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$_2$. 
• Consider the combustion of octane (C\textsubscript{8}H\textsubscript{18}), a component of gasoline:

\[
2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g)
\]

• The balanced chemical equation shows that 16 mol of CO\textsubscript{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 → 5 pancakes
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?

\[
8 \text{ eggs} \times \frac{5 \text{ pancakes}}{2 \text{ eggs}} = 20 \text{ pancakes}
\]
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 (NH₃).

\[ 3 \text{H}_2(g) + \text{N}_2(g) \rightarrow 2 \text{NH}_3(g) \]
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

Since we do not ordinarily deal with individual molecules, we can express the same ratios in moles.

### 

3 mol \(H_2\) : 1 mol \(N_2\) : 2 mol \(NH_3\)
If we have 3 mol of N\(_2\), and more than enough H\(_2\), how much NH\(_3\) can we make?
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:
  \[ 2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g) \]

• In the case of octane, a shortage of \( \text{O}_2 \) 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, \( \text{CH}_3\text{OC(\text{CH}_3)_3} \)) 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 (\( \text{C}_2\text{H}_5\text{OH} \)), 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.
• 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:
What mass of carbon dioxide is emitted by an automobile per $5.0 \times 10^2$ 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.
\[ 2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g) \]

**SOLUTION MAP:**

\[
\begin{align*}
g \text{C}_8\text{H}_{18} & \rightarrow \text{mol C}_8\text{H}_{18} \\
\frac{1 \text{ mol C}_8\text{H}_{18}}{114.3 \text{ g C}_8\text{H}_{18}} & \rightarrow \frac{16 \text{ mol CO}_2}{2 \text{ mol C}_8\text{H}_{18}} \\
\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} & \rightarrow g \text{ CO}_2
\end{align*}
\]
2 C₈H₁₈(l) + 25 O₂(g) → 16 CO₂(g) + 18 H₂O(g)

SOLUTION:

\[5.0 \times 10^2 \text{ g } C₈H₁₈ \times \frac{1 \text{ mol } C₈H₁₈}{114.3 \text{ g } C₈H₁₈} \times \frac{16 \text{ mol } CO₂}{2 \text{ mol } C₈H₁₈} \times \frac{44.01 \text{ g } CO₂}{1 \text{ mol } CO₂} = 1.5 \times 10^3 \text{ g CO₂}\]
More pancakes
Recall the original equation:

\[1 \text{ cup flour } + 2 \text{ eggs } + \frac{1}{2} \text{ tsp baking powder} \rightarrow 5 \text{ pancakes}\]
Limiting Reactant, Theoretical Yield, and Percent Yield

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

\[
3 \text{ cups flour} \times \frac{5 \text{ pancakes}}{1 \text{ cup flour}} = 15 \text{ pancakes}
\]

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

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

We have enough flour for 15 pancakes, enough eggs for 25 pancakes, and enough baking powder for 40 pancakes.
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.
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** — \((\text{actual yield/ theoretical yield}) \times 100\%\)
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:

1. \(\frac{1 \text{ mol TiCl}_4}{1 \text{ mol Ti}}\)
2. \(\frac{1 \text{ mol TiCl}_4}{2 \text{ mol Cl}_2}\)

Smallest amount determines limiting reactant.
Example: Ti(s) + 2 Cl₂(g) → TiCl₄(s)

Given (moles): 1.8 mol Ti and 3.2 mol Cl₂

Find: limiting reactant and theoretical yield

**SOLUTION:**

\[ \text{Limiting reactant} \]

\[ 1.8 \text{ mol Ti} \times \frac{1 \text{ mol TiCl}_4}{1 \text{ mol Ti}} = 1.8 \text{ mol TiCl}_4 \]

\[ 3.2 \text{ mol Cl}_2 \times \frac{1 \text{ mol TiCl}_4}{2 \text{ mol Cl}_2} = 1.6 \text{ mol TiCl}_4 \]

Limiting reactant: Ti

Least amount of product: 1.6 mol TiCl₄
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: \( \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2 \text{NaCl}(s) \)

Given (grams): 53.2 g Na and 65.8 g Cl\(_2\)

Find: limiting reactant and theoretical yield

**SOLUTION MAP:**

- For Na:
  - \(1 \text{ mol Na} \rightarrow 22.99 \text{ g Na} \)
  - \(\frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \rightarrow \frac{2 \text{ mol NaCl}}{2 \text{ mol Na}} \rightarrow 58.44 \text{ g NaCl} \)

- For Cl\(_2\):
  - \(1 \text{ mol Cl}_2 \rightarrow 70.90 \text{ g Cl}_2 \)
  - \(\frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \rightarrow \frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2} \rightarrow 58.44 \text{ g NaCl} \)

Smallest amount determines limiting reactant.
Limiting Reactant and Percent Yield: Gram to Gram

Example: \( \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2 \text{NaCl}(s) \)

Given (grams): 53.2 g Na and 65.8 g Cl\(_2\)

Find: limiting reactant and theoretical yield

**SOLUTION:**

\[
\begin{align*}
53.2 \text{ g Na} & \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{2 \text{ mol NaCl}}{2 \text{ mol Na}} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 135 \text{ g NaCl} \\
65.8 \text{ g Cl}_2 & \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \times \frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 108 \text{ g NaCl}
\end{align*}
\]

Limiting reactant

Least amount of product
Theoretical Yield and Percent Yield

Example: $\text{Na(s)} + \text{Cl}_2(g) \rightarrow 2 \text{NaCl(s)}$

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.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \frac{86.4 \text{ g}}{108 \text{ g}} \times 100\% = 80.0\%$$
Example 8.6: \( \text{Cu}_2\text{O}(s) + \text{C}(s) \rightarrow 2 \text{Cu}(s) + \text{CO}(g) \)

Given (grams): 11.5 g \( \text{Cu}_2\text{O} \) and 114.5 g \( \text{C} \)
Find: limiting reactant and theoretical yield

**SOLUTION MAP:**

1. **g C** → **mol C**
   - 1 mol C
   - 12.01 g C

2. **mol C** → **mol Cu**
   - 2 mol Cu
   - 1 mol C

3. **mol Cu** → **g Cu**
   - 63.55 g Cu
   - 1 mol Cu

4. **g \( \text{Cu}_2\text{O} \)** → **mol \( \text{Cu}_2\text{O} \)**
   - 1 mol \( \text{Cu}_2\text{O} \)
   - 143.10 g \( \text{Cu}_2\text{O} \)

5. **mol \( \text{Cu}_2\text{O} \)** → **mol Cu**
   - 2 mol Cu
   - 1 mol \( \text{Cu}_2\text{O} \)

6. **mol Cu** → **g Cu**
   - 63.55 g Cu
   - 1 mol Cu

Smallest amount determines limiting reactant.
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.

\[ 1 \text{ mol } \text{Cu}_2\text{O} : 2 \text{ mol } \text{Cu} \]
\[ 1 \text{ mol } \text{C} : 2 \text{ mol } \text{Cu} \]
\[ \text{Molar mass } \text{Cu}_2\text{O} = 143.10 \text{ g/mol} \]
\[ \text{Molar mass } \text{C} = 12.01 \text{ g/mol} \]
\[ \text{Molar mass } \text{Cu} = 63.55 \text{ g/mol} \]
Limiting Reactant and Percent Yield: Gram to Gram

Example 8.6: \( \text{Cu}_2\text{O}(s) + \text{C}(s) \rightarrow 2 \text{Cu}(s) + \text{CO}(g) \)

Given (grams): 11.5 g \( \text{Cu}_2\text{O} \) and 114.5 g C

Find: limiting reactant and theoretical yield

**SOLUTION:**

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

\[
114.5 \text{ g Cu}_2\text{O} \times \frac{1 \text{ mol Cu}_2\text{O}}{143.10 \text{ g Cu}_2\text{O}} \times \frac{2 \text{ mol Cu}}{1 \text{ mol Cu}_2\text{O}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 101.7 \text{ g Cu}
\]

Limiting reactant

Least amount of product
Actual Yield and Percent Yield

Example 8.6: \( \text{Cu}_2\text{O(s)} + \text{C(s)} \rightarrow 2 \text{Cu(s)} + \text{CO(g)} \)

Given (grams): actual yield 87.4 g Cu

Find: percent yield

**SOLUTION:**

Theoretical yield = 101.7 g Cu

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

\[
= \frac{87.4 \text{ g}}{101.7 \text{ g}} \times 100\% = 85.9\%
\]
Enthalpy: 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: A Measure of the Heat Evolved or Absorbed in a Reaction

- We define the enthalpy of reaction, $\Delta H_{\text{rxn}}$, as the amount of thermal energy (or heat) that flows when a reaction occurs at constant pressure.
Sign of $\Delta H_\text{rxn}$

- The *sign* of $\Delta H_\text{rxn}$ (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

• **(a)** In an exothermic reaction, energy is released into the surroundings. **(b)** In an endothermic reaction, energy is absorbed from the surroundings.
Sign of $\Delta H_{\text{rxn}}$

• When thermal energy flows out of the reaction and into the surroundings (as in an exothermic reaction), then $\Delta H_{\text{rxn}}$ is negative.

• The enthalpy of reaction for the combustion of CH$_4$, the main component in natural gas, is as follows:

$$\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)$$

$\Delta H_{\text{rxn}} = -802.3 \text{ kJ}$

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

• The magnitude of $\Delta H_{\text{rxn}}$ tells us that 802.3 kJ of heat are emitted when 1 mol CH$_4$ reacts with 2 mol O$_2$. 
Sign of $\Delta H_{\text{rxn}}$

- When thermal energy flows into the reaction and out of the surroundings (as in an endothermic reaction), then $\Delta H_{\text{rxn}}$ is positive.
- The enthalpy of reaction for the reaction between nitrogen and oxygen gas to form nitrogen monoxide is as follows:

$$\text{N}_2(g) + \text{O}_2(g) \rightarrow 2 \text{NO}(g) \quad \Delta H_{\text{rxn}} = +182.6 \text{ kJ}$$

- This reaction is endothermic and therefore has a positive enthalpy of reaction.
- The magnitude of $\Delta H_{\text{rxn}}$ tells us that 182.6 kJ of heat are absorbed from the surroundings when 1 mol N$_2$ reacts with 1 mol O$_2$. 
Stoichiometry of $\Delta H_{\text{rxn}}$

- The amount of heat emitted or absorbed when a chemical reaction occurs depends on the *amounts* of reactants that actually react.
- We usually specify $\Delta H_{\text{rxn}}$ in combination with the balanced chemical equation for the reaction.
- The magnitude of $\Delta H_{\text{rxn}}$ is for the stoichiometric amounts of reactants and products for the reaction *as written*. 
Stoichiometry of $\Delta H_{\text{rxn}}$

- For example, the balanced equation and $\Delta H_{\text{rxn}}$ for the combustion of propane (the fuel used in LP gas) is as follows:

$$C_3H_8(g) + 5 \text{O}_2(g) \rightarrow 3 \text{CO}_2(g) + 4 \text{H}_2\text{O}(g)$$

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

- 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 are emitted.

- These ratios can be used to construct conversion factors between amounts of reactants or products and the quantity of heat exchanged.
Stoichiometry of $\Delta H_{\text{rxn}}$

- To find out how much heat is emitted upon the combustion of a certain mass in grams of propane $\text{C}_3\text{H}_8$, we can use the following solution map:

\[
\frac{1 \text{ mol } \text{C}_3\text{H}_8}{44.11 \text{ g } \text{C}_3\text{H}_8} \quad \frac{-2044 \text{ kJ}}{1 \text{ mol } \text{C}_3\text{H}_8}
\]
Example 8.7: Stoichiometry Involving $\Delta H_{\text{rxn}}$

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

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

Example: Complete combustion of $11.8 \times 10^3$ g of propane ($C_3H_8$)

**SOLUTION MAP:**

\[
\frac{1 \text{ mol } C_3H_8}{44.11 \text{ g } C_3H_8} \quad \text{and} \quad \frac{-2044 \text{ kJ}}{1 \text{ mol } C_3H_8}
\]

**RELATIONSHIPS USED:**

1 mol $C_3H_8$ : $-2044$ kJ (from balanced equation)

Molar mass $C_3H_8$ = 44.11 g/mol
Example: Complete combustion of $11.8 \times 10^3$ g of propane ($\text{C}_3\text{H}_8$)

**SOLUTION:**

$$11.8 \times 10^3 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.11 \text{ g C}_3\text{H}_8} \times \frac{-2044 \text{ kJ}}{1 \text{ mol C}_3\text{H}_8} = -5.47 \times 10^5 \text{ kJ}$$

Often in the homework, the absolute value of $Q$, $|Q|$, is requested and words are used to convey the sign of the heat absorbed or given off in the reaction.
• 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
• **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 ($\Delta H_{\text{rxn}}$).

• The magnitude of $\Delta H_{\text{rxn}}$ 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.
• 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 $9.0 \times 10^{12}$ 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.
2 C₈H₁₈(ℓ) + 25 O₂(g) → 16 CO₂(g) + 18 H₂O(g)

- The balanced chemical equation shows that 16 mol of CO₂ are produced for every 2 mol of octane burned.
- If the atmosphere currently contains approximately 3.0 \times 10^{15} \text{ kg of CO}_2, how long will it take for the world’s fossil fuel combustion to double the amount of atmospheric carbon dioxide?