

Chapter 14 Chemical Kinetics

Chemical reactions

What we need to know about a chemical reaction are:

- Stoichiometry
 - Balanced chemical equation
- Thermodynamic
 - Endothermic or exothermic Chapter 5
 - Spontaneous Chapter 19
- Kinetic Chapter 14
 - Reaction rate

Kinetics

- In kinetics we study the <u>rate</u> at which a chemical process occurs.
- Besides information about the speed at which reactions occur, kinetics also sheds light on the reaction mechanism (exactly how the reaction occurs).

- Physical state of the reactants
 - In order to react, molecules must come in contact with each other.
 - The more <u>homogeneous</u> the mixture of reactants, the faster the molecules can react.



Steel wool heated in air (about 20% O₂) glows red-hot but oxidizes to Fe₂O₃ slowly



Red-hot steel wool in 100% O_2 burns vigorously, forming Fe_2O_3 quickly © 2012 Pearson Education, Inc.

Concentration of reactants

 As the concentration of reactants increases, so does the likelihood that reactant molecules will collide.

Temperature

 At higher temperatures, reactant molecules have more <u>kinetic energy</u>, move faster, and collide more often and with greater energy.

- Presence of a <u>catalyst</u>
 - Catalysts speed up reactions by changing the mechanism of the reaction.
 - Catalysts are not consumed during the course of the reaction.

The Dust Explosion

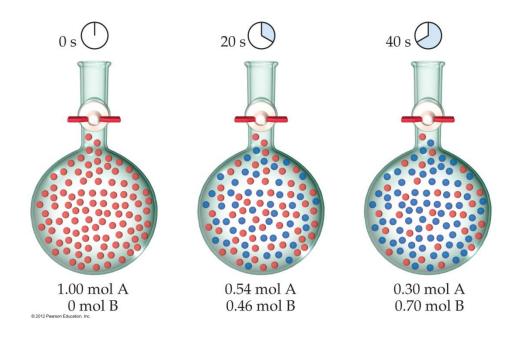
https://youtu.be/TAdEIO1FCSM

$$(C_6H_{10}O_5)_n(s) + 6nO_2(g) \rightarrow 6nCO_2(g) + 5nH_2O(f)$$

starch







Rates of reactions can be determined by monitoring the change in concentration of either reactants or products as a function of time.

$$C_4H_9CI(aq) + H_2O(I) \longrightarrow C_4H_9OH(aq) + HCI(aq)$$

TABLE 14.1 • Rate Data for Reaction of C ₄ H ₉ Cl with Water			
Time, $t(s)$	$[C_4H_9Cl](M)$	Average Rate (M/s)	
0.0	0.1000	$2:2:= 1.9 \times 10^{-4}$	
50.0	0.0905 -:=:=:	1.9×10^{-4}	
100.0	0.0820 -:=:=:	1.7×10^{-4} 1.6×10^{-4}	
150.0	0.0741 -:=:=:	1.6×10^{-4}	
200.0	0.0671 -:=:=:	1.4×10^{-4} 1.22×10^{-4}	
300.0	0.0549 -:=:=:	1.22×10 1.01×10^{-4}	
400.0	0.0448 -:=:=:	$2.2.3.6 \times 10^{-4}$	
500.0	0.0368 -:=:=:	0.80×10^{-4}	
800.0	0.0200	0.360 × 10	
10,000	0		

In this reaction, the concentration of butyl chloride, C₄H₉Cl, was measured at various times.

© 2012 Pearson Education, In

$$C_4H_9CI(aq) + H_2O(I) \longrightarrow C_4H_9OH(aq) + HCI(aq)$$

TABLE 14.1 •	Rate Data for B	Reaction of C ₄ H	Cl with Water
	riute Duta ioi i	icuction of Carr	got with water

Time, $t(s)$	$[C_4H_9Cl](M)$	Average Rate (M/s)
0.0	0.1000	1.9×10^{-4}
50.0	0.0905 -:=:=:	
100.0	0.0820 -:=:	1.7×10^{-4}
150.0	0.0741 -:=:=:	1.6×10^{-4}
200.0	0.0671 -:=:=:	1.4×10^{-4}
300.0	0.0549 -:=:=:	1.22×10^{-4}
400.0	0.0448 -:=:	1.01×10^{-4}
500.0	0.0368	$2:2:= 0.80 \times 10^{-4}$
800.0	0.0200	0.560×10^{-4}
10,000	0	

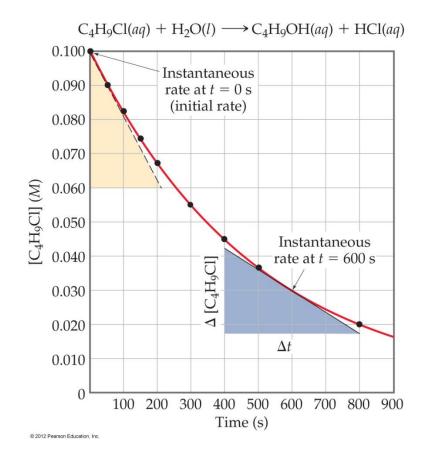
The average rate of the reaction over each interval is the change in concentration divided by the change in time:

© 2012 Pearson Education, Inc.

Average rate =
$$\frac{\Delta[C_4H_9CI]}{\Delta t}$$

$$C_4H_9CI(aq) + H_2O(I) \longrightarrow C_4H_9OH(aq) + HCI(aq)$$

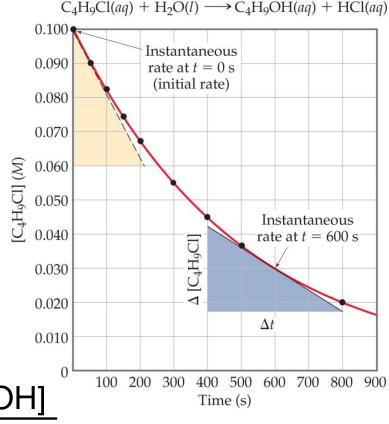
- A plot of [C₄H₉Cl]
 versus time for this
 reaction yields a curve
 like this.
- The slope of a line tangent to the curve at any point is the instantaneous rate at that time.



12

$$C_4H_9CI(aq) + H_2O(I) \longrightarrow C_4H_9OH(aq) + HCI(aq)$$

- In this reaction, the ratio of C₄H₉Cl to C₄H₉OH is 1:1.
- Thus, the rate of disappearance of C₄H₉Cl is the same as the rate of appearance of C₄H₉OH.



Rate =
$$\frac{-\Delta[C_4H_9CI]}{\Delta t} = \frac{\Delta[C_4H_9OH]}{\Delta t}$$

Reaction Rates and Stoichiometry

• What if the ratio is *not* 1:1?

$$2 HI(g) \longrightarrow H_2(g) + I_2(g)$$

In such a case,

Rate =
$$-\frac{1}{2} \frac{\Delta[HI]}{\Delta t} = \frac{\Delta[I_2]}{\Delta t}$$

Reaction Rates and Stoichiometry

To generalize, then, for the reaction

$$aA + bB \longrightarrow cC + dD$$

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

Rate Law

One can gain information about the rate of a reaction by seeing how the rate changes with changes in concentration.

$$aA + bB \longrightarrow cC + dD$$

Rate =
$$-\frac{1}{a} \frac{\Delta[A]}{\Delta t} = k[A]^m [B]^n$$

- *m* and *n* must be determined experimentally!
- m or n = 0 (zero order), 1 (first order), 2 (second order)
- Overall reaction order = m + n

Types of Rate Laws

- Differential Rate Law (rate law) shows how the rate of a reaction depends on concentrations.
 - Reaction rate vs. concentration
 - Rate = k [A]ⁿ[B]^m
- Integrated Rate Law shows how the concentrations of species in the reaction depend on time.
 - Concentration vs. time

$$2 NO_2(g) \rightarrow 2 NO(g) + O_2(g)$$

Rate =
$$-\frac{\Delta[NO_2]}{\Delta t} = k [NO_2]^n$$

Time interval is infinite small

$$-\frac{d[NO_2]}{dt} = k [NO_2]^n$$

$$\int_0^t \frac{d[NO_2]}{[NO_2]^n} = \int_0^t k \, dt$$

Integrated Rate Laws

For a first-order process

$$\frac{\mathsf{d}[\mathsf{A}]}{\mathsf{d}t} = -k[\mathsf{A}]$$

Using calculus to integrate the rate law gives us

$$\ln \frac{[A]_t}{[A]_0} = -kt$$

where

[A]₀ is the initial concentration of A, and

 $[A]_t$ is the concentration of A at a certain time, t, during the course of the reaction.

Integrated Rate Laws

Manipulating this equation produces

$$\ln \frac{[A]_t}{[A]_0} = -kt$$

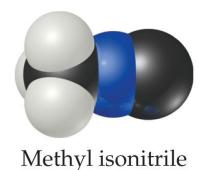
$$\ln [A]_t - \ln [A]_0 = -kt$$

$$\ln [A]_t = -kt + \ln [A]_0$$

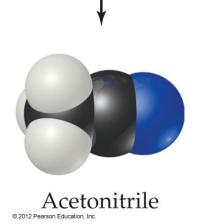
$$\ln [A]_t = -kt + \ln [A]_0$$

$$y = mx + b$$

Therefore, if a reaction is first-order, a plot of In[A] vs. t will yield a straight line, and the slope of the line will be -k.



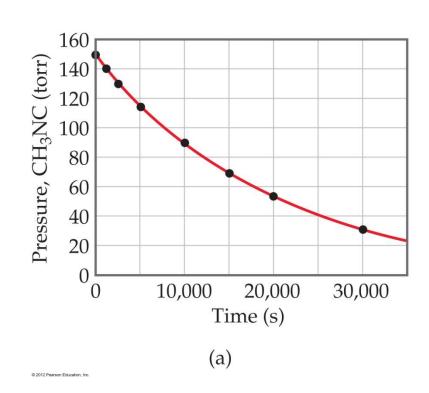
Consider the process in which methyl isonitrile is converted to acetonitrile.



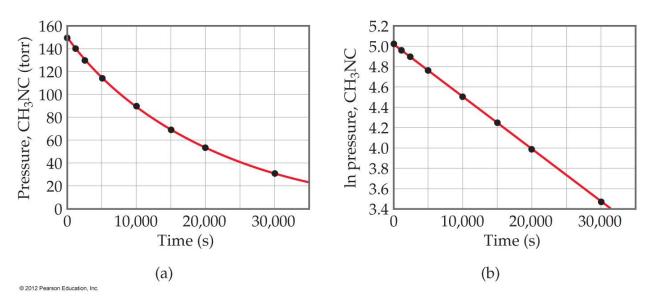
CH₃NC — CH₃CN

$$CH_3NC \longrightarrow CH_3CN$$

This data were collected for this reaction at 198.9 °C.



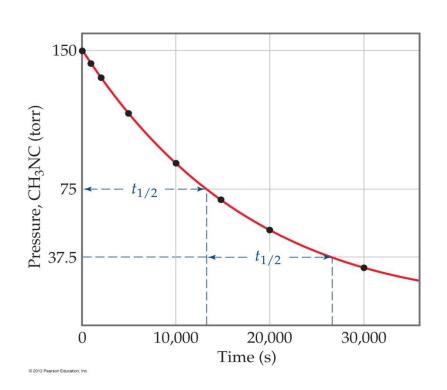
22



- When In P is plotted as a function of time, a straight line results.
 Therefore In [A]_t = -kt + In [A]₀
- · Therefore,
 - The process is first-order.
 - k is the negative of the slope: 5.1×10^{-5} s⁻¹.

Half-Life $(t_{1/2})$: $[A]_t = 0.5 [A]_0$

For a first-order process $\ln [A]_t = -kt + \ln [A]_0$



In
$$\frac{0.5 \text{ [A]}_0}{\text{[A]}_0} = -kt_{1/2}$$
In $0.5 = -kt_{1/2}$

$$-0.693 = -kt_{1/2}$$

$$\frac{0.693}{k} = t_{1/2}$$

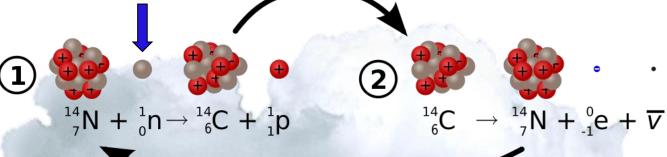
Note: For a first-order process, the half-life does not depend on [A]₀.

Section 21.4



Radiometric Dating

Discovered in 1939 by Serge Korff



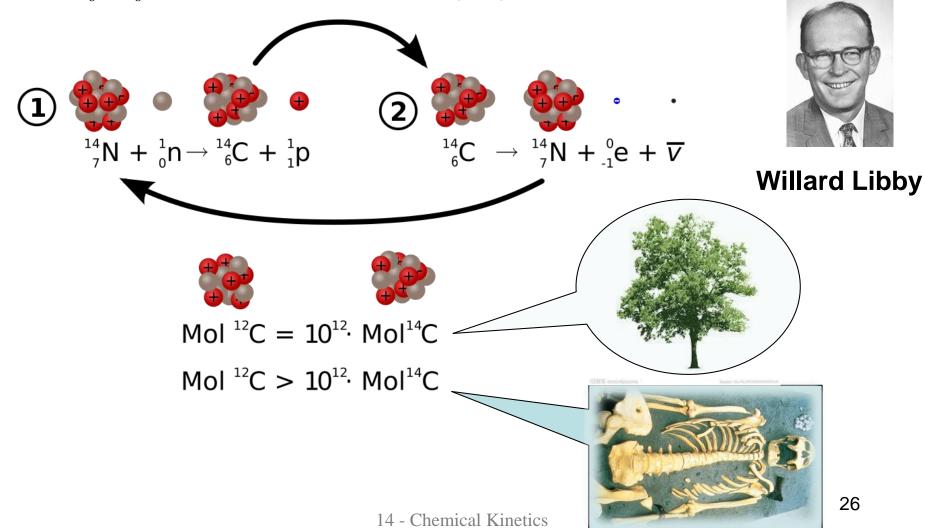
1: Formation of ¹⁴C

2: Decay of ¹⁴C

 $t_{1/2} = 5,730 \pm 40 \text{ years}$

Radiometric Dating

¹⁴₆C/¹²₆C remains constant for living organisms



Radiometric Dating

* $CO_2 \rightarrow$ *Organic molecule in plants \rightarrow *animals

Age determination from ¹⁴C/¹²C ratio

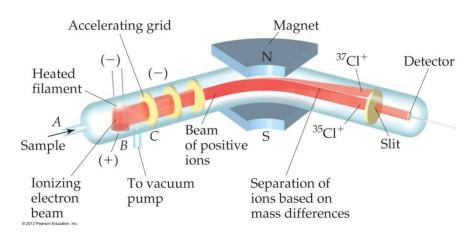
← decay ← died



Willard Libby Nobel Prize 1960

Checked with tree growth: accurate within 10%

Is it a must to use a sensitive radiation detector as proposed by Libby? i.e. how do we separate ¹⁴C from ¹²C?



Rock Age

A rock with 0.115 mg ²⁰⁶Pb / 1.000 mg ²³⁸U

Assuming ²⁰⁶Pb is coming from ²³⁸U

$$^{238}U \rightarrow ^{206}Pb$$

$$t_{1/2} = 4.5 \times 10^9 \text{ yr}$$



Hudson Bay, Canada

The original amount of ²³⁸U was

$$\left(0.115 \times \frac{238}{206}\right) + 1.000 = 1.133 \text{ mg}$$

The amount decayed

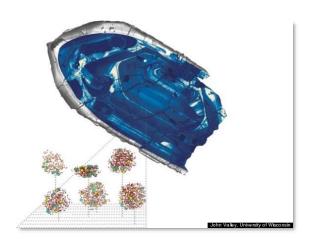
$$\ln \frac{N}{N_0} = -kt$$
 $\implies \ln \left(\frac{1.000}{1.133} \right) = -\left(\frac{0.693}{4.5 \times 10^9 \text{ yr}} \right) t$ $\implies t = 8.1 \times 10^8 \text{ yr}$

Rock Age

Oldest rock to date: $4.375 \pm 0.006 \times 10^9$ yr (Western Australia)

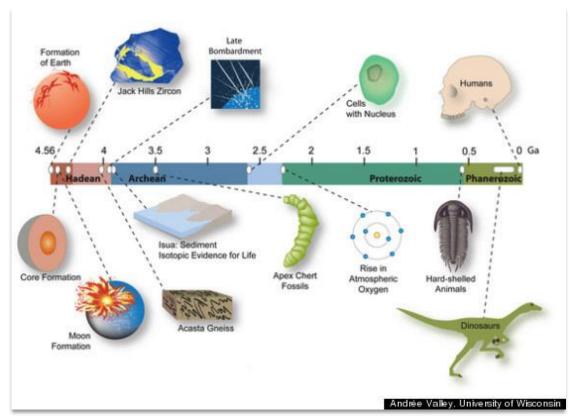
Cooling time for earth surface: $1-1.5 \times 10^9$ yr

Age of earth at about: $4.0-4.5 \times 10^9$ yr



10-nm clusters of radiogenic atoms of lead embedded in a crystal of zirconium silicate

Nature Geoscience (Feb 23, 2014)



Second-Order Processes

$$\frac{\mathsf{d}[\mathsf{A}]}{\mathsf{d}t} = -k[\mathsf{A}]^2$$

Similarly, integrating the rate law for a process that is second-order in reactant A, we get

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$
 Integrated rate law
$$y = mx + b$$

If a process is second-order in A, a plot of $1/[A]_t$ vs. t yields a straight line, and the slope of that line is k.

Second-Order Processes

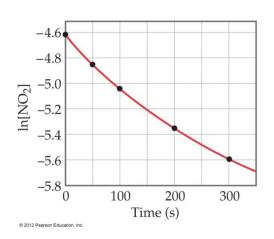
The decomposition of NO₂ at 300 °C is described by the equation

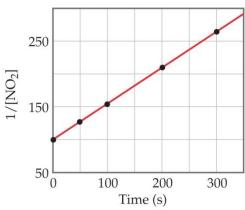
$$NO_2(g) \longrightarrow NO(g) + \frac{1}{2}O_2(g)$$

and yields data comparable to this table:

Time (s)	[NO ₂], <i>M</i>
0.0	0.01000
50.0	0.00787
100.0	0.00649
200.0	0.00481
300.0	0.00380

Second-Order Processes





•	Graphing	$\frac{1}{[NO_2]}$ vs. t
	gives a sti	raight line

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

Time (s)	[NO ₂], M	1/[NO ₂]
0.0	0.01000	100
50.0	0.00787	127
100.0	0.00649	154
200.0	0.00481	208
300.0	0.00380	263

 The process is second-order in [NO₂].

Summary of the Rate Laws

Summary of the Kinetics for Reactions of the Type $aA \longrightarrow Products$ That Are Zero, First, or Second Order in [A]

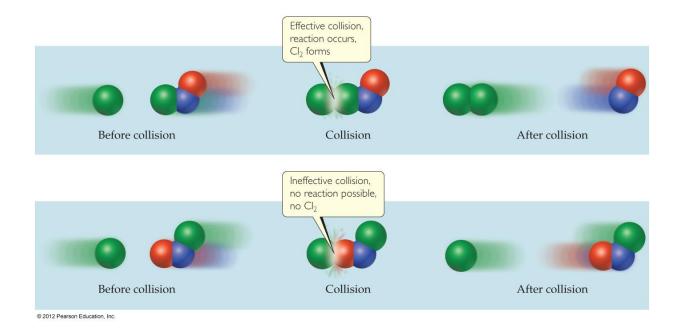
	Order		
	Zero	First	Second
Rate law	Rate = k	Rate = $k[A]$	$Rate = k[A]^2$
Integrated rate law	$[\mathbf{A}] = -kt + [\mathbf{A}]_0$	$\ln[\mathbf{A}] = -kt + \ln[\mathbf{A}]_0$	$\frac{1}{[\mathbf{A}]} = kt + \frac{1}{[\mathbf{A}]_0}$
Plot needed to give a straight line	[A] versus t	ln[A] versus t	$\frac{1}{[A]}$ versus t
Relationship of rate constant to the slope of the straight line	Slope = $-k$	Slope = $-k$	Slope = k
Half-life	$t_{1/2} = \frac{[\mathbf{A}]_0}{2k}$	$t_{1/2} = \frac{0.693}{k}$	$t_{1/2} = \frac{1}{k[\mathbf{A}]_0}$

The Collision Model

- In a chemical reaction, <u>bonds</u> are <u>broken</u> and new bonds are <u>formed</u>.
- Molecules can only react if they collide with each other.

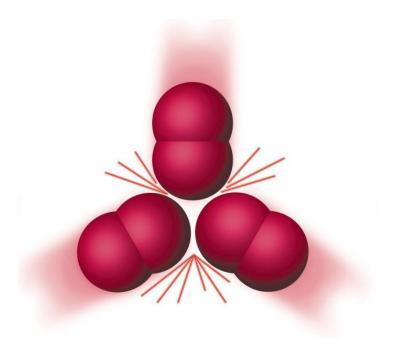
The Collision Model

Furthermore, molecules must collide with the correct **orientation** and with enough **energy** to cause bond breakage and formation.



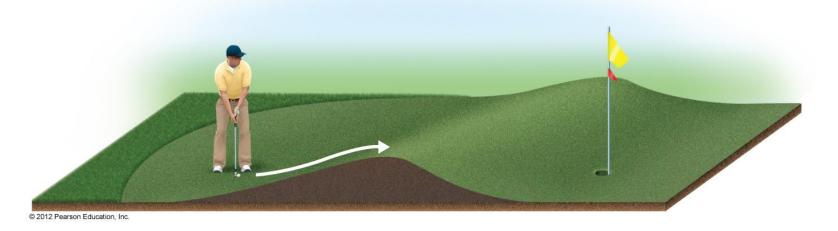
The Collision Model

Termolecular processes are very unlikely to occur!



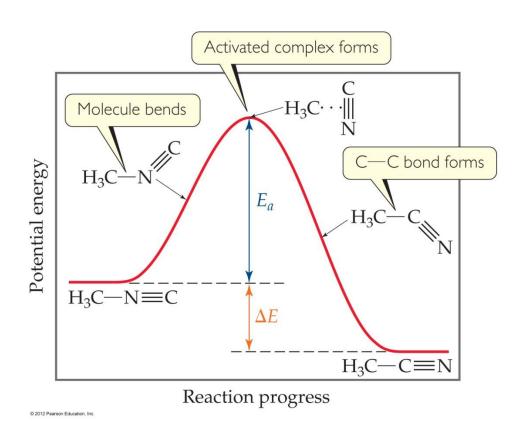
Activation Energy

- There is a minimum amount of energy required for reaction: the **activation energy**, E_a .
- Just as a ball cannot get over a hill if it does not roll up the hill with enough energy, a reaction cannot occur unless the molecules possess sufficient energy to get over the activation-energy barrier.



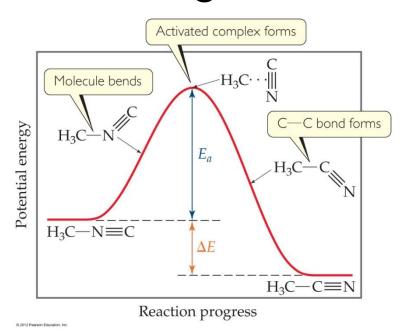
Reaction Coordinate Diagrams

It is helpful to visualize energy changes throughout a process on a reaction coordinate diagram like this one for the rearrangement of methyl isonitrile.



Reaction Coordinate Diagrams

- The diagram shows the energy of the reactants and products (and, therefore, ΔE).
- The high point on the diagram is the transition state.

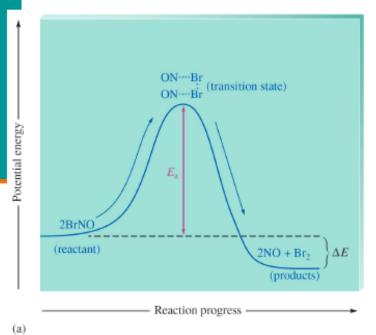


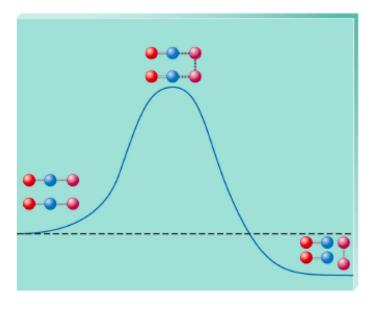
- The species present at the transition state is called the activated complex.
- The energy gap between the reactants and the activated complex is the activation-energy barrier.

Activation Energy

The activation energy

 Is the minimum energy needed for a reaction to take place upon proper collision of reactants.





(b)

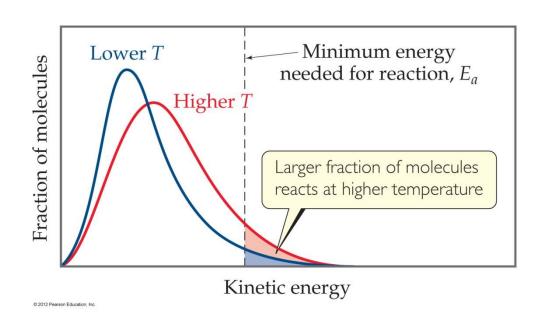


GIVE IT SOME THOUGHT

Suppose we have two reactions, A \longrightarrow B and B \longrightarrow C. You can isolate B, and it is stable. Is B the transition state for the reaction A \longrightarrow C?

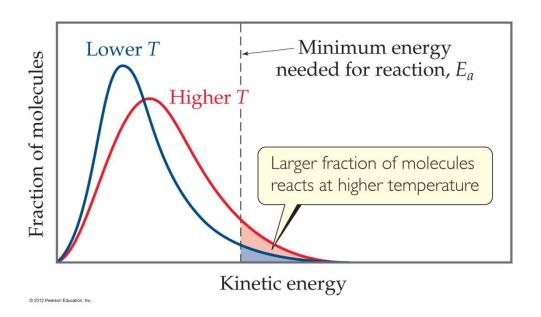
- A. Yes, because B is an intermediate.
- B. Yes, because B can be isolated.
- C. No, because B is not stable.
- D. No, because B can be isolated and transition states are by definition not stable.

Maxwell-Boltzmann Distributions



- Temperature is defined as a measure of the average kinetic energy of the molecules in a sample.
- At any temperature there is a wide distribution of kinetic energies.

Maxwell-Boltzmann Distributions



 So more molecules can overcome the activationenergy barrier.

- As the temperature increases, the curve flattens and broadens.
- Thus, at higher temperatures, a larger population of molecules has higher energy.

Arrhenius Equation

Svante Arrhenius developed a mathematical relationship between k and E_a :

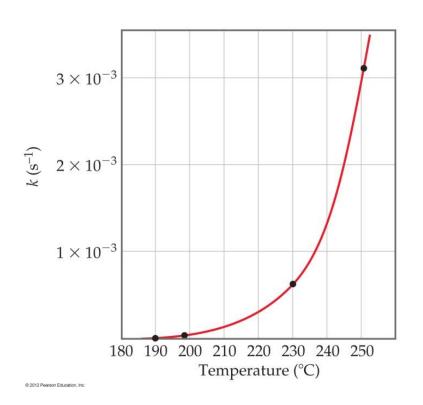
Thermal energy

$$k = Ae^{-E_a/RT}$$

$$E = \frac{3}{2}RT$$

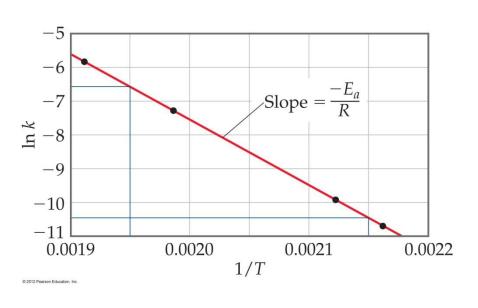
- where A is the frequency factor, a number that represents the likelihood that collisions would occur with the proper orientation for reaction.
- The term e^{-E_a/RT} represents the fraction of molecular collisions sufficiently energetic to produce a reaction.

Temperature and Rate



- Generally, as temperature increases, so does the reaction rate.
- This is because k is temperature-dependent.

Arrhenius Equation



Taking the natural logarithm of both sides, the equation becomes

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

$$y = mx + b$$

Therefore, if k is determined experimentally at several temperatures, E_a can be calculated from the slope of a plot of $\ln k$ vs. $\frac{1}{T}$.

Reaction Mechanisms

The sequence of events that describes the actual process by which reactants become products is called the **reaction mechanism**.

Reaction Mechanisms

- Reactions may occur all at once or through several discrete steps.
- Each of these processes is known as an elementary reaction or elementary process.

Reaction Mechanisms

TABLE 14.3 • Elementary Reactions and Their Rate Laws		
Molecularity	Elementary Reaction	Rate Law
<i>Uni</i> molecular	$A \longrightarrow products$	Rate = k[A]
<i>Bi</i> molecular	$A + A \longrightarrow products$	$Rate = k[A]^2$
<i>Bi</i> molecular	$A + B \longrightarrow products$	Rate = k[A][B]
<i>Ter</i> molecular	$A + A + A \longrightarrow products$	$Rate = k[A]^3$
<i>Ter</i> molecular	$A + A + B \longrightarrow products$	$Rate = k[A]^2[B]$

© 2012 Pearson Education, Inc.

Not likely

 $A + B + C \longrightarrow products$

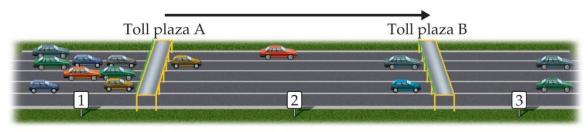
*Ter*molecular

The **molecularity** of a process tells how many molecules are involved in the process.

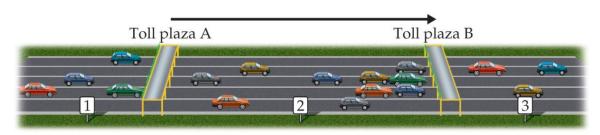
Rate = k[A][B][C]

Multistep Mechanisms

- In a multistep process, one of the steps will be slower than all others.
- The overall reaction cannot occur faster than this slowest, rate-determining step.



(a) Cars slowed at toll plaza A, rate-determining step is passage through A



(b) Cars slowed at toll plaza B, rate-determining step is passage through B

© 2012 Pearson Education, Inc.

Slow Initial Step

$$NO_2(g) + CO(g) \longrightarrow NO(g) + CO_2(g)$$

 The rate law for this reaction is found experimentally to be

Rate =
$$k[NO_2]^2$$

- CO is necessary for this reaction to occur, but the rate of the reaction does not depend on its concentration.
- This suggests that the reaction occurs in two steps.

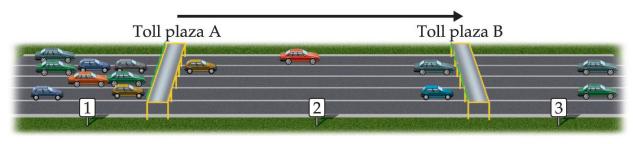
Slow Initial Step

A proposed mechanism for this reaction is

Step 1:
$$NO_2 + NO_2 \longrightarrow NO_3 + NO$$
 (slow)

Step 2:
$$NO_3 + CO \longrightarrow NO_2 + CO_2$$
 (fast)

- The NO₃ intermediate is consumed in the second step.
- As CO is not involved in the slow, rate-determining step, it does not appear in the rate law.



(a) Cars slowed at toll plaza A, rate-determining step is passage through A

$$2 \text{ NO}(g) + \text{Br}_2(g) \longrightarrow 2 \text{ NOBr}(g)$$

The rate law for this reaction is found to be

Rate =
$$k[NO]^2[Br_2]$$

 Because termolecular processes are rare, this rate law suggests a two-step mechanism.

A proposed mechanism is

Step 1: NO + Br₂
$$\longrightarrow$$
 NOBr₂ (fast)

Step 2:
$$NOBr_2 + NO \longrightarrow 2 NOBr$$
 (slow)

Step 1 includes the forward *and* reverse (backward) reactions.

- Step 1: $NO + Br_2 \longrightarrow NOBr_2$ (fast)
- Step 2: $NOBr_2 + NO \longrightarrow 2 NOBr$ (slow)
- The rate of the overall reaction depends upon the rate of the slow step.
- The rate law for that step would be

Rate =
$$k_2$$
[NOBr₂] [NO]

But how can we find [NOBr₂]?

Step 1:
$$NO + Br_2 \longrightarrow NOBr_2$$
 (fast)

Step 2:
$$NOBr_2 + NO \longrightarrow 2 NOBr$$
 (slow)

- NOBr₂ can react two ways:
 - With NO to form NOBr.
 - By decomposition to reform NO and Br₂.
- The reactants and products of the first step are in equilibrium with each other.
- Therefore,

$$Rate_f = Rate_r$$

This is the "fast equilibrium" approximation.

Step 1:
$$NO + Br_2 \longrightarrow NOBr_2$$
 (fast)

Step 2:
$$NOBr_2 + NO \longrightarrow 2 NOBr$$
 (slow)

• Because $Rate_f = Rate_r$

$$k_1[NO][Br_2] = k_{-1}[NOBr_2]$$

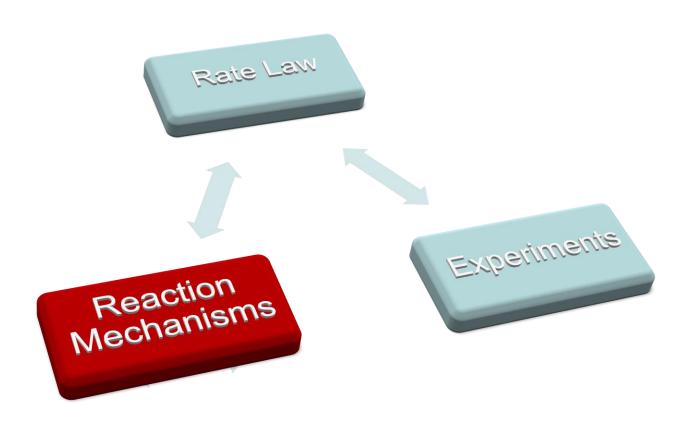
Solving for [NOBr₂], gives us

$$\frac{k_1}{k_{-1}}$$
[NO] [Br₂] = [NOBr₂]

Substituting this expression for [NOBr₂] in the rate law for the rate-determining step, gives

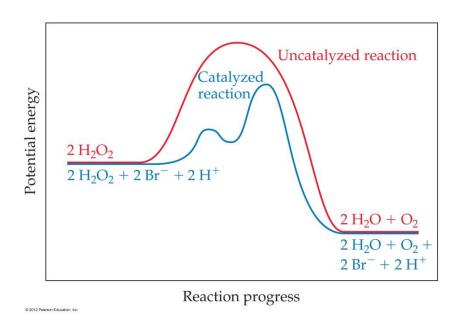
Rate =
$$\frac{k_2 k_1}{k_{-1}}$$
 [NO] [Br₂] [NO]
= k [NO]² [Br₂]

Deduce Reaction Mechanism



Catalysts

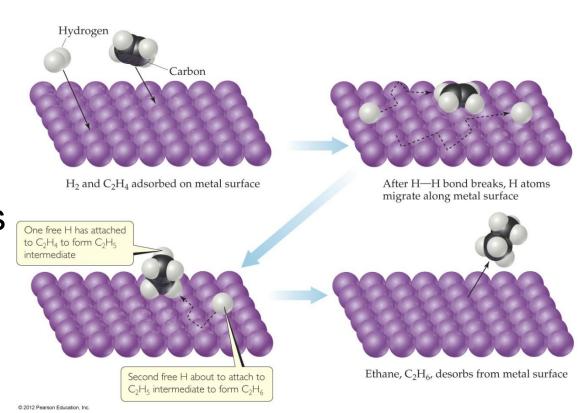
- Catalysts increase the rate of a reaction by decreasing the activation energy of the reaction.
- Catalysts change the mechanism by which the process occurs.



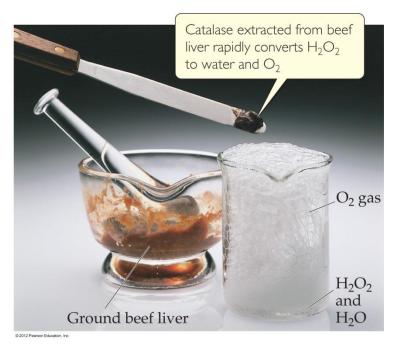
60

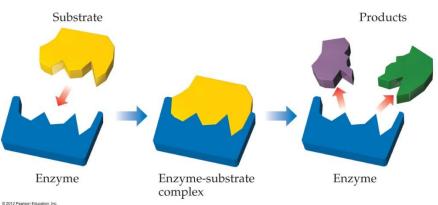
Catalysts

One way a catalyst can speed up a reaction is by holding the reactants together and helping bonds to break.



Enzymes

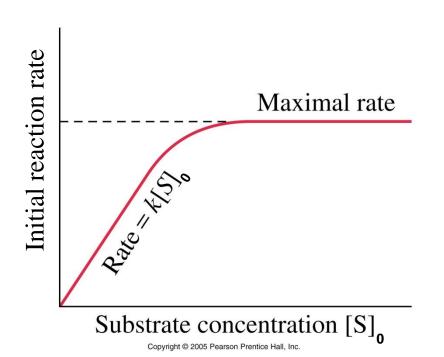




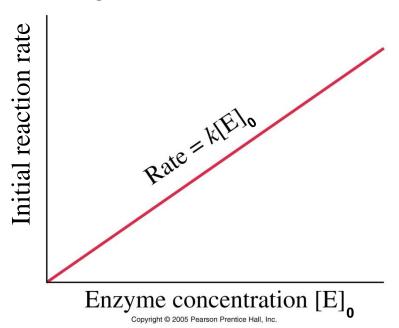
- Enzymes are catalysts in biological systems → ~ 10¹⁶ times faster
- The substrate fits into the active site of the enzyme much like a key fits into a lock.

Factors Influencing Enzyme Activity

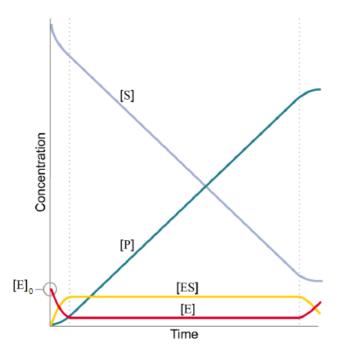
The rates of enzyme-catalyzed reactions are influenced by factors such as concentration of the substrate, concentration of the enzyme, acidity of the medium, and temperature.



At high substrate concentrations:



$$E + S \xrightarrow{k_1 \atop k_2} ES \xrightarrow{k_3} E + P$$



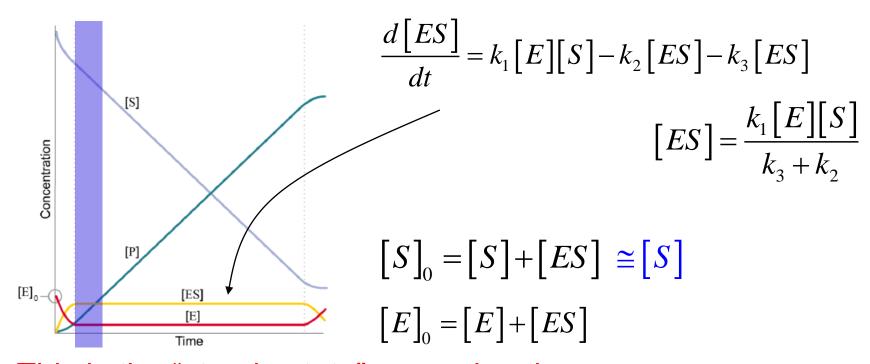
$$\frac{d[ES]}{dt} = k_1[E][S] - k_2[ES] - k_3[ES]$$

$$\frac{d[P]}{dt} = k_3[ES]$$

$$\frac{d[S]}{dt} = -k_1[E][S] + k_2[ES]$$

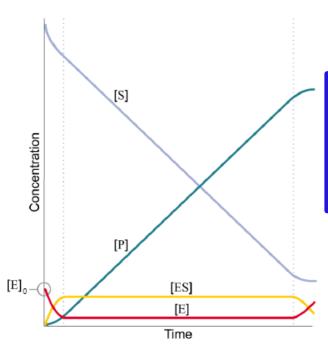
Numerical simulations

$$E + S \xrightarrow{k_1 \atop k_2} ES \xrightarrow{k_3} E + P$$



This is the "steady-state" approximation.

$$E + S \xrightarrow{k_1 \atop k_2} ES \xrightarrow{k_3} E + P$$



$$[ES] = \frac{k_1 \{ [E]_0 - [ES] \} [S]_0}{k_3 + k_2}$$

$$[ES] = \frac{[E]_0}{1 + \frac{k_3 + k_2}{k_1 [S]_0}}$$

$$[ES] = \frac{k_1 \lfloor E \rfloor \lfloor S \rfloor}{k_3 + k_2}$$

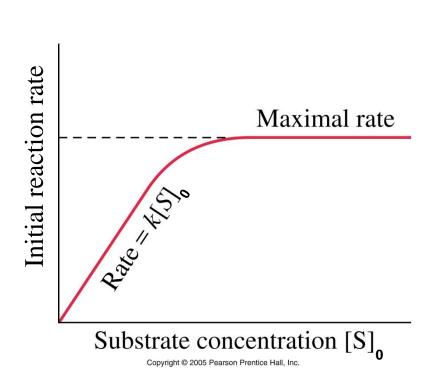
$$[S]_0 = [S] + [ES] \cong [S]$$

$$[E]_0 = [E] + [ES]$$

$$E + S \xrightarrow{k_{1} \atop k_{2}} ES \xrightarrow{k_{3}} E + P$$

$$v_{0} = \frac{d[P]}{dt} = k_{3}[ES] \qquad v_{0} = k_{3} \frac{[E]_{0}}{1 + \frac{k_{3} + k_{2}}{k_{1}[S]_{0}}} = \frac{k_{3}[E]_{0}}{1 + \frac{K_{M}}{[S]_{0}}}$$

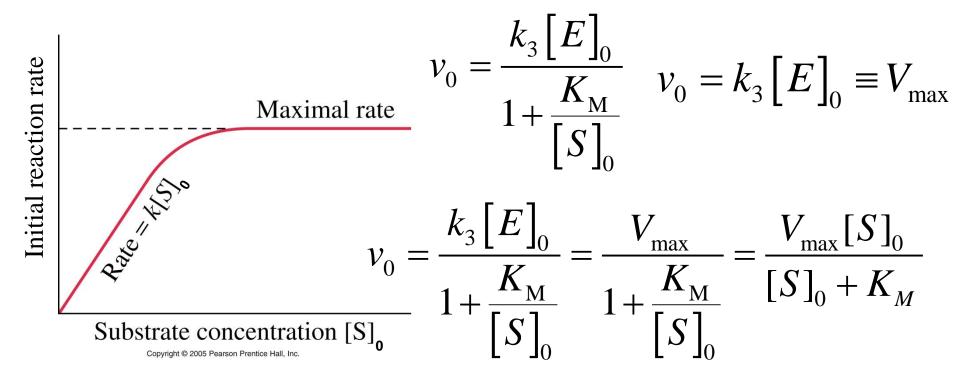
$$E + S \xrightarrow{k_1 \atop k_2} ES \xrightarrow{k_3} E + P$$



$$v_0 = \frac{k_3 [E]_0}{1 + \frac{K_M}{[S]_0}}$$

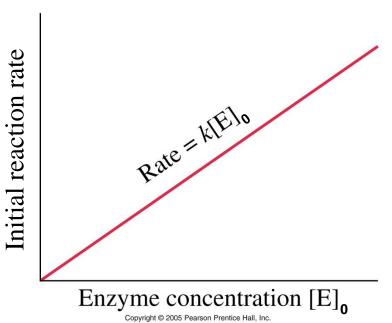
$$v_0 = \frac{d[P]}{dt} = k_3[ES]$$

$$E + S \xrightarrow{k_1} ES \xrightarrow{k_2} ES \xrightarrow{k_3} E + P$$



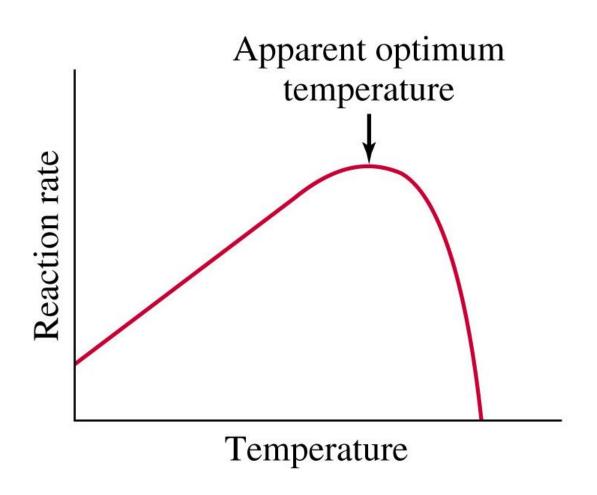
$$E + S \xrightarrow{k_1 \atop k_2} ES \xrightarrow{k_3} E + P$$

At high substrate concentrations:

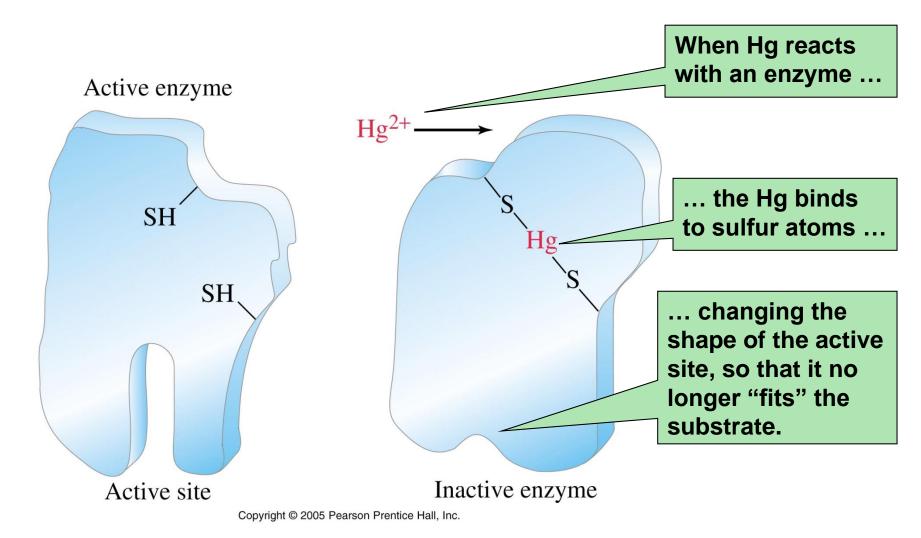


Si
$$v_0 = \frac{k_3 [E]_0}{1 + \frac{K_M}{[S]_0}}$$
 $v_0 = k_3 [E]_0 \equiv V_{\text{max}}$

Enzyme Activity as a Function of Temperature



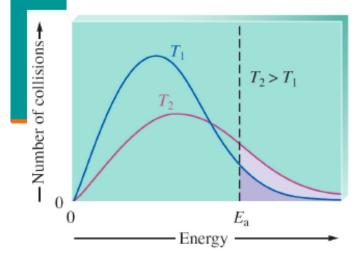
Enzyme Inhibition



Reaction Rate and Temperature

Reaction rate

 Increases when temperature rises because reacting molecules move faster, providing more colliding molecules with energy higher than E_a.



Number of collisions with E_a

= (total number of collisions)e-Ea/RT

E_a: activation energy

R: universal gas constant

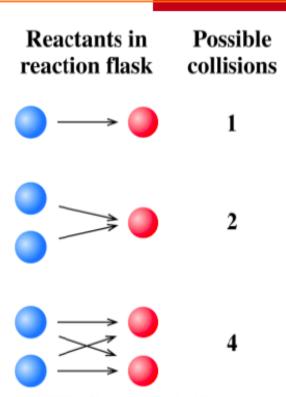
T: temperature in K

Reaction Rate and Concentration

Increasing the

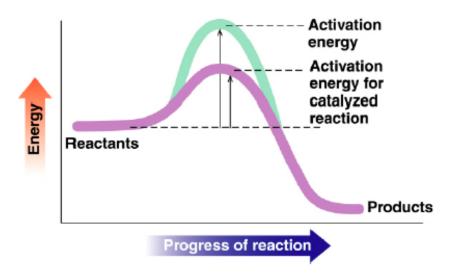
concentration of reactants

- Increases the number of collisions.
- Increases the reaction rate.



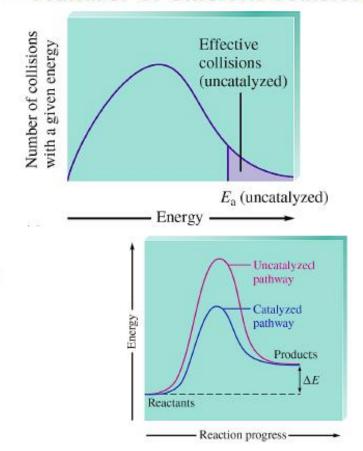
Catalyst

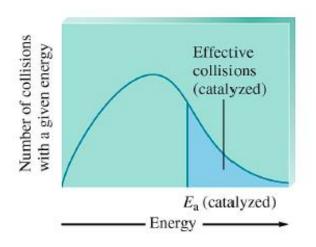
- A substance that speeds up a reaction without being consumed itself.
- Provides a new pathway for the reaction with a lower activation energy.



Effect of Catalysts

Number of efficient collisions





Both forward and backward reactions are accelerated.

Factors That Increase Reaction Rate

Factor	Reason
More reactants	More collisions
Higher temperature	More collisions, more collision with energy of activation
Adding a catalyst	Lowers energy of activation

