

## Lecture Presentation

# Chapter 14 Chemical Kinetics

## Ectotherms

- Lizards and other cold-blooded creatures are **ectotherms**—animals whose body temperature matches their environment's temperature.
- When a lizard's body temperature drops, the chemical reactions that occur in its body slow down, as do all chemical reactions when cooled.
- This causes the lizard to become lethargic and to slow down.
- **Chemical kinetics** is the study of the factors that affect the *rates* of chemical reactions, such as temperature.

## Chemical Kinetics

- The speed of a chemical reaction is called its **reaction rate**.
- The rate of a reaction is a measure of how fast the reaction makes products or uses reactants.
- The ability to control the speed of a chemical reaction is important.

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## Defining Rate

- **Rate** is how much a quantity changes in a given period of time.
- The speed you drive your car is a rate—the distance your car travels (miles) in a given period of time (1 hour).
  - So, the rate of your car has units of mi/hr.

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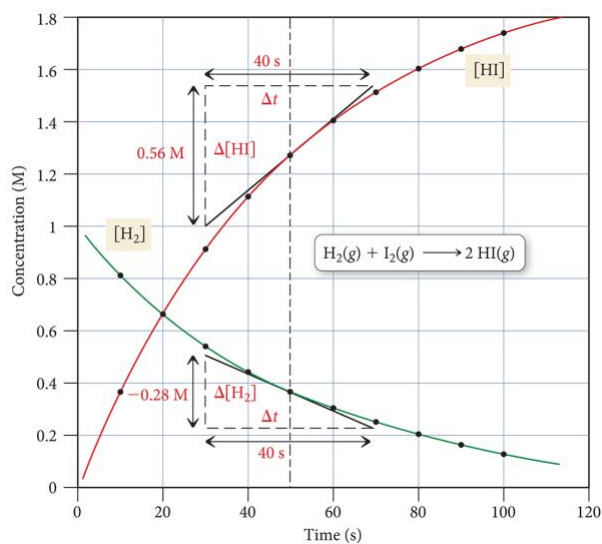
## Defining Reaction Rate

- The rate of a chemical reaction is generally measured in terms of how much the concentration of a reactant decreases (or product concentration increases) in a given period of time.
- For reactants, a negative sign is placed in front of the definition.
  - For the reaction  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightarrow 2 \text{HI}(\text{g})$

$$\text{Rate} = -\frac{\Delta[\text{H}_2]}{\Delta t} = -\frac{[\text{H}_2]_{t_2} - [\text{H}_2]_{t_1}}{t_2 - t_1}$$

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## Reactant and Product Concentrations as a Function of Time



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## The Rate of a Chemical Reaction

$$\text{Rate} = -\frac{\Delta[\text{H}_2]}{\Delta t} = -\frac{[\text{H}_2]_{t_2} - [\text{H}_2]_{t_1}}{t_2 - t_1}$$

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## Average Rate

- The average rate is the change in measured concentrations in any particular time period.
  - Linear approximation of a curve
- The larger the time interval, the more the average rate deviates from the instantaneous rate.

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## The Average Rate of the Reaction

Time (s)	[H <sub>2</sub> ] (M)	$\Delta[\text{H}_2]$ (M)	$\Delta t$ (s)	Rate = $-\Delta[\text{H}_2]/\Delta t$ (M/s)
0.000	1.000	}		
10.000	0.819		10.000	0.0181
20.000	0.670		10.000	0.0149
30.000	0.549		10.000	0.0121
40.000	0.449		10.000	0.0100
50.000	0.368		10.000	0.0081
60.000	0.301		10.000	0.0067
70.000	0.247		10.000	0.0054
80.000	0.202		10.000	0.0045
90.000	0.165		10.000	0.0037
100.000	0.135		10.000	0.0030

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## Instantaneous Rate

- The instantaneous rate is the change in concentration at any one particular time.
  - Slope at one point of a curve
- The instantaneous rate is determined by taking the slope of a line tangent to the curve at that particular point.
  - First derivative of the function (for all of you calculus fans)

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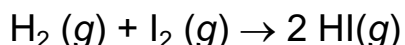
## Reaction Rate Changes over Time

- As time goes on, the rate of a reaction generally slows down because the concentration of the reactants decreases.
- At some time, the reaction stops, either because the reactants run out or because the system has reached equilibrium.

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## Reaction Rate and Stoichiometry

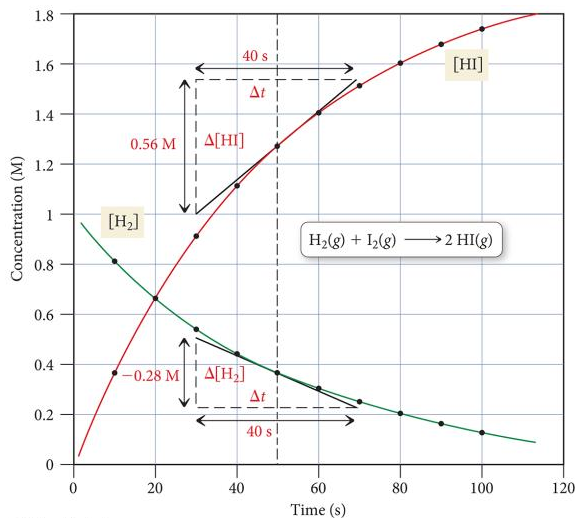
- In most reactions, the coefficients of the balanced equation are not all the same.



- For these reactions, the change in the number of molecules of one substance is a multiple of the change in the number of molecules of another.
  - For the above reaction, for every 1 mole of  $\text{H}_2$  used, 1 mole of  $\text{I}_2$  will also be used and 2 moles of  $\text{HI}$  made.
  - Therefore, the rate of change will be different.
- To be consistent, the change in the concentration of each substance is multiplied by 1/coefficient.

$$\text{Rate} = -\frac{1}{a} \frac{\Delta[\text{A}]}{\Delta t} = -\frac{1}{b} \frac{\Delta[\text{B}]}{\Delta t} = +\frac{1}{c} \frac{\Delta[\text{C}]}{\Delta t} = +\frac{1}{d} \frac{\Delta[\text{D}]}{\Delta t}$$

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Using  $[\text{H}_2]$ , the instantaneous rate at 50 s is as follows:

$$\text{Rate} = -\frac{-0.28 \text{ M}}{40 \text{ s}}$$

$$\text{Rate} = 0.0070 \frac{\text{M}}{\text{s}}$$

Using  $[\text{HI}]$ , the instantaneous rate at 50 s is as follows:

$$\text{Rate} = \left(\frac{1}{2}\right) \frac{0.56 \text{ M}}{40 \text{ s}}$$

$$\text{Rate} = 0.0070 \frac{\text{M}}{\text{s}}$$

## Measuring Reaction Rate

- To measure the reaction rate you need to be able to measure the concentration of at least one component in the mixture at many points in time.
- Some reactions occur slowly enough that samples can be periodically withdrawn from the reaction vessel and analyzed to determine the progress of the reaction.
- Three techniques are commonly used to monitor a reaction mixture: polarimetry, spectroscopy, and pressure measurement.

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## Continuous Monitoring

- **Polarimetry**—This measures the change in the degree of rotation of plane-polarized light caused by one of the components over time.
- **Spectrophotometry**—This measures the amount of light of a particular wavelength absorbed by one component over time.
  - The component absorbs its complementary color.
- **Total pressure**—The total pressure of a gas mixture is stoichiometrically related to partial pressures of the gases in the reaction.

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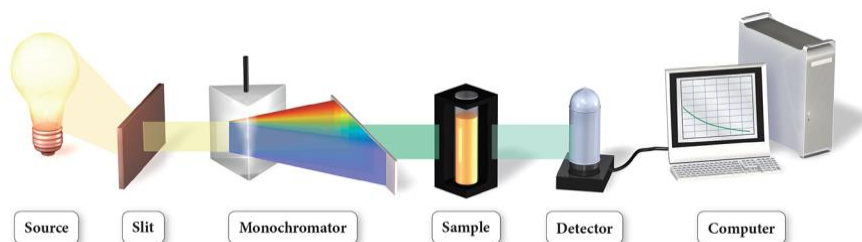
## Sampling the Reaction Mixture at Specific Times

- **Aliquots** (samples from the reaction mixture) are drawn off at specific times during the reaction, and quantitative analysis is performed.
  - Titration for one of the components
  - Gravimetric analysis
- **Gas chromatography** can measure the concentrations of various components in a mixture.
  - For samples that have volatile components
  - Separates mixture by adherence to a surface

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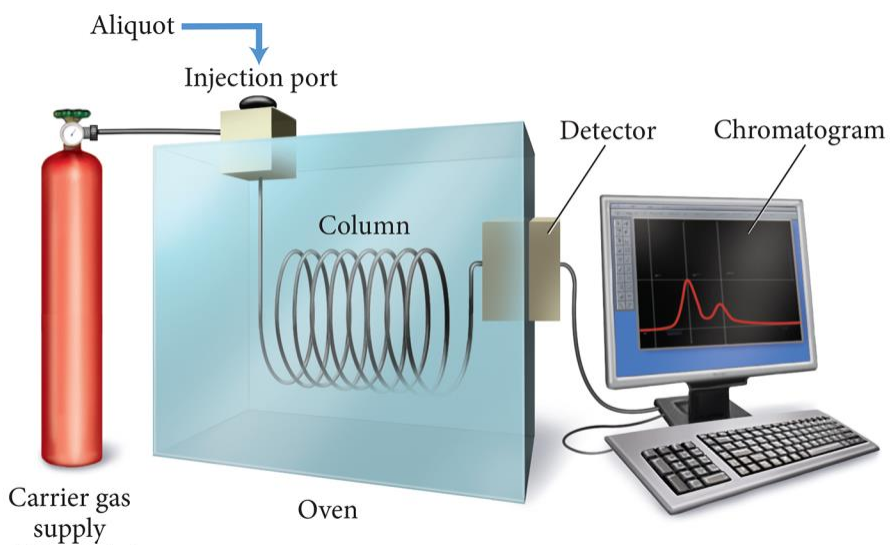


## Methods for Determining Concentrations in a Mixture



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## Methods for Determining Concentrations in a Mixture



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## Factors Affecting Reaction Rate: Reactant Concentration

- Rate often depends on the concentration of one or more of the reactant molecules.
- **Rate law** is an equation relating concentration of reactants to rate when the reverse reaction is negligible.

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## The Rate Law

- The rate of a reaction is directly proportional to the concentration of each reactant raised to a power.
- For the reaction **A → products**, the rate law would have the form given below.

$$\text{Rate} = k[\text{A}]^n$$

- $n$  is called the **order**; usually, it is an integer that determines rate dependence on reactant concentration.
- $k$  is called the **rate constant**.

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## Reaction Order

- The exponent on each reactant in the rate law is called the **order** with respect to that reactant.
- The sum of the exponents on the reactants is called the **order of the reaction**.
- In the rate law,  $\text{Rate} = k[\text{NO}]^2[\text{O}_2]$ , the reaction is second order with respect to  $[\text{NO}]$ , first order with respect to  $[\text{O}_2]$ , and third order overall.

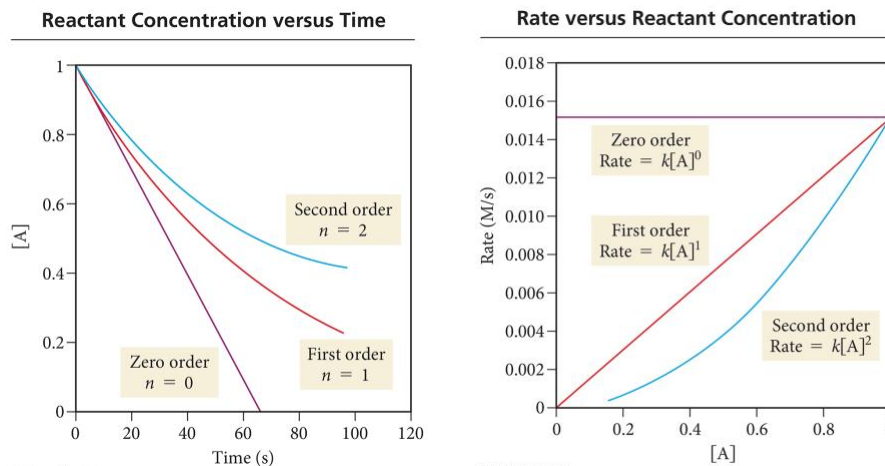
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## Rate = $k[\text{A}]^n$

- If a reaction is zero order, the rate of the reaction is always the same.
  - Doubling  $[\text{A}]$  will have no effect on the reaction rate.
- If a reaction is first order, the rate is directly proportional to the reactant concentration.
  - Doubling  $[\text{A}]$  will double the rate of the reaction.
- If a reaction is second order, the rate is directly proportional to the square of the reactant concentration.
  - Doubling  $[\text{A}]$  will quadruple the rate of the reaction.

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## Reactant Concentration versus Time $A \rightarrow \text{Products}$



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## Determining the Order of a Reaction

- The rate law *must* be determined experimentally.
- We can use the **method of initial rates**, where data from different experiments with varying starting concentrations of reactants and the corresponding initial rates are given.
- Determine how rate is impacted by change in a single reactant in two different experiments.

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## Determining the Rate Law When There Are Multiple Reactants

- Changing each reactant will affect the overall rate of the reaction.
- By changing the initial concentration of one reactant at a time, the effect of each reactant's concentration on the rate can be determined.
- In examining results, we compare differences in rate for reactions that differ only in the concentration of one reactant.

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## Integrated Rate Laws

- The rate law shows the relationship between rate and concentration.
- It is useful to have an equation relating concentration with time.
- Using calculus we can obtain the **integrated rate law** that shows the relationship between the concentration of A and the time of the reaction.

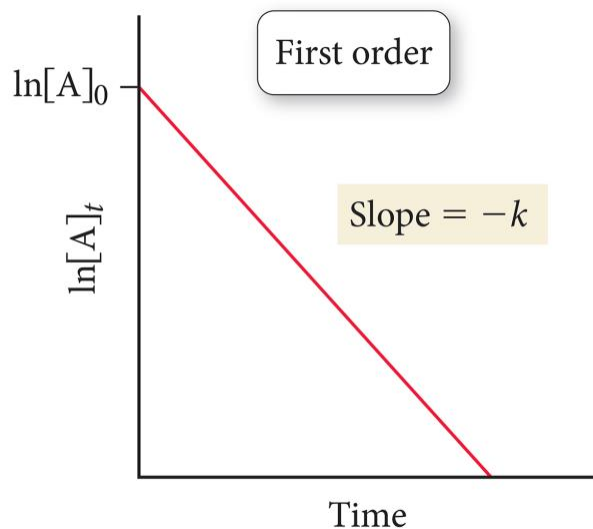
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## First-Order Reactions

- Rate law:  $\text{rate} = k[A]^1 = k[A]$
- Integrated rate law:  $\ln[A] = -kt + \ln[A]_{\text{initial}}$
- Graph  $\ln[A]$  versus time—straight line with **slope =  $-k$**  and **y-intercept =  $\ln[A]_{\text{initial}}$** 
  - Used to determine the rate constant
- When rate = M/sec,  **$k = \text{s}^{-1}$**

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## First-Order Integrated Rate Law



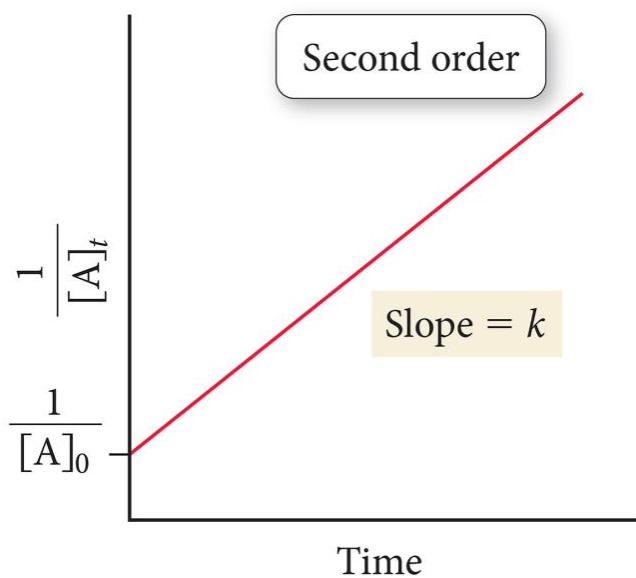
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## Second-Order Reactions

- Rate =  $k[A]^2$
- $1/[A] = kt + 1/[A]_{\text{initial}}$
- Graph  $1/[A]$  versus time—straight line with **slope =  $k$**  and **y-intercept =  $1/[A]_{\text{initial}}$** 
  - Used to determine the rate constant
- $t_{1/2} = 1/(k[A]_0)$
- When Rate = M/sec,  $k = \text{M}^{-1} \cdot \text{s}^{-1}$

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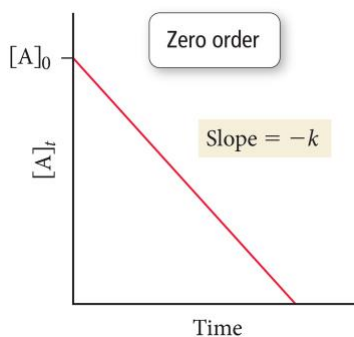
## Second-Order Integrated Rate Law



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## Zero-Order Reactions

- Rate =  $k[A]^0 = k$   
– Constant rate reactions
- $[A] = -kt + [A]_{\text{initial}}$
- Graph of  $[A]$  versus time—straight line with **slope =  $-k$**  and **y-intercept =  $[A]_{\text{initial}}$**
- $t_{1/2} = [A]_{\text{initial}}/2k$
- When Rate = M/sec,  **$k = \text{M/sec}$**



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## Half-Life

- The **half-life**,  $t_{1/2}$ , of a reaction is the time required for the concentration of the reactant to fall to half its initial value.
- The half-life of the reaction depends on the order of the reaction.

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## The Half-Life of a First-Order Reaction Is Constant

### Half-Life for a First-Order Reaction

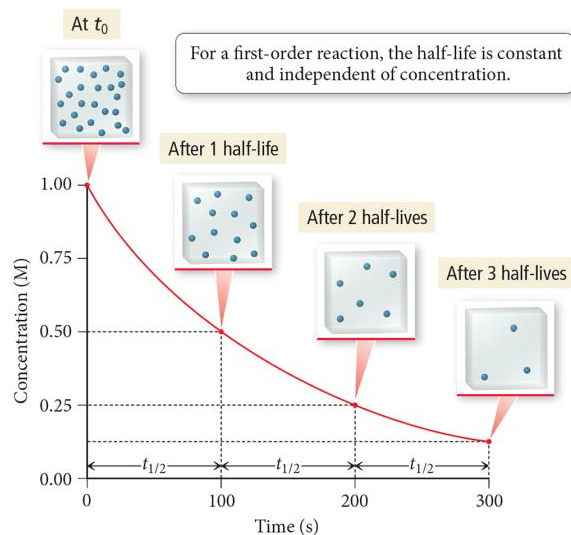


TABLE 14.2 Rate Law Summary Table

Order	Rate Law	Units of $k$	Integrated Rate Law	Straight-Line Plot	Half-Life Expression
0	Rate = $k[A]^0$	$M \cdot s^{-1}$	$[A]_t = -kt + [A]_0$	<p>y-intercept = <math>[A]_0</math> Slope = <math>-k</math></p>	$t_{1/2} = \frac{[A]_0}{2k} = \frac{1}{k} \frac{[A]_0}{2}$
1	Rate = $k[A]^1$	$s^{-1}$	$\ln[A]_t = -kt + \ln[A]_0$ $\ln \frac{[A]_t}{[A]_0} = -kt$	<p>y-intercept = <math>\ln[A]_0</math> Slope = <math>-k</math></p>	$t_{1/2} = \frac{0.693}{k} = \frac{1}{k} (0.693)$
2	Rate = $k[A]^2$	$M^{-1} \cdot s^{-1}$	$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$	<p>Slope = <math>k</math> y-intercept = <math>1/[A]_0</math></p>	$t_{1/2} = \frac{1}{k[A]_0} = \frac{1}{k} \frac{1}{[A]_0}$

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## Graphical Determination of the Rate Law for $A \rightarrow \text{Product}$

- Plots of  $[A]$  versus time,  $\ln[A]$  versus time, and  $1/[A]$  versus time allow determination of whether a reaction is zero, first, or second order.
- Whichever plot gives a straight line determines the order with respect to  $[A]$ .
  - If linear is  $[A]$  versus time,  $\text{Rate} = k[A]^0$ .
  - If linear is  $\ln[A]$  versus time,  $\text{Rate} = k[A]^1$ .
  - If linear is  $1/[A]$  versus time,  $\text{Rate} = k[A]^2$ .

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## The Effect of Temperature on Rate

- The rate constant of the rate law,  $k$ , is temperature dependent.
- The Arrhenius equation shows the relationship:

$$k = A e^{\frac{-E_a}{RT}}$$

Diagram labels:   
 - **Frequency factor** points to  $A$ .   
 - **Exponential factor** points to  $e^{\frac{-E_a}{RT}}$ .   
 - **Activation energy** points to  $E_a$ .

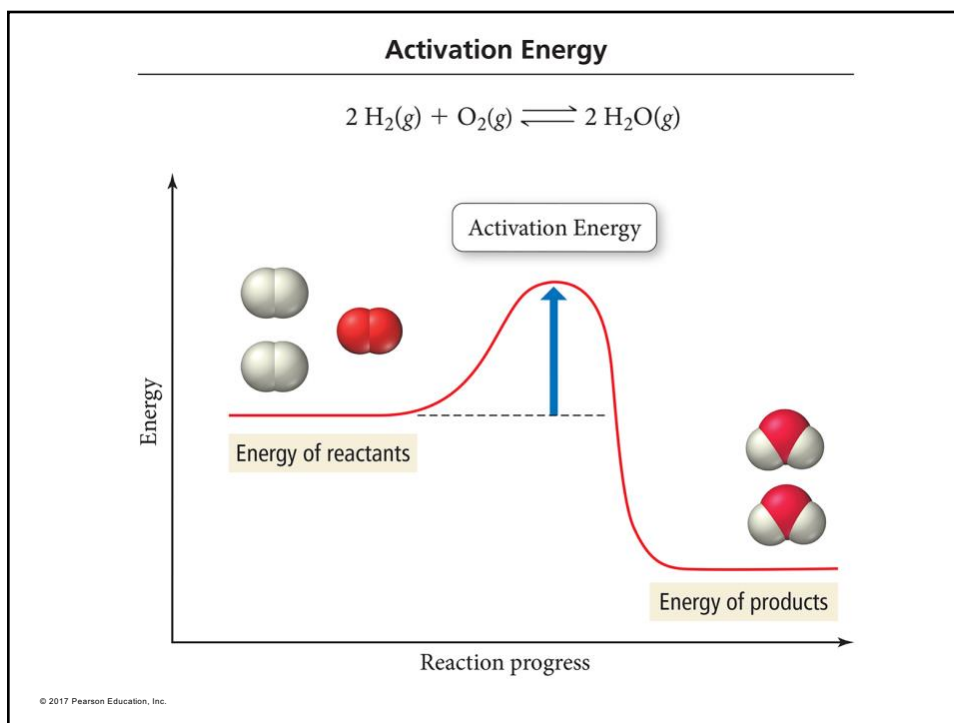
where  $T$  is the temperature in kelvin.

$R$  is the gas constant in energy units,  $8.314 \text{ J}/(\text{mol} \cdot \text{K})$ .

$A$  is called the **frequency factor**, the rate the reactant energy approaches the activation energy.

$E_a$  is the **activation energy**, the minimum energy needed to start the reaction.

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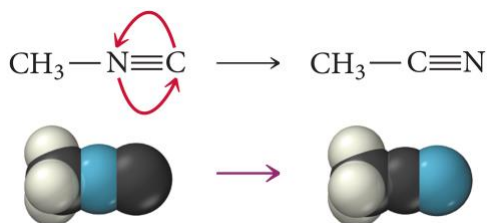


## Activation Energy and the Activated Complex

- There is an energy barrier to almost all reactions.
- The **activation energy** is the amount of energy needed to convert reactants into the activated complex.
  - Also known as transition state
- The **activated complex** is a chemical species with partially broken and partially formed bonds.
  - Always very high in energy because of its partial bonds

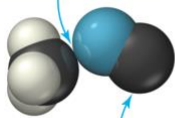
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## Isomerization of Methyl Isonitrile



Methyl isonitrile rearranges to acetonitrile.

Bond weakens

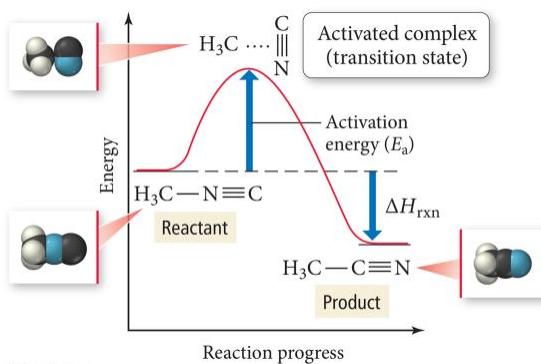


NC group begins to rotate

For the reaction to occur, the  $\text{H}_3\text{C}-\text{N}$  bond must break, and a new  $\text{H}_3\text{C}-\text{C}$  bond must form.

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## Energy Profile for the Isomerization of Methyl Isonitrile

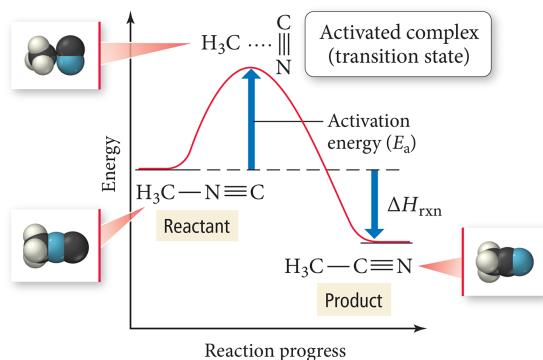


As the reaction begins, the  $\text{C}-\text{N}$  bond weakens enough for the  $\text{C}\equiv\text{N}$  group to start to rotate.

The **activated complex** is a chemical species with partial bonds.

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## Energy Profile for the Isomerization of Methyl Isonitrile

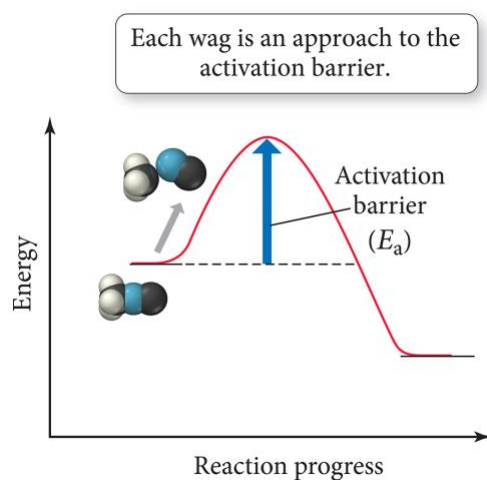


The **activation energy** is the difference in energy between the reactants and the activated complex.

The **frequency** is the number of molecules that begin to form the activated complex in a given period of time.

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## The Frequency Factor



The frequency factor represents the number of approaches to the activation barrier per unit time.

For this reaction, the frequency factor represents the rate at which the NC part of the molecule wags (vibrates side to side).

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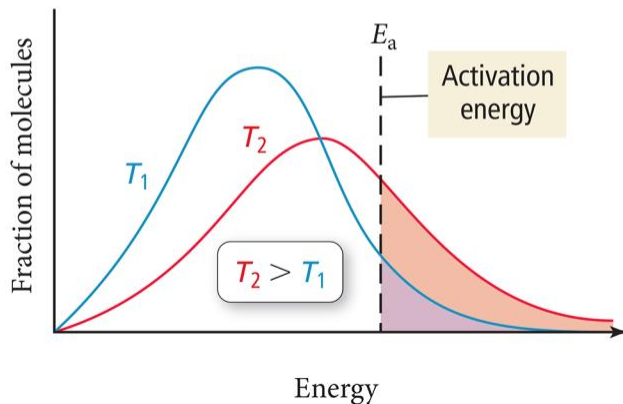
## Arrhenius Equation: The Exponential Factor

- The exponential factor in the Arrhenius equation,  $e^{-E_a/RT}$ , is a number between 0 and 1.
- It represents the fraction of reactant molecules with sufficient energy that can make it over the energy barrier.
  - The higher the energy barrier (larger activation energy), the fewer molecules that have sufficient energy to overcome it.
- Increasing temperature will increase the number of molecules with sufficient energy to overcome the energy barrier.
- Increasing the temperature will increase the reaction rate.

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## Thermal Energy Distribution

As temperature increases, the fraction of molecules with enough energy to surmount the activation energy barrier also increases.



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## Arrhenius Plots

- The Arrhenius equation can be algebraically solved to give the following form:

$$\ln k = -\frac{E_a}{R} \left( \frac{1}{T} \right) + \ln A$$

$$y = mx + b$$

This equation is in the form  $y = mx + b$ ,  
where  $y = \ln(k)$  and  $x = (1/T)$ .

A graph of  $\ln(k)$  versus  $(1/T)$  is a straight line.

$(-8.314 \text{ J/mol} \cdot \text{K})(\text{slope of the line}) = E_a$ , (in Joules)

$e^{y\text{-intercept}} = A$  (unit is the same as  $k$ )

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## Arrhenius Equation: Two-Point Form

- If you only have two  $(T, k)$  data points, the following forms of the Arrhenius equation can be used:

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

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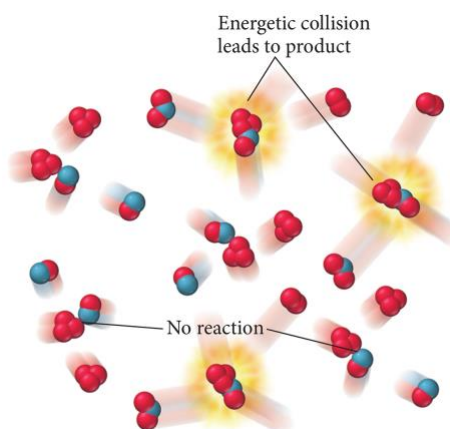
## Collision Theory

- In **collision theory**, for a reaction to take place, the reacting molecules must collide with one another.
  - On average, about  $10^9$  collisions per second
- Once molecules collide, they may react together or they may not, depending on two factors:
  - Whether the collision has enough energy to “break the bonds holding reactant molecules together”
  - Whether the reacting molecules collide in the proper orientation for new bonds to form

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## Effective Collisions: Kinetic Energy Factor

For a collision to lead to overcoming the energy barrier, the reacting molecules must have sufficient kinetic energy so that when they collide, the activated complex can form.



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## The Frequency Factor of the Arrhenius Equation

- The Arrhenius equation includes a term, **A**, called the **frequency factor**.
- The frequency factor can be broken into two terms that relate to the two factors that determine whether a collision will be effective.

$$k = Ae^{\frac{-E_a}{RT}}$$

$$= pze^{\frac{-E_a}{RT}}$$

Orientation factor
Collision frequency

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## Effective Collisions: Orientation Effect



Ineffective collision



Ineffective collision



Effective collision

The molecules must be aligned in a very specific way for the reaction to occur.

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## Effective Collisions

- Collisions that lead to reaction are called **effective collisions**.
- The higher the frequency of effective collisions, the faster the reaction rate.

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## Collision Frequency

- The collision frequency is the number of collisions that happen per second.
- The more collisions there are per second, the more collisions can be effective and lead to product formation.

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## Orientation Factor

- The orientation factor,  $p$ , is a statistical term relating the frequency factor to the collision frequency.
- For most reactions,  $p < 1$ .
- Generally, the more complex the reactant molecules, the smaller the value of  $p$ .
- For reactions involving atoms colliding,  $p \approx 1$  because of the spherical nature of the atoms.
- Some reactions actually can have a  $p > 1$ .
  - Generally involve electron transfer

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## Orientation Factor

- The proper orientation results when the atoms are aligned in such a way that the old bonds can break and the new bonds can form.
- For most reactions, the orientation factor is less than 1.

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## Reaction Mechanisms

- When we write a chemical equation to represent a chemical reaction, we usually represent the overall reaction, not the series of individual steps by which the reaction occurs.
- The **reaction mechanism** is the series of individual chemical steps by which an overall chemical reaction occurs.
- Knowing the rate law of the reaction helps us understand the sequence of reactions in the mechanism.

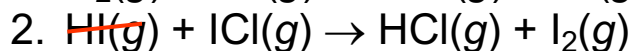
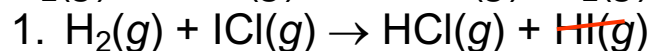
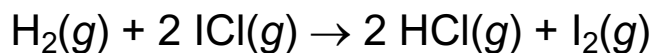
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## An Example of a Reaction Mechanism

- Overall reaction:
$$\text{H}_2(g) + 2 \text{ICl}(g) \rightarrow 2 \text{HCl}(g) + \text{I}_2(g)$$
- Mechanism:
  1.  $\text{H}_2(g) + \text{ICl}(g) \rightarrow \text{HCl}(g) + \text{HI}(g)$
  2.  $\text{HI}(g) + \text{ICl}(g) \rightarrow \text{HCl}(g) + \text{I}_2(g)$
- The reactions in this mechanism are **elementary steps**, meaning that they cannot be broken down into simpler steps and that the molecules actually interact directly in this manner without any other steps.

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## Elements of Mechanism Intermediates



- Notice that the HI is a product in step 1, but then a reactant in step 2.
- Because HI is made but then consumed, HI does not show up in the overall reaction.
- Materials that are products in an early mechanism step, but then reactants in a later step, are called **reaction intermediates**.

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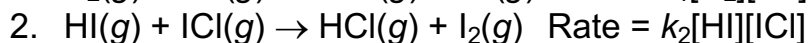
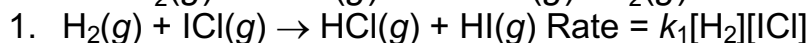
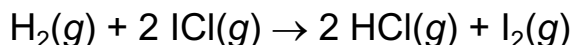
## Molecularity

- The number of reactant particles in an elementary step is called its **molecularity**.
- A **unimolecular** step involves one particle.
- A **bimolecular** step involves two particles.
  - However, they may be the same kind of particle.
- A **termolecular** step involves three particles.
  - However, these are exceedingly rare in elementary steps.

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## Rate Laws for Elementary Steps

- Each step in the mechanism is like its own little reaction with its own activation energy and own rate law.
- The rate law for an overall reaction must be determined experimentally.
- But, the rate law of an elementary step can be deduced from the equation of the step.



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## Rate Laws of Elementary Steps

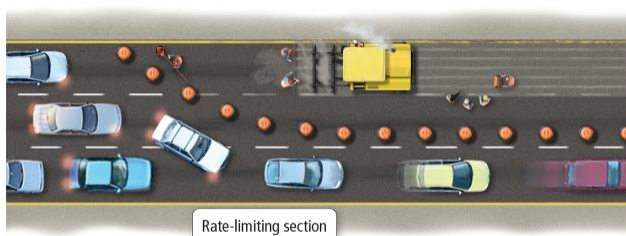
**TABLE 14.3** Rate Laws for Elementary Steps

Elementary Step	Molecularity	Rate Law
$\text{A} \longrightarrow \text{products}$	1	$\text{Rate} = k[\text{A}]$
$\text{A} + \text{A} \longrightarrow \text{products}$	2	$\text{Rate} = k[\text{A}]^2$
$\text{A} + \text{B} \longrightarrow \text{products}$	2	$\text{Rate} = k[\text{A}][\text{B}]$
$\text{A} + \text{A} + \text{A} \longrightarrow \text{products}$	3 (rare)	$\text{Rate} = k[\text{A}]^3$
$\text{A} + \text{A} + \text{B} \longrightarrow \text{products}$	3 (rare)	$\text{Rate} = k[\text{A}]^2[\text{B}]$
$\text{A} + \text{B} + \text{C} \longrightarrow \text{products}$	3 (rare)	$\text{Rate} = k[\text{A}][\text{B}][\text{C}]$

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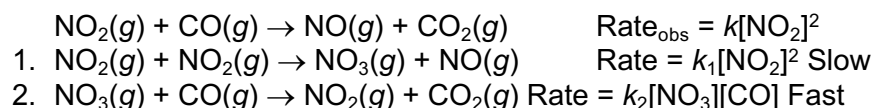
## Rate-Determining Step

- In most mechanisms, one step occurs much slower than the other steps.
- We call the slowest step in the mechanism the **rate-determining step (RDS)**.
  - The slowest step has the largest activation energy.
- The rate law of the rate-determining step determines the rate law for the overall reaction.



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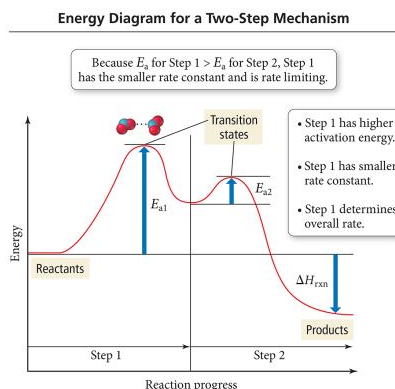
## Another Reaction Mechanism



The first step is slower than the second step because its activation energy is larger.

The first step in this mechanism is the **RDS**.

The rate law of the first step is the same as the rate law of the overall reaction.



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## Validating a Mechanism

To validate (not prove) a mechanism, two conditions must be met:

1. The elementary steps must sum to the overall reaction.
2. The rate law predicted by the mechanism must be consistent with the experimentally observed rate law.

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## Mechanisms with a Fast Initial Step

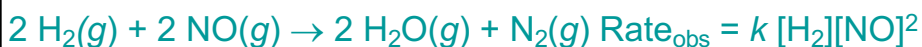
- When a mechanism contains a fast initial step, the RDS may contain intermediates.
- When a previous step is rapid and reaches equilibrium, the forward and reverse reaction rates are equal, so the concentrations of reactants and products of the step are related, and the product is an intermediate.
- Substituting into the rate law of the RDS will produce a rate law in terms of just reactants.

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## An Example

1.  $2 \text{NO}(g) \xrightleftharpoons[k_{-1}]{k_1} \text{N}_2\text{O}_2(g)$  Fast
2.  $\text{H}_2(g) + \text{N}_2\text{O}_2(g) \xrightarrow{k_2} \text{H}_2\text{O}(g) + \text{N}_2\text{O}(g)$  Slow Rate =  $k_2[\text{H}_2][\text{N}_2\text{O}_2]$
3.  $\text{H}_2(g) + \text{N}_2\text{O}(g) \xrightarrow{k_3} \text{H}_2\text{O}(g) + \text{N}_2(g)$  Fast



for Step 1  $\text{Rate}_{\text{forward}} = \text{Rate}_{\text{reverse}}$

$$k_1[\text{NO}]^2 = k_{-1}[\text{N}_2\text{O}_2]$$

$$\text{Rate} = k_2[\text{H}_2][\text{N}_2\text{O}_2]$$

$$[\text{N}_2\text{O}_2] = \frac{k_1}{k_{-1}}[\text{NO}]^2$$

$$\text{Rate} = k_2[\text{H}_2] \frac{k_1}{k_{-1}}[\text{NO}]^2$$

Substituting this definition into the rate law, we get:

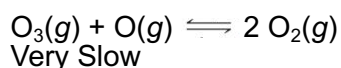
$$\text{Rate} = \frac{k_2 k_1}{k_{-1}}[\text{H}_2][\text{NO}]^2$$

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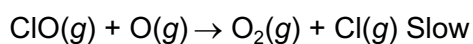
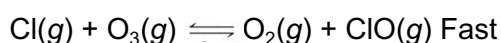
## Catalysts

- **Catalysts** are substances that affect the rate of a reaction without being consumed.
- A catalyst works by providing an alternative mechanism for the reaction with a lower activation energy.
- A catalyst is consumed in an early mechanism step and then made in a later step.

Mechanism without catalyst:

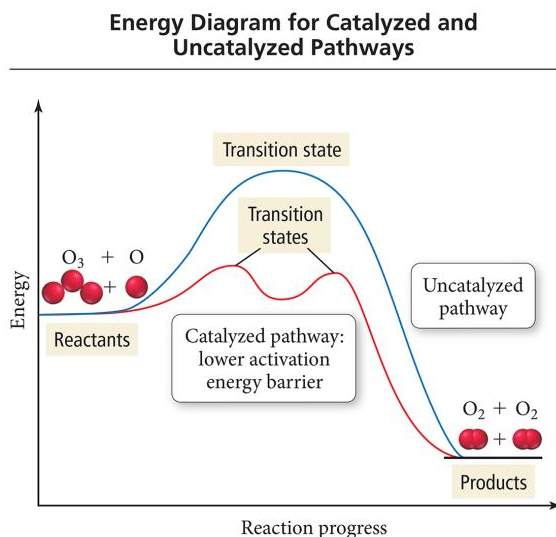


Mechanism with catalyst:



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## Energy Profile of a Catalyzed Reaction



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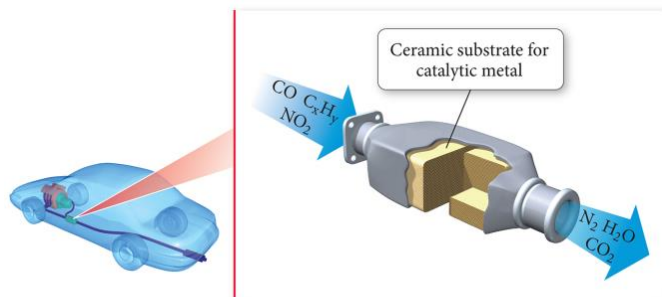
## Homogeneous and Heterogeneous Catalysis

- Catalysts give the reactant molecules a different path to follow with a lower activation energy.
- **Heterogeneous catalysts** are in the same phase as the reactant particles.
  - They hold one reactant molecule in proper orientation for reaction to occur when the collision takes place, sometimes helping to start breaking bonds.
- **Homogeneous catalysts** are in a different phase than the reactant particles.
  - They react with one of the reactant molecules to form a more stable activated complex with a lower activation energy.

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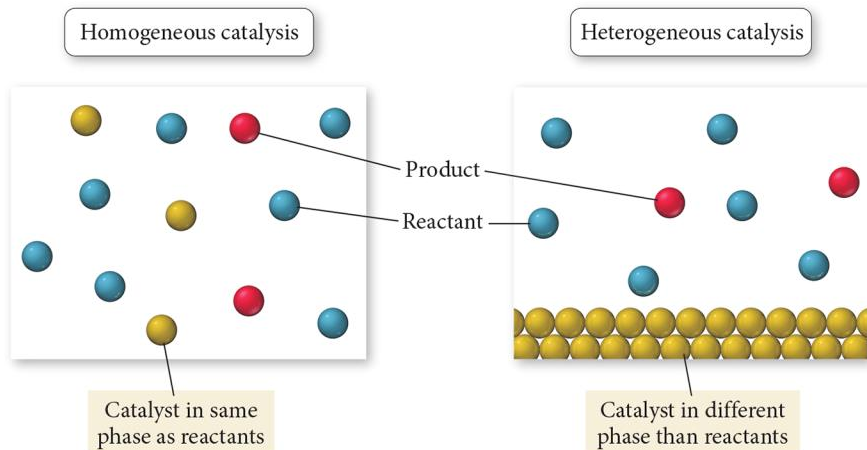
## Catalysts

The catalytic converter in the exhaust system of a car helps eliminate pollutants in the exhaust.



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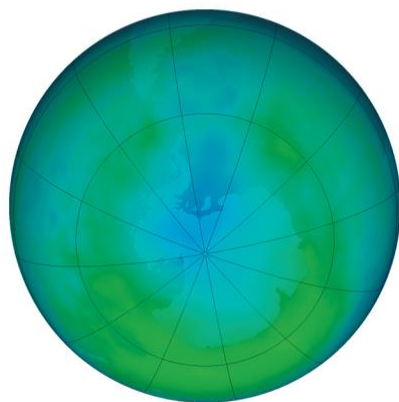
## Types of Catalysts



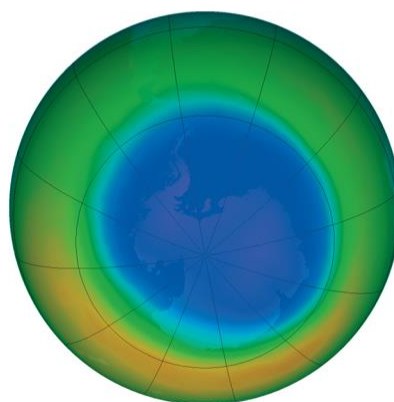
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## Ozone Depletion over the Antarctic

May



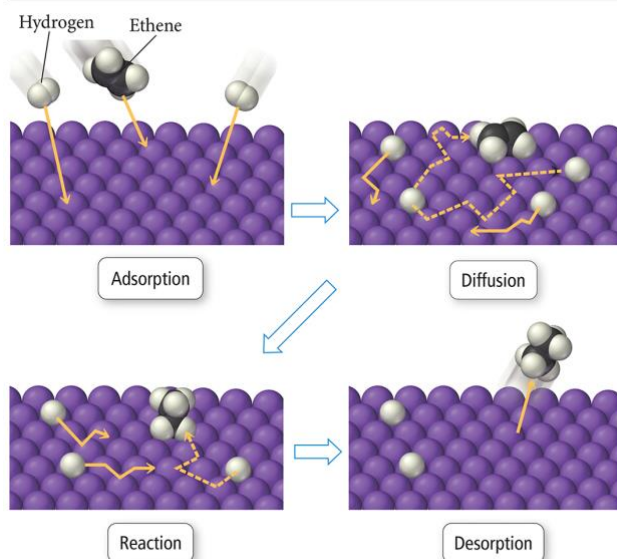
September



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## Catalytic Hydrogenation of Ethane

### Heterogeneous Catalysis



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## Enzymes: Biological Catalysts

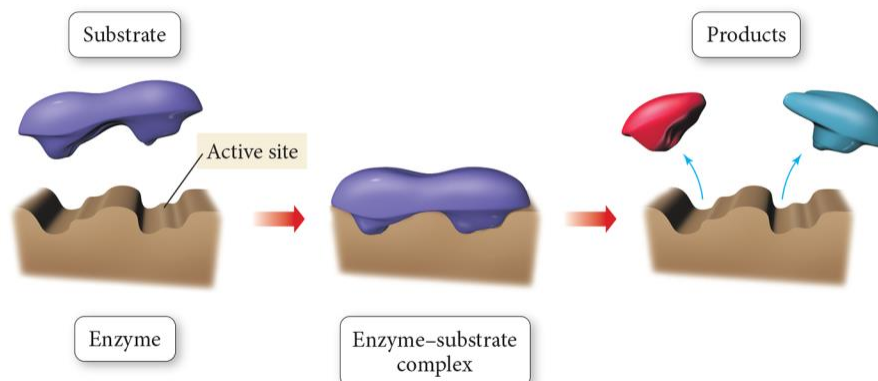
- Because many of the biological molecules are large and complex, most biological reactions require a catalyst to proceed at a reasonable rate.
- Protein molecules that catalyze biological reactions are called **enzymes**.
- Enzymes work by adsorbing the substrate reactant onto an **active site** that orients the **substrate** for reaction.

1. Enzyme + Substrate  $\rightleftharpoons$  Enzyme–Substrate Fast
2. Enzyme–Substrate  $\rightarrow$  Enzyme + Product Slow

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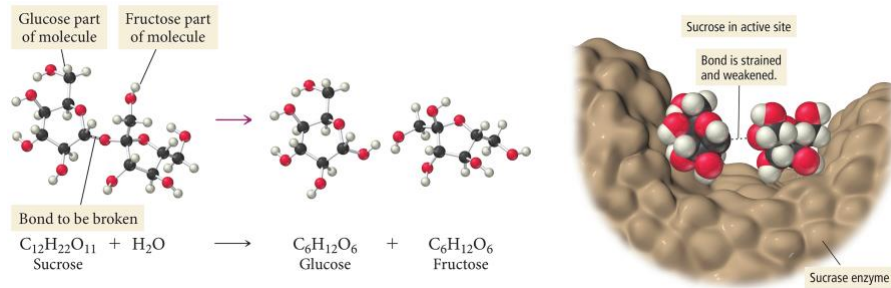
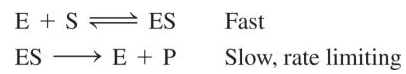
## Enzyme–Substrate Binding: The Lock and Key Mechanism

### Enzyme–Substrate Binding



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## Enzymatic Hydrolysis of Sucrose



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