Introduction to Chemical Reactions

Chapter 6
Instructional Goals

1. Given the reactants and product in a chemical reaction, the student will be able to write and balance chemical equations.

2. Identify oxidation, reduction, combustion, hydrogenation reactions.

3. Identify hydrolysis, hydration, and dehydration reactions of organic compounds.

4. Use stoichiometric calculations to determine the theoretical yield, and percent yield of a reaction.

5. Describe the difference in energy changes ($\Delta G$) for spontaneous and nonspontaneous reactions, and list the factors that affect the rate of a chemical reaction.
Chemical Reactions

What is Chemistry?

Chemistry is the study of matter and the **changes** it undergoes.

Chemical Changes

Physical Changes
Chemical Reactions

- Reactions involve changes in matter resulting in new substances.

- In chemical reactions, the covalent and ionic bonds that hold elements and compounds together are broken and/or new bonds are formed.

Reactants \rightarrow \text{Products}
Evidence of Chemical Reactions

1) The **color** changes
2) A **solid** is formed.
3) A **gas** is formed.
4) **Heat** is given off or the appearance of **flames**.
5) **Light** is emitted.
6) A new **odor** is emitted.
7) The **temperature** changes.
8) A permanent, new **state** is formed.
Reactions and Energy Changes

Energy can be **released** in a chemical reaction.

\[
\text{methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} + \text{energy}
\]

Energy can be **absorbed** in a chemical reaction.

Example: Reaction of solid barium hydroxide octahydrate (\(\text{Ba(OH)}_2.8\text{H}_2\text{O}\)) and ammonium thiocyanate absorbs energy (\(\text{NH}_4\text{SCN}\)).
Chemical changes are represented using chemical equations.

\[ 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) \]
State of product or reactant

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

(g) = gas
(l) = liquid
(s) = solid
(aq) = aqueous
(dissolve in water)
Things to remember when writing equations:

Identify the state of each chemical

- Metals are solids, except for Hg which is liquid
**Things to remember when writing equations:**

Identify elements that come as **diatomic** molecules

- The following elements come as diatomic elements:
  
  \[ \text{H}_2 \quad \text{N}_2 \quad \text{O}_2 \quad \text{F}_2 \quad \text{Cl}_2 \quad \text{Br}_2 \quad \text{I}_2 \]
Balancing Chemical Equations

Let’s think about “reactions” with something more familiar than chemicals:

For example, if explaining to a small child how to make a cheese sandwich, you might say, “start with two pieces of bread and one slice of cheese.”

\[2 \text{ bread slices} + 1 \text{ cheese slice} \rightarrow 1 \text{ sandwich}\]

\[2B + C \rightarrow B_2C\]
Balancing Chemical Equations

• In a balanced chemical equation, the same number of each atom appears on each side.

• Mr. and Mrs. Lavoisier showed that matter is “conserved”.

• Matter is neither created or destroyed in a chemical reaction.
Balancing Chemical Equations

Example:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

The equation is now “balanced”
Example: Combustion of Methane

- Methane gas burns to produce carbon dioxide gas and water vapor.
  
  - Whenever something burns it combines with $O_2(g)$
  
  $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Fill-in:

$\_\_C + \_\_H + \_\_O \rightarrow \_\_C + \_\_H + \_\_O$
Method for Balancing

1) Make a table and count all atoms on each side of the equation
   • If hydrogen and/or oxygen is present, list them last
     • hydrogen before oxygen.
   • polyatomic ions may be counted as one “element” if it does not change in the reaction

2) Balance an element in the table by adding coefficients to the equation (start with the first element on the list)

3) Recount each atom and Repeat step 2 for all atoms until balanced.
Balancing Chemical Equations

• Nitrogen gas ($\text{N}_2$) can be reacted with oxygen gas ($\text{O}_2$) to produce the dental anesthetic nitrous oxide ($\text{N}_2\text{O}$) also known as *laughing gas*.

$$\text{N}_2 + \text{O}_2 \rightarrow \text{N}_2\text{O}$$

Reactants  Atom  Products

\[
\begin{array}{ccc}
4 \cdot 2 & \leftarrow & \text{N} & \rightarrow & 2 \cdot 4 \\
2 & \leftarrow & \text{O} & \rightarrow & 1 \cdot 2 \\
\end{array}
\]

\[
2\text{N}_2 + \text{O}_2 \rightarrow 2\text{N}_2\text{O}
\]
Balancing Group Work

When gasoline is burned in a car engine, some nitrogen gas \((N_2)\) from the air also burns (reacts with \(O_2\)) to produce nitric oxide (NO), a colorless, toxic gas. Balance the chemical equation for this reaction.

\[ N_2 + O_2 \rightarrow NO \]
Example:

\[ \text{Mg}(s) + \text{O}_2(g) \rightarrow \text{MgO}(s) \]

\[ 2 = 2 \times 1 \quad \leftarrow \quad \text{Mg} \rightarrow \quad 1 \times 2 \]

\[ 2 \quad \leftarrow \quad \text{O} \quad \rightarrow \quad 1 \times 2 = 2 \]

\[ 2\text{Mg}(s) + \text{O}_2(g) \rightarrow \quad 2\text{MgO}(s) \]
Example:
• polyatomic ions may be counted as one “element” if it does not change in the reaction

\[ \text{Al} + \text{FeSO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + \text{Fe} \]

\[ 3 = 3 \times 1 \leftarrow (\text{SO}_4) \rightarrow 3 \]

\[ 2 = 2 \times 1 \leftarrow \text{Al} \rightarrow 2 \]

\[ 3 \leftarrow \text{Fe} \rightarrow 1 \times 3 = 3 \]

\[ 2\text{Al} + 3\text{FeSO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{Fe} \]
Group Work

Balance the following reaction for the combustion of propane:

\[ \text{C}_3\text{H}_8(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \]
Avoiding Common Errors When Balancing Chemical Equations

1. Do not change the formula of a reactant or product.

\[ \text{N}_2 + \text{O}_2 \rightarrow \text{N}_2\text{O} \]

2. Do not add new reactants or products.

\[ \text{N}_2 + \text{O}_2 \rightarrow \text{N}_2\text{O}_2 \]

3. Do not use multiples of the coefficients when writing the balanced equation.

\[ 4\text{N}_2 + 2\text{O}_2 \rightarrow 4\text{N}_2\text{O} \]  \[ 2\text{N}_2 + \text{O}_2 \rightarrow 2\text{N}_2\text{O} \]
Stoichiometry
Stoichiometry allows a chemist or scientist to know how much of an element or reactant to use and how much product is expected to come out of the reaction. Stoichiometry deals with calculations about the masses of reactants and products involved in a chemical reaction.
A balanced equation allows us to predict which reactant (if any) will run out first and how much product can be expected to form.

Before we do calculations with chemicals, let’s do something we are all familiar with:

Food!!!
Suppose you want to make as many cheese sandwiches as possible for lunch. If you have 20 slices of bread, how many slices of cheese do you need?

\[2 \text{ bread slices} + 1 \text{ cheese slice} \rightarrow 1 \text{ sandwich}\]

Balanced equation:

\[
\frac{20 \text{ bread slices}}{2 \text{ bread slices}} = \frac{1 \text{ cheese slice}}{2 \text{ bread slices}} = 10 \text{ cheese slices}
\] 

Relationship Between Bread and Cheese:

\[
\frac{20 \text{ bread slices}}{2 \text{ bread slices}} = \frac{1 \text{ cheese slice}}{2 \text{ bread slices}} = 10 \text{ cheese slices}
\]
Let’s do the same calculation for a chemical reaction now!
Carbon monoxide reacts with oxygen to produce carbon dioxide:

\[ 2\text{CO} + 1\text{O}_2 \rightarrow 2\text{CO}_2 \]

How many \(\text{CO}_2\) molecules are produced from 2 CO molecules? __________

How many \(\text{O}_2\) molecules are needed to produce 2 \(\text{CO}_2\) molecules? __________

How many CO molecules are needed to react with 1 \(\text{O}_2\) molecule? __________
2CO + O₂ → 2CO₂

How many moles of CO₂ are produced from 2 moles of CO?  

How many moles of O₂ are needed to produce 2 moles of CO₂?  

How many moles of CO are needed to react with 1 mole of O₂?
Beginning with 6.0 moles of CO, how many moles of CO\(_2\) will be made if all the CO is used up?

\[
2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2
\]

\[
\frac{2 \text{ moles CO}_2}{2 \text{ moles CO}} = \frac{6.0 \text{ moles CO}}{6.0 \text{ moles CO}_2}
\]

Relationship between moles of CO and moles of CO\(_2\)

\[
\frac{6.0 \text{ moles CO}}{2 \text{ moles CO}_2} = 6.0 \text{ moles CO}_2
\]
How many moles of \( O_2 \) will be needed to react with 6.0 moles of CO?

\[
2\text{CO} + 1\text{O}_2 \rightarrow 2\text{CO}_2
\]

Relationship between moles of CO and moles of \( O_2 \):

\[
\begin{array}{c}
6.0 \text{ moles CO} \\
\frac{1 \text{ mole } O_2}{2 \text{ moles CO}} \\
= 3.0 \text{ moles } O_2
\end{array}
\]

\[
\begin{array}{c}
6.0 \text{ moles CO} \\
\frac{1 \text{ mole } O_2}{2 \text{ moles CO}} \\
= 3.0 \text{ moles } CO_2
\end{array}
\]
How many moles of CO are needed to produce 7.50 moles of CO$_2$?

Assume that you have an unlimited supply of O$_2$.

\[ 2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2 \]

\[
\begin{array}{c|c|c}
7.50 \text{ moles CO}_2 & 2 \text{ moles CO} & 7.50 \text{ moles CO} \\
\hline
2 \text{ moles CO}_2 & 2 \text{ moles CO}_2 & \\
\end{array}
\]

Relationship between moles of CO$_2$ and moles of CO.
Since we can not measure the number of moles on a balance in the lab, we are usually given the number of grams of a reactant or product and need to determine the number of grams of other reactants or products in a reaction.
These calculations only work with **moles**, not with mass (grams)!

We will see why that is using our cheese sandwich example.

If you have 20 grams of bread, how many grams of cheese do you need?

2 bread slices + 1 cheese slice $\rightarrow$ 1 sandwich
Since the coefficients in an equation give us the ratio of moles, we must convert grams to moles in our calculations.
The re-breather units used by some firefighters convert exhaled carbon dioxide ($\text{CO}_2$) into oxygen gas ($\text{O}_2$).

How many grams of oxygen ($\text{O}_2$) can be produced from 1.40 grams of potassium superoxide ($\text{KO}_2$)?

\[4\text{KO}_2(s) + 2\text{CO}_2(g) \rightarrow 2\text{K}_2\text{CO}_3(s) + 3\text{O}_2(g)\]
How many grams of oxygen ($O_2$) can be produced from 1.40 grams of potassium superoxide ($KO_2$)?

$$4KO_2 + 2CO_2 \rightarrow 2K_2CO_3 + 3O_2$$

| 1.40 g $KO_2$ | 1 mole $KO_2$ | 3 mole $O_2$ | 32.0 g $O_2$ | $= 0.473$ g $O_2$
<table>
<thead>
<tr>
<th></th>
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<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>71.1 g $KO_2$</td>
<td>4 mole $KO_2$</td>
<td>1 mole $O_2$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Relationship between moles of $O_2$ and moles of $KO_2$
How many grams of \((\text{CO}_2)\) will react with 1.40 grams of potassium superoxide \((\text{KO}_2)\)?

\[
4\text{KO}_2 + 2\text{CO}_2 \rightarrow 2\text{K}_2\text{CO}_3 + 3\text{O}_2
\]

<table>
<thead>
<tr>
<th>grams KO(_2)</th>
<th>moles KO(_2)</th>
<th>grams CO(_2)</th>
<th>moles CO(_2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.40 g KO(_2)</td>
<td>1 mole KO(_2)</td>
<td>44.0 g CO(_2)</td>
<td>1 mole CO(_2)</td>
</tr>
<tr>
<td>71.1 g KO(_2)</td>
<td>4 mole KO(_2)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Relationship between moles of CO\(_2\) and moles of KO\(_2\)
How many grams of \((\text{KO}_2)\) are needed to produce 5.70 g of \(\text{O}_2\)?

\[
4\text{KO}_2 + 2\text{CO}_2 \rightarrow 2\text{K}_2\text{CO}_3 + 3\text{O}_2
\]

<table>
<thead>
<tr>
<th>5.70 g (\text{O}_2)</th>
<th>1 mole (\text{O}_2)</th>
<th>4 mole (\text{KO}_2)</th>
<th>71.1 g (\text{KO}_2)</th>
<th>= 16.9 g (\text{KO}_2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>32.0 g (\text{O}_2)</td>
<td>3 mole (\text{O}_2)</td>
<td>1 mole (\text{KO}_2)</td>
<td>71.1 g (\text{KO}_2)</td>
<td></td>
</tr>
</tbody>
</table>

Relationship between moles of \(\text{O}_2\) and moles of \(\text{KO}_2\)
How many grams of silver will completely react with 10.0 g of sulfur?

$$2\text{Ag}(s) + \text{S}(s) \rightarrow \text{Ag}_2\text{S}(s)$$
Grams of reactant

Use the molecular weight of the reactant to convert

Moles of reactant

Use the balanced equation to convert

Moles of product

Use the molecular weight of the product to convert

Grams of product
In this course, we will always be given the amount of one reactant and we will assume an excess supply of all other reactants.

If you are curious as to how to do stoichiometric calculations with a limited amount of all reactants, read the section in the text about limiting reagents. (email me for workseets)
Percent Yield

• The amount of product obtained from a reaction is called the **actual yield**.

• To indicate how well the actual yield agrees with the theoretical (calculated) yield, chemists report the **percent yield**:

\[
\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100\%
\]
Example:
A drug company runs a large scale reaction to prepare the pain reliever acetaminophen. The theoretical yield of the reaction is 42 kg of acetaminophen. If 33 kg of acetaminophen are obtained, what is the percent yield of the reaction?

\[
\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100\%
\]
General Types of Reactions

• Read the textbook (section 6 of chapter 6) and take notes on the 4 general types of reactions.
Chemical Reactions

- In this chapter we will look at 5 classes of chemical reactions:
  1) Oxidation-Reduction
     - Combustion
     - Standard Oxidation-Reduction
  2) Hydrogenation
  3) Hydrolysis
  4) Hydration
  5) Dehydration

- You will need to be able to identify the type of reaction and predict the product(s).
Combustion Reactions

• Combustion reactions may occur when organic molecules reacts with oxygen gas ($O_2$).

• This is also called burning!!! In order to burn something you need the 3 things:

  1) A Fuel (hydrocarbon)
  2) Oxygen to burn it with
  3) Something to ignite the reaction (spark)
Combustion Reactions

Example: Combustion of a hydrocarbon :
\[ C_xH_y + O_2(g) \rightarrow CO_2(g) + H_2O(g) \]

- Products in combustion are carbon dioxide and \( H_2O \). (although *incomplete* burning does cause some by-products like carbon monoxide)
- Combustion is a type of oxidation-reduction reactions.
Oxidation and Reduction
Oxidation-Reduction Reactions
All reactions that involve a transfer of one or more electrons are called oxidation-reduction reactions.

• Also called redox reactions.

We say that the substance that loses electrons in the reaction is oxidized and the substance that gains electrons in the reaction is reduced.

• Use OIL RIG to remember
Example: Sodium Reacting with Chlorine Gas

\[ 2Na(s) + Cl_2 (g) \rightarrow 2NaCl(s) \]

- In this process, neutral sodium atoms (charge = 0) lose an electron (are oxidized) to form Na\(^{+}\) and each of the neutral chlorine atoms in Cl\(_2\) gain an electron (are reduced) to form Cl\(^{-}\).

- Since Cl\(_2\) removes electrons from Na, Cl\(_2\) is called the oxidizing agent.

- Because Na gives electrons to Cl\(_2\), Na is the reducing agent.
Oxidation and Reduction of Organic Molecules
Method for determining oxidation or reduction in organic compounds:

An atom in a molecule is oxidized if it:

• gains oxygen (is attached to more oxygen atoms in the product than in the reactant) OR
• loses hydrogen (is attached to fewer hydrogen atoms in the product than in the reactant)

An atom in a molecule is reduced if it:

• loses oxygen (is attached to fewer oxygen atoms) OR
• gains hydrogen (is attached to more hydrogen atoms)
Example
Combustion of methane:

\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}
\]

Is carbon oxidized or is it reduced during this reaction?

When \(\text{CH}_4\) is converted to \(\text{CO}_2\), carbon is oxidized – it loses hydrogen and gains oxygen.
Example

Combustion of methane:

\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

What is reduced? Where did the electrons that carbon lost go to?

\[ \text{O}_2 \text{ is reduced} \] – each oxygen in \( \text{O}_2 \) “lost” an oxygen.
Why do you say carbon is oxidized (loses electrons)?
How is carbon oxidized when this happens?

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

C goes from $\delta^-$ in CH$_4$ to $\delta^+$ in CO$_2$

It *loses* electrons to do so. $\rightarrow$ Oxidation!
How is oxygen reduced when this happens?

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

Non-polar bond
Oxygen atoms have ZERO charge.

O goes from ZERO charge in $\text{O}_2$ to $\delta^-$ in $\text{CO}_2$

It gains electrons to do so. $\rightarrow$ Reduction!
Reactions of Organic Molecules
1) Hydrogenation: Reduction of Alkenes

- Alkenes and other unsaturated hydrocarbons undergo a reduction reaction called catalytic hydrogenation, in which hydrogen gas (H$_2$) in the presence of platinum (Pt), catalyst, acts as a reducing agent.

- During hydrogenation, the carbon atoms in an alkene are reduced because they gain hydrogen atoms.
1) Hydrogenation: Reduction of Alkenes

An alkene + Hydrogen $\rightarrow$ An alkane

\[ \text{Hydrocarbon} \quad \text{C} = \text{C} \quad \text{Hydrocarbon} \]

\[ \text{Hydrocarbon} \quad \text{C} - \text{C} \quad \text{Hydrocarbon} \]

\[ \text{An alkene} \quad \text{H} - \text{H} \quad \text{An alkane} \]
Group Work

Draw and name the product formed when the alkenes react with $\text{H}_2$ and Pt.

a. 1-butene
b. cis-2-butene
c. trans-2-pentene
2) Reactions Involving Water

Water is a reactant or product in a number of reactions important to organic and biochemistry.

In this section we will take a look at three of them –

a) hydrolysis

b) hydration

c) dehydration
2a) Hydrolysis of an Ester

In a hydrolysis reaction, water (hydro) is used to split (lyse) a molecule.

- Esters undergo hydrolysis – when treated with water in the presence of hydroxide ion (OH\(^-\)) they split to form a carboxylate ion and an alcohol.

\[
\text{An ester} \quad \text{CH}_3\text{CH}_2\text{CH}_2\text{C} \quad \text{OCH}_2\text{CH}_3 + \text{H}_2\text{O} \quad \text{OH}^- \quad \rightarrow \quad \text{CH}_3\text{CH}_2\text{CH}_2\text{C} \quad \text{O}^-\quad + \quad \text{HOCH}_2\text{CH}_3
\]

An ester \quad A \text{ carboxylate ion} \quad \text{An alcohol}
2a) Hydrolysis of an Ester

\[ \text{Hydrocarbon} \overset{\text{O}}{\underset{\text{O}}{\text{C}} \text{O}} \text{Hydrocarbon} + \text{HO-O-H} \]

\[ \text{Hydrocarbon} \overset{\text{O}}{\underset{\text{O}}{\text{C}} \text{OH}} + \text{HO-O-H} \text{Hydrocarbon} \]

A carboxylic acid An alcohol
Carboxylic acid vs. Carboxylate ion

A carboxylic acid

This exists as a carboxylate ion in the presence of OH⁻
Hydrolysis is one of the factors that determine the length of time that some drugs remain active.

- The local anesthetics Procaine (also known as Novacain) is a short acting, and remains effective for little more than an hour. A particular enzyme present in the blood serum deactivates this anesthetics by catalyzing hydrolysis of the ester groups.

\[
\text{Procaine} + \text{H}_2\text{O} \rightarrow \text{Procaine}^- + \text{HOCH}_2\text{CH}_2\text{NCH}_2\text{CH}_3
\]
Predict the products formed by the hydrolysis of this ester:

\[
\text{CH}_3\text{CH}_2\text{C} \equiv \text{O} \text{O} - \text{CH}_3 \quad + \quad \text{H}_2\text{O}
\]
2b) Hydration of an Alkene

In a hydration reaction, water is added to a double bond.
2b) Hydration of an Alkene

H–OH

An alcohol

Hydrocarbon

C=C

C

H

H

H

An alkene

Hydrocarbon

C

H

H

H

Hydrocarbon

C

H

H

H

Hydrocarbon
Predict the product formed by the hydration of this alkene:

\[
\text{CH}_3\text{CH}_2\text{C}==\text{CCH}_2\text{CH}_2\text{CH}_3
\]
2c) Dehydration of an Alcohol

Dehydration is the reverse of hydration.

• Water is removed from an alcohol to form a double bond (alkene).
2c) Dehydration of an Alcohol

An alcohol + Hydrocarbon → C\(\text{=C}\)C

Hydrocarbon + C\text{OH} → An alkene

Hydrocarbon
2c) Dehydration Example:

TCA CYCLE
2c) Dehydration Example:

Biochemical dehydration and hydration. The enzyme-catalyzed conversion of citrate into isocitrate involves the dehydration of citrate to form cis-aconitate and the hydration of cis-aconitate to form isocitrate.
Reactions Involving Water: Summary

- Esters can be hydrolyzed.
- Alkenes can be hydrated.
- Alcohols can be dehydrated.
Group Work

Draw the product of each reaction.

a. \[
\text{\text{ }} + \text{H}_2\text{O} \rightarrow
\]

b. \[
\text{CH}_3\text{C} \equiv \text{OCH}_3 + \text{H}_2\text{O} \rightarrow
\]

c. \[
\text{OH} \quad \longrightarrow
\]

\[
\text{CH}_3\text{CHCH}_3
\]
Free Energy and Reaction Rate
Free Energy (G)

For the purposes of this course, you can think of free energy as the total energy:

\[
\text{Total energy (Free energy)} = \text{kinetic energy} + \text{potential energy}
\]

If the free energy of the products \(G_{\text{prod}}\) is less than the free energy of the reactants \(G_{\text{react}}\), then a chemical reaction will occur.

- We live in a universe where matter tends to be at the lowest possible energy (or free energy).
Free Energy $\Delta G$

Consider the combustion of propane:

$$\text{CH}_3\text{CH}_2\text{CH}_3(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$$

$$\Delta G = G_f - G_i$$
Free Energy $\Delta G$

$\text{CH}_3\text{CH}_2\text{CH}_3(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$

- $\Delta G$ for this reaction has a value of -531 kcal/mol
- Because $\Delta G$ is negative, the reaction is spontaneous (happens).

A spontaneous reaction has a negative $\Delta G$ and a nonspontaneous reaction has a positive $\Delta G$. 
Reaction Energy Diagrams

ΔG is negative (-) = Spontaneous

ΔG is positive (+) = Non-Spontaneous
Reaction Rates: How Fast

The rate of a reaction depends on the **activation energy**:

Large Activation Energy = Slow Reaction Rate

Small Activation Energy = Fast Reaction Rate
Reaction Rates: How Fast

The rate of a reaction depends on temperature:

- When temperature increases, the reaction rate increases.
- When temperature decreases, the reaction rate decreases.
Reaction Rates: Catalysts

A catalyst increases the rate of a reaction by lowering the activation energy.
Reaction Rates: Catalysts

Catalysis Example: Lactase (β-galactosidase)

Lactose

Galactose (β 1->4) Glucose

β-galactosidase

Glucose
End of Chapter 6 lecture