{3\ mol\ \cancel{Si}} = 0.400\ mol\ Si_3N_4$ Limiting Reactant $1.00\ \cancel{mol}\ N_2 \times \frac{1\ mol\ Si_3N_4}{2\ mol\ \cancel{N_2}} = 0.500\ mol\ Si_3N_4$ Theoretical Yield

More Making Pancakes

- Let's now assume that as we are making pancakes, we spill some of the batter, burn a pancake, drop one on the floor, or other uncontrollable events happen so that we only make 11 pancakes. The actual amount of product made in a chemical reaction is called the **actual yield**.
- We can determine the efficiency of making pancakes by calculating the percentage of the maximum number of pancakes we actually make. In chemical reactions, we call this the **percent yield**.

$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \text{Percent Yield} \qquad \frac{11\ \text{pancakes}}{15\ \text{pancakes}} \times 100\% = 73\%$$

Theoretical and Actual Yield

- As we did with the pancakes, in order to determine the theoretical yield, we should use reaction stoichiometry to determine the amount of product each of our reactants could make.
- The theoretical yield will always be the least possible amount of product.
  - ✓ The theoretical yield will always come from the limiting reactant.
- Because of both controllable and uncontrollable factors, the actual yield of product will always be less than the theoretical yield.

Measuring Amounts in the Lab

- In the lab, our balances do not measure amounts in moles, unfortunately, they measure amounts in grams.
- This means we must add two steps to each of our calculations: first convert the amount of each reactant to moles, then convert the amount of product into grams.

Example 8.6—When 11.5 g of C Are Allowed to React with 114.5 g of Cu_2O in the Reaction Below, 87.4 g of Cu Are Obtained.

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

Given: 11.5 g C, 114.5 g Cu_2O, 87.4 g Cu

Find: Limiting reactant, theoretical yield, percent yield

Solution Map:

$$\text{g C} \xrightarrow{\frac{1\ \text{mol}}{12.01\ \text{g}}} \text{mol C} \xrightarrow{\frac{2\ \text{mol Cu}}{1\ \text{mol C}}} \text{mol Cu} \xrightarrow{\frac{1\ \text{mol}}{63.54\ \text{g}}} \text{g Cu}$$

$$\text{g } Cu_2O \xrightarrow{\frac{1\ \text{mol}}{143.02\ \text{g}}} \text{mol } Cu_2O \xrightarrow{\frac{2\ \text{mol Cu}}{1\ \text{mol } Cu_2O}} \text{mol Cu} \xrightarrow{\frac{1\ \text{mol}}{63.54\ \text{g}}} \text{g Cu}$$

Choose smallest

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

Relationships: 1 mol C = 12.01g, 1 mol Cu_2O = 143.02g, 1 mol Cu = 63.54 g, 2 mol Cu = 1 mol Cu, 2 mol Cu = 1 mol Cu_2O

Example 8.6—When 11.5 g of C Are Allowed to React with 114.5 g of Cu_2O in the Reaction Below, 87.4 g of Cu Are Obtained.

$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$, Continued

Solution:

$$11.5\ \text{g}\ \cancel{\text{C}} \times \frac{1\ \cancel{\text{mol C}}}{12.01\ \cancel{\text{g C}}} \times \frac{2\ \cancel{\text{mol Cu}}}{1\ \cancel{\text{mol C}}} \times \frac{63.54\ \text{g Cu}}{1\ \cancel{\text{mol Cu}}} = 122\ \text{g Cu}$$

$$114.5\ \text{g}\ \cancel{Cu_2O} \times \frac{1\ \cancel{\text{mol } Cu_2O}}{143.02\ \cancel{\text{g } Cu_2O}} \times \frac{2\ \cancel{\text{mol Cu}}}{1\ \cancel{\text{mol } Cu_2O}} \times \frac{63.54\ \text{g Cu}}{1\ \cancel{\text{mol Cu}}} = 101.7\ \text{g Cu}$$

The smallest amount is 101.7 g Cu, therefore that is the theoretical yield.

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

The reactant that produces 101.7 g Cu is the Cu_2O,

Therefore, Cu_2O is the limiting reactant.

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

Check: Since the percentage yield is < 100, the answer makes sense.

Example 8.6:

- When 11.5 g of C are allowed to react with 114.5 g of Cu_2O in the reaction below, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.

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

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\,Cu(s) + CO(g)$

- Write down the given quantity and its units.

Given: 11.5 g C
114.5 g Cu_2O
87.4 g Cu produced

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced

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

Find: limiting reactant
theoretical yield
percent yield

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.

- Collect needed conversion factors:

Molar mass Cu_2O = 143.02 g/mol
Molar mass Cu = 63.54 g/mol
Molar mass C = 12.01 g/mol
1 mol $Cu_2O \equiv$ 2 mol Cu (from the chem. equation)
1 mol C $\equiv$ 2 mol Cu (from the chem. equation)

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.
Conversion Factors: 1 mol C = 12.01 g; 1 mol Cu = 63.54 g; 1 mol Cu_2O = 143.08 g; 1 mol Cu_2O ≡ 2 mol Cu; 1 mol C ≡ 2 mol Cu

- Write a solution map:

$$\left.\begin{array}{l} \text{g C} \xrightarrow[\text{12.01 g C}]{\text{1 mol C}} \text{mol C} \xrightarrow[\text{1 mol C}]{\text{2 mol Cu}} \text{mol Cu} \xrightarrow[\text{1 mol Cu}]{\text{63.54 g Cu}} \text{g Cu} \\ \text{g } Cu_2O \xrightarrow[\text{143.08 g } Cu_2O]{\text{1 mol } Cu_2O} \text{mol } Cu_2O \xrightarrow[\text{1 mol } Cu_2O]{\text{2 mol Cu}} \text{mol Cu} \xrightarrow[\text{1 mol Cu}]{\text{63.54 g Cu}} \text{g Cu} \end{array}\right\}$$

Smallest amount is theoretical Yield.

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.
Conversion Factors: 1 mol C = 12.01 g; 1 mol Cu = 63.54 g; 1 mol Cu_2O = 143.08 g; 1 mol Cu_2O ≡ 2 mol Cu; 1 mol C ≡ 2 mol Cu
Solution Map:
g rct → mol rct → mol Cu → g Cu

- Apply the solution map:

$$11.5\text{ g C} \times \frac{1\text{ mole C}}{12.01\text{ g C}} \times \frac{2\text{ mol Cu}}{1\text{ mol C}} \times \frac{63.54\text{ g Cu}}{1\text{ mole Cu}} = 122\text{ g Cu}$$

$$114.5\text{ g } Cu_2O \times \frac{1\text{ mole } Cu_2O}{143.08\text{ g } Cu_2O} \times \frac{2\text{ mol Cu}}{1\text{ mol } Cu_2O} \times \frac{63.54\text{ g Cu}}{1\text{ mole Cu}} = 101.7\text{ g Cu}$$

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.
Conversion Factors: 1 mol C = 12.01 g; 1 mol Cu = 63.54 g; 1 mol Cu_2O = 143.08 g; 1 mol Cu_2O ≡ 2 mol Cu; 1 mol C ≡ 2 mol Cu
Solution Map:
g rct → mol rct → mol Cu → g Cu

- Apply the solution map:

11.5 g C can make 122 g Cu

114.5 g Cu_2O can make **101.7 g Cu**

Least amount

Theoretical yield = 101.7 g Cu

Limiting reactant = Cu_2O

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.
Conversion Factors: 1 mol C = 12.01 g; 1 mol Cu = 63.54 g; 1 mol Cu_2O = 143.08 g; 1 mol Cu_2O ≡ 2 mol Cu; 1 mol C ≡ 2 mol Cu

- Write a solution map:

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

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.
Conversion Factors: 1 mol C = 12.01 g; 1 mol Cu = 63.54 g; 1 mol Cu_2O = 143.08 g; 1 mol Cu_2O ≡ 2 mol Cu; 1 mol C ≡ 2 mol Cu
Solution Map: $\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \text{Percent Yield}$

- Apply the solution map:

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

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

Example:
When 11.5 g of C reacts with 114.5 g of Cu_2O, 87.4 g of Cu are obtained. Find the limiting reactant, theoretical yield, and percent yield.
$Cu_2O(s) + C(s) \rightarrow 2\ Cu(s) + CO(g)$

Information:
Given: 11.5 g C, 114.5 g Cu_2O
87.4 g Cu produced
Find: lim. rct., theor. yld., % yld.
Conversion Factors: 1 mol C = 12.01 g; 1 mol Cu = 63.54 g; 1 mol Cu_2O = 143.08 g; 1 mol Cu_2O ≡ 2 mol Cu; 1 mol C ≡ 2 mol Cu

- Check the solutions:

Limiting reactant = Cu_2O
Theoretical yield = 101.7 g
Percent yield = 85.9%

The percent yield makes sense as it is less than 100%.

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?

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?, Continued

Given: 9.05 g NH_3, 45.2 g CuO

Find: g N_2

Solution Map:

$$\text{g } NH_3 \xrightarrow{\frac{1\text{ mol}}{17.03\text{ g}}} \text{mol } NH_3 \xrightarrow{\frac{1\text{ mol } N_2}{2\text{ mol } NH_3}} \text{mol } N_2 \xrightarrow{\frac{1\text{ mol}}{28.02\text{ g}}} \text{g } N_2$$

$$\text{g CuO} \xrightarrow{\frac{1\text{ mol}}{79.55\text{ g}}} \text{mol CuO} \xrightarrow{\frac{1\text{ mol } N_2}{3\text{ mol CuO}}} \text{mol } N_2 \xrightarrow{\frac{1\text{ mol}}{28.02\text{ g}}} \text{g } N_2$$

Choose smallest

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

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?, Continued

Solution:

$$9.05\ \cancel{\text{g NH}_3} \times \frac{1\ \cancel{\text{mol NH}_3}}{17.03\ \cancel{\text{g NH}_3}} \times \frac{1\ \cancel{\text{mol N}_2}}{2\ \cancel{\text{mol NH}_3}} \times \frac{28.02\ \text{g N}_2}{1\ \cancel{\text{mol N}_2}} = 7.42\ \text{g N}_2$$

$$45.2\ \cancel{\text{g CuO}} \times \frac{1\ \cancel{\text{mol CuO}}}{79.55\ \cancel{\text{g CuO}}} \times \frac{1\ \cancel{\text{mol N}_2}}{3\ \cancel{\text{mol CuO}}} \times \frac{28.02\ \text{g N}_2}{1\ \cancel{\text{mol N}_2}} = \boxed{5.30\ \text{g N}_2 \text{ Theoretical yield}}$$

$$\frac{4.61\ \text{g N}_2}{5.30\ \text{g N}_2} \times 100\% = 87.0\%\ \text{Yield}$$

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

Enthalpy Change

- We previously described processes as exothermic if they released heat, or endothermic if they absorbed heat.
- The **enthalpy of reaction** is the amount of thermal energy that flows through a process.
  - ✓ At constant pressure.
  - ✓ ΔH_{rxn}

Sign of Enthalpy Change

- For exothermic reactions, the sign of the enthalpy change is negative when:
  - ✓ Thermal energy is produced by the reaction.
  - ✓ The surroundings get hotter.
  - ✓ $\Delta H = -$
  - ✓ For the reaction $CH_4(s) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$, the $\Delta H_{rxn} = -802.3$ kJ per mol of CH_4.
- For endothermic reactions, the sign of the enthalpy change is positive when:
  - ✓ Thermal energy is absorbed by the reaction.
  - ✓ The surroundings get colder.
  - ✓ $\Delta H = +$
  - ✓ For the reaction $N_2(s) + O_2(g) \rightarrow 2\ NO(g)$, the $\Delta H_{rxn} = +182.6$ kJ per mol of N_2.

Enthalpy and Stoichiometry

- The amount of energy change in a reaction depends on the amount of reactants.
  - ✓ You get twice as much heat out when you burn twice as much CH_4.
- Writing a reaction implies that amount of energy changes for the stoichiometric amount given in the equation.

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

So 1 mol $C_3H_8 \equiv 5$ mol $O_2 \equiv 3$ mol $CO_2 \equiv 4$ mol $H_2O \equiv$ -2044 kJ.

Example 8.7—How Much Heat Is Associated with the Complete Combustion of 11.8 x 10^3 g of $C_3H_8(g)$?

Given:	11.8 x 10^3 g C_3H_8,
Find:	*heat,* kJ
Solution Map:	g C_3H_8 → mol C_3H_8 → kJ; $\frac{1\ \text{mol}\ C_3H_8}{44.09\ \text{g}}$, $\frac{-2044\ \text{kJ}}{1\ \text{mol}\ C_3H_8}$
Relationships:	1 mol C_3H_8 = -2044 kJ, Molar mass = 44.11 g/mol

Solution:

$$11.8\times10^3\ \text{g}\ \cancel{C_3H_8}\times\frac{1\ \text{mol}\ \cancel{C_3H_8}}{44.11\ \text{g}\ \cancel{C_3H_8}}\times\frac{-2044\ \text{kJ}}{1\ \cancel{\text{mol}\ C_3H_8}}=-5.47\times10^5\ \text{kJ}$$

Check: The sign is correct and the value is reasonable.

Practice—How Much Heat Is Evolved When a 0.483 g Diamond Is Burned?

($\Delta H_{combustion}$ = −395.4 kJ/mol C)

Practice—How Much Heat Is Evolved When a 0.483 g Diamond Is Burned? ($\Delta H_{combustion}$ = −395.4 kJ/mol C), Continued

Given:	0.483 g C
Find:	*heat*, kJ
Solution Map:	g C $\xrightarrow{\frac{1\text{ mol C}}{12.01\text{ g}}}$ mol C $\xrightarrow{\frac{-395.4\text{ kJ}}{1\text{ mol C}}}$ kJ
Relationships:	1 mol C = −395.4 kJ, Molar mass = 12.01 g/mol

Solution:

$$0.483\ \cancel{\text{g C}} \times \frac{1\ \cancel{\text{mol}}\ \text{C}}{12.01\ \cancel{\text{g}}} \times \frac{-395.4\ \text{kJ}}{1\ \cancel{\text{mol}}} = -15.9\ \text{kJ}$$

Check: The sign is correct and the value is reasonable since there is less than 0.1 mol C.