

Lecture Presentation

Chapter 13

Chemical Kinetics

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Reaction Rates

Chemical Kinetics: The area of chemistry concerned with reaction rates and the sequence of steps by which reactions occur

Reaction Rate: Either the *increase* in the concentration of a product per unit time or the *decrease* in the concentration of a reactant per unit time

Reaction Rates

TABLE 13.1 Concentrations as a Function of Time at 55 °C for the Reaction
 $2 \text{N}_2\text{O}_5(g) \longrightarrow 4 \text{NO}_2(g) + \text{O}_2(g)$

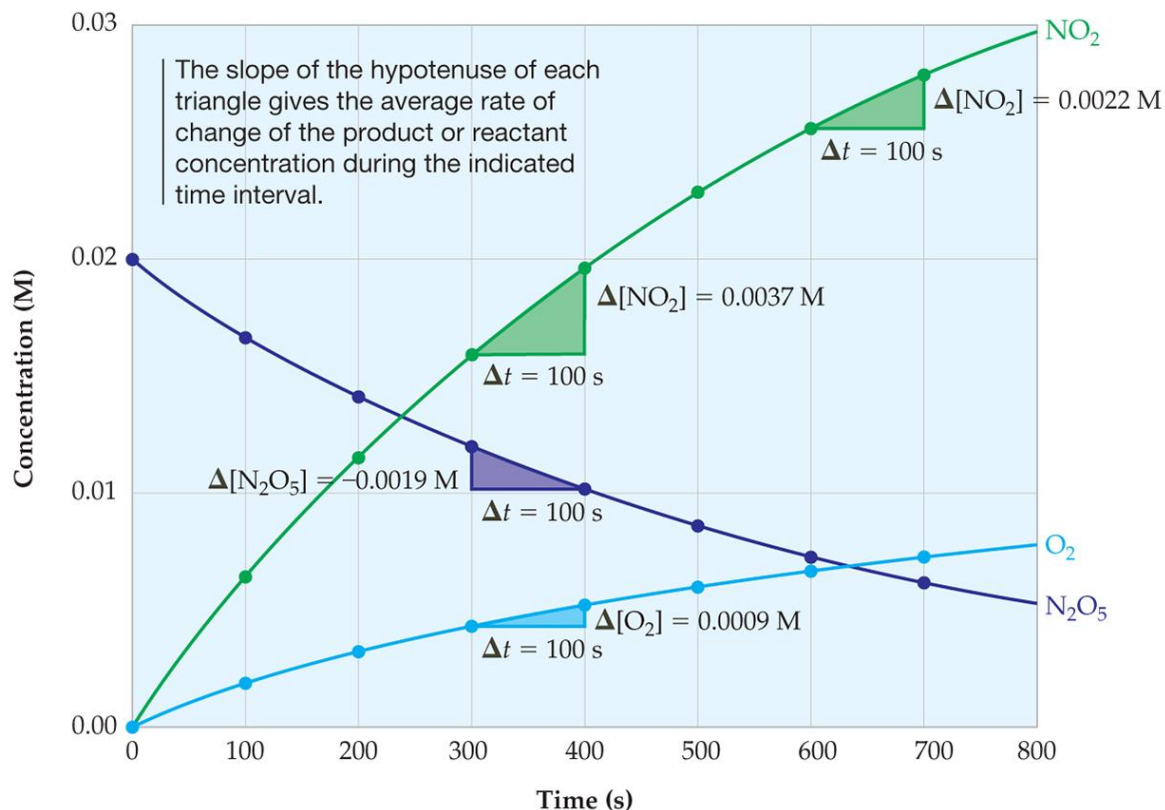
Time (s)	Concentration (M)		
	N_2O_5	NO_2	O_2
0	0.0200	0	0
100	0.0169	0.0063	0.0016
200	0.0142	0.0115	0.0029
300	0.0120	0.0160	0.0040
400	0.0101	0.0197	0.0049
500	0.0086	0.0229	0.0057
600	0.0072	0.0256	0.0064
700	0.0061	0.0278	0.0070

Note that the concentrations of NO_2 and O_2 increase as the concentration of N_2O_5 decreases.

Reaction Rates



The concentrations of O_2 and NO_2 increase as the concentration of N_2O_5 decreases.



The rate of **formation of O_2** is one-fourth the rate of **formation of NO_2** and one-half the rate of **decomposition of N_2O_5** .

Reaction Rates



Rate of decomposition of N_2O_5 :

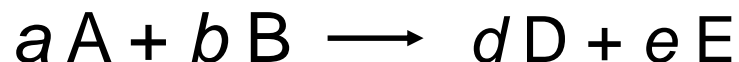
$$\begin{aligned} \frac{\Delta[\text{N}_2\text{O}_5]}{\Delta t} &= \frac{-(0.0101 \text{ M} - 0.0120 \text{ M})}{(400 \text{ s} - 300 \text{ s})} \\ &= 1.9 \times 10^{-5} \frac{\text{M}}{\text{s}} \end{aligned}$$

Reaction Rates



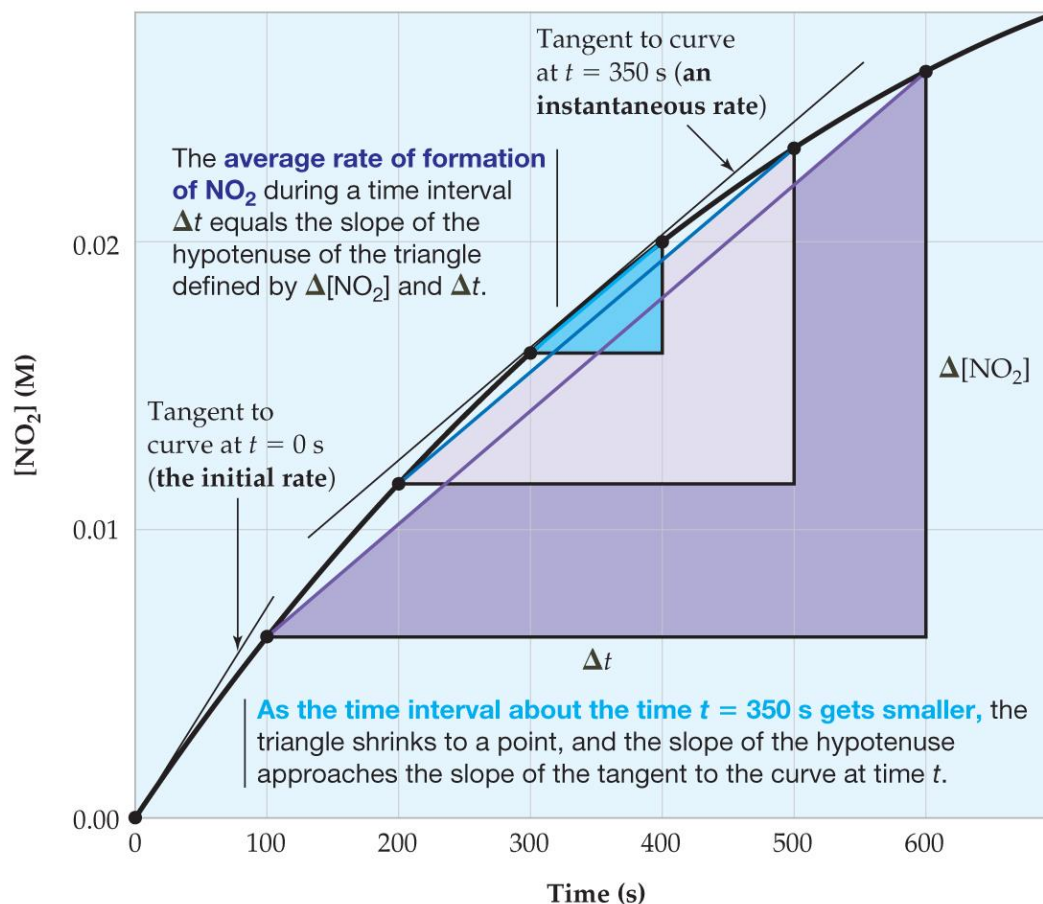
General rate of reaction:

$$\text{rate} = -\frac{1}{2} \frac{\Delta[\text{N}_2\text{O}_5]}{\Delta t} = \frac{1}{4} \frac{\Delta[\text{NO}_2]}{\Delta t} = \frac{\Delta[\text{O}_2]}{\Delta t}$$



$$\text{rate} = -\frac{1}{a} \frac{\Delta[\text{A}]}{\Delta t} = -\frac{1}{b} \frac{\Delta[\text{B}]}{\Delta t} = \frac{1}{d} \frac{\Delta[\text{D}]}{\Delta t} = \frac{1}{e} \frac{\Delta[\text{E}]}{\Delta t}$$

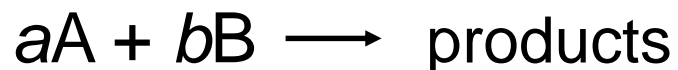
Reaction Rates



The slope of the tangent at time t is defined as the instantaneous rate at that particular time. The initial rate is the slope of the tangent to the curve at $t = 0$.

Rate Laws and Reaction Order

Rate Law: An equation that shows the dependence of the reaction rate on the concentration of each reactant



$$\text{rate} \propto [A]^m[B]^n$$

$$\text{rate} = k[A]^m[B]^n$$

k is the rate constant.

Rate Laws and Reaction Order

The values of the exponents in the rate law **must be determined by experiment**; they cannot be deduced from the stoichiometry of the reaction.

TABLE 13.2 Balanced Chemical Equations and Experimentally Determined Rate Laws for Some Reactions

Reaction*	Rate Law
$(\text{CH}_3)_3\text{CBr}(\text{soln}) + \text{H}_2\text{O}(\text{soln}) \longrightarrow (\text{CH}_3)_3\text{COH}(\text{soln}) + \text{H}^+(\text{soln}) + \text{Br}^-(\text{soln})$	Rate = $k[(\text{CH}_3)_3\text{CBr}]$
$\text{HCO}_2\text{H}(\text{aq}) + \text{Br}_2(\text{aq}) \longrightarrow 2 \text{H}^+(\text{aq}) + 2 \text{Br}^-(\text{aq}) + \text{CO}_2(\text{g})$	Rate = $k[\text{Br}_2]$
$\text{BrO}_3^-(\text{aq}) + 5 \text{Br}^-(\text{aq}) + 6 \text{H}^+(\text{aq}) \longrightarrow 3 \text{Br}_2(\text{aq}) + 3 \text{H}_2\text{O}(\text{l})$	Rate = $k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]^2$
$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \longrightarrow 2 \text{HI}(\text{g})$	Rate = $k[\text{H}_2][\text{I}_2]$

In general, the exponents in the rate law **are not the same** as the stoichiometric coefficients in the balanced chemical equation for the reaction.

*In the first reaction, “(soln)” denotes a nonaqueous solution.

Experimental Determination of a Rate Law



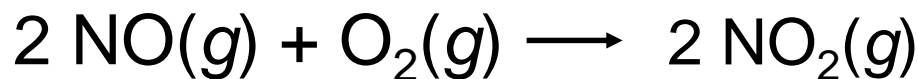
$$\text{rate} = k[\text{NO}]^m [\text{O}_2]^n$$

Compare the initial rates to the changes in initial concentrations.

TABLE 13.3 Initial Concentration and Rate Data for the Reaction
 $2 \text{NO}(g) + \text{O}_2(g) \longrightarrow 2 \text{NO}_2(g)$

Experiment	Initial [NO]	Initial [O ₂]	Initial Reaction Rate (M/s)
1	0.015	0.015	0.024
2	0.030	0.015	0.096
3	0.015	0.030	0.048
4	0.030	0.030	0.192

Experimental Determination of a Rate Law



$$\text{rate} = k[\text{NO}]^2 [\text{O}_2]^n$$

The concentration of NO **doubles**, the concentration of O₂ remains constant, and the rate **quadruples**.

$$2^m = 4 \quad m = 2$$

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Experimental Determination of a Rate Law



$$\text{rate} = k[\text{NO}]^2 [\text{O}_2]$$

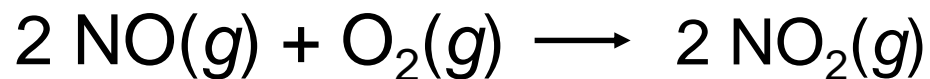
The concentration of O_2 **doubles**, the concentration of NO remains constant, and the rate **doubles**.

$$2^n = 2 \quad n = 1$$

TABLE 13.3 Initial Concentration and Rate Data for the Reaction
 $2 \text{NO}(g) + \text{O}_2(g) \longrightarrow 2 \text{NO}_2(g)$

Experiment	Initial [NO]	Initial [O ₂]	Initial Reaction Rate (M/s)
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Experimental Determination of a Rate Law



$$\text{rate} = k[\text{NO}]^2 [\text{O}_2]$$

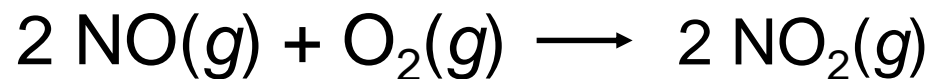
Reaction Order With Respect to a Reactant

- NO: second-order
- O₂: first-order

Overall Reaction Order

- 2 + 1 = 3 (third-order)

Experimental Determination of a Rate Law

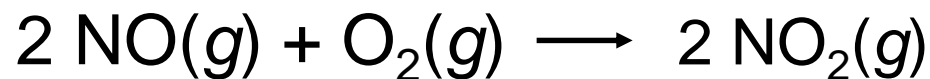


$$\text{rate} = k[\text{NO}]^2 [\text{O}_2]$$

Units of k for this third-order reaction:

$$k = \frac{\text{rate}}{[\text{NO}]^2 [\text{O}_2]} = \frac{\frac{\text{M}}{\text{s}}}{(\text{M}^2) (\text{M})} = \frac{1}{\text{M}^2 \text{s}}$$

Experimental Determination of a Rate Law

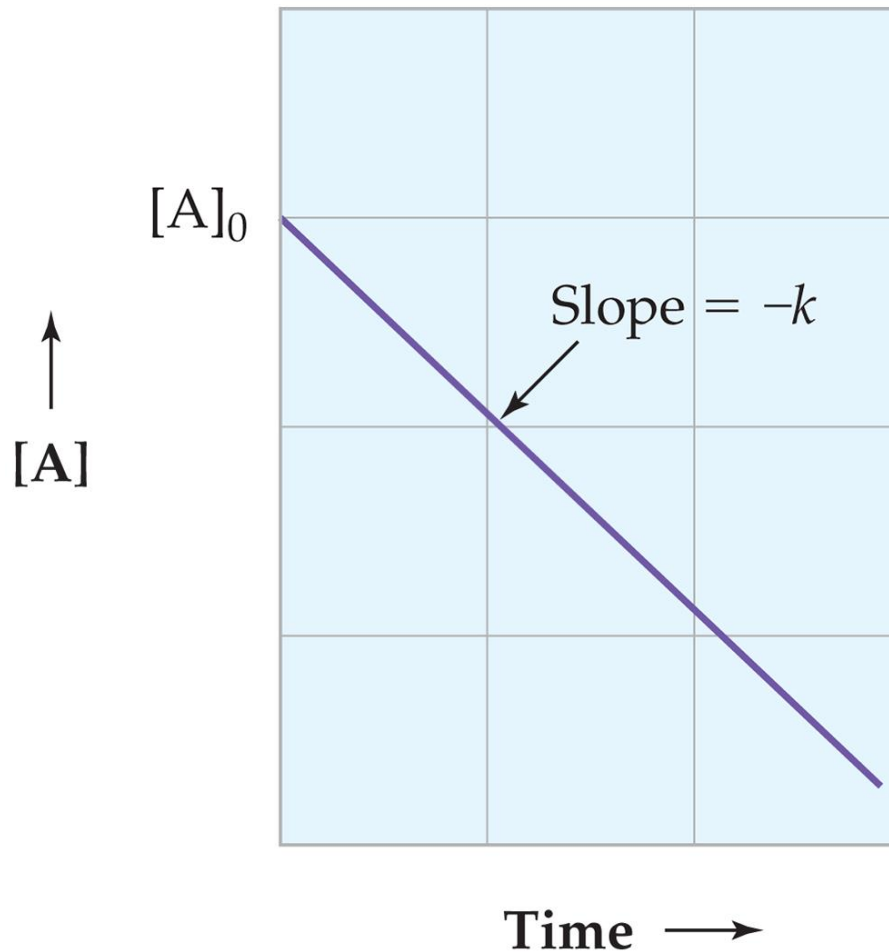


$$\text{rate} = k[\text{NO}]^2 [\text{O}_2]$$

Rate Law	Overall Reaction Order	Units for k
Rate = k	Zeroth order	M/s or M s^{-1}
Rate = $k[\text{A}]$	First order	1/s or s^{-1}
Rate = $k[\text{A}][\text{B}]$	Second order	$1/(\text{M} \cdot \text{s})$ or $\text{M}^{-1} \text{s}^{-1}$
Rate = $k[\text{A}][\text{B}]^2$	Third order	$1/(\text{M}^2 \cdot \text{s})$ or $\text{M}^{-2} \text{s}^{-1}$

Zeroth-Order Reactions

A plot of **[A]** versus **time** gives a straight-line fit and the **slope** will be **$-k$** .



Zeroth-Order Reactions

For a zeroth-order reaction, the rate is independent of the concentration of the reactant.



$$\text{rate} = k[A]^0 = k \qquad -\frac{\Delta[A]}{\Delta t} = k$$

Calculus can be used to derive an **integrated** rate law.

$$\boxed{[A]_t = -kt + [A]_0}$$

$y = mx + b$

$[A]_t$ concentration of A at time t

$[A]_0$ initial concentration of A

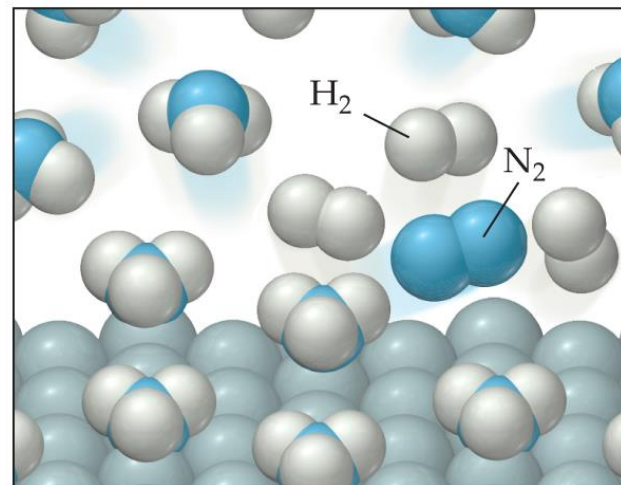
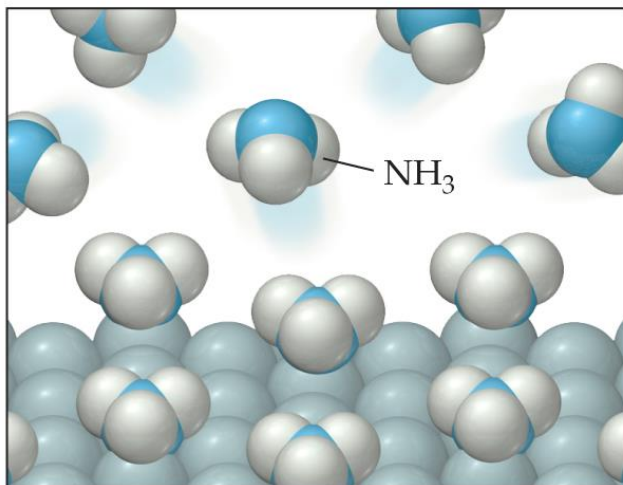
Zeroth-Order Reactions



$$\text{rate} = k[\text{NH}_3]^0 = k$$

Most of the NH_3 molecules are in the gas phase above the surface and are unable to react.

As NH_3 molecules on the surface decompose, they are replaced by molecules from the gas phase, so the number of NH_3 molecules on the surface remains constant.



Because only the NH_3 molecules on the surface react under these conditions, the reaction rate is independent of the total concentration of NH_3 .

Integrated Rate Law for a First-Order Reaction



$$\text{rate} = k[A] \qquad -\frac{\Delta[A]}{\Delta t} = k[A]$$

Calculus can be used to derive an **integrated** rate law.

$$\ln \left(\frac{[A]_t}{[A]_0} \right) = -kt \left\{ \begin{array}{l} [A]_t \quad \text{concentration of A at time } t \\ [A]_0 \quad \text{initial concentration of A} \end{array} \right.$$

Using: $\ln \left(\frac{x}{y} \right) = \ln(x) - \ln(y)$

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

Integrated Rate Law for a First-Order Reaction

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

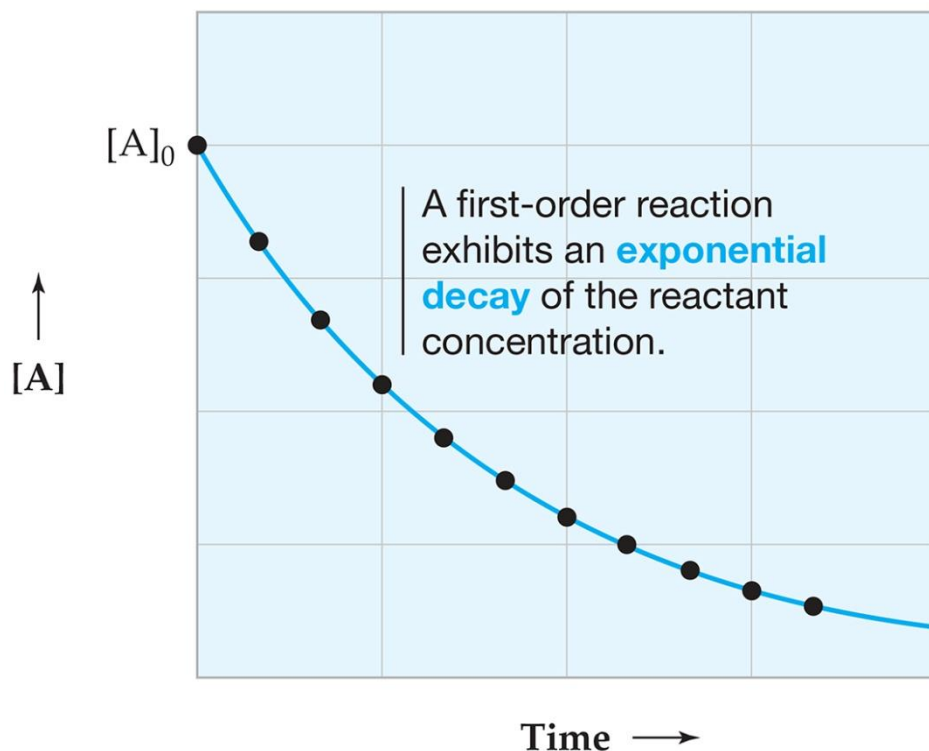
$$y = mx + b$$

A plot of $\ln[A]$ versus **time** gives a straight-line fit and the **slope** will be $-k$.

Integrated Rate Law for a First-Order Reaction

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

(a) Reactant concentration versus time



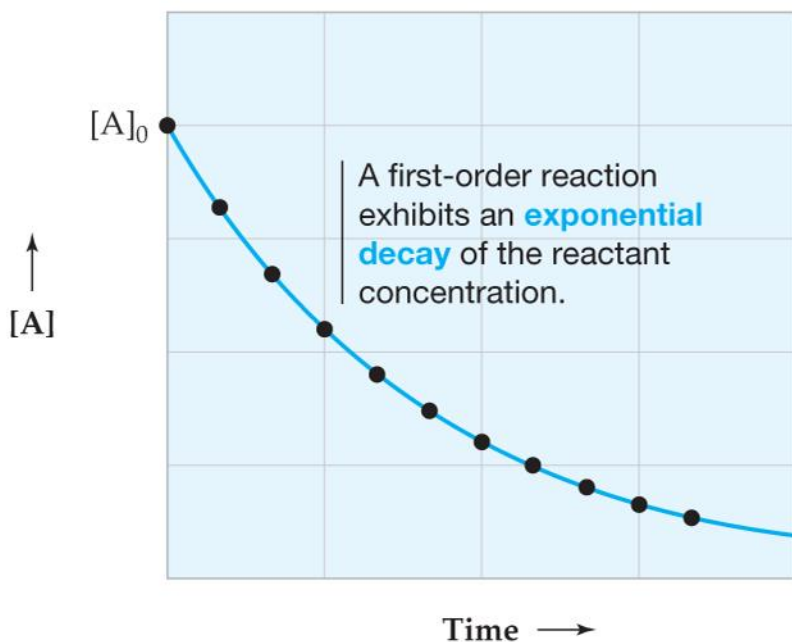
This is a plot of $[A]$ versus **time**.

The best-fit is a *curve* and not a *line*.

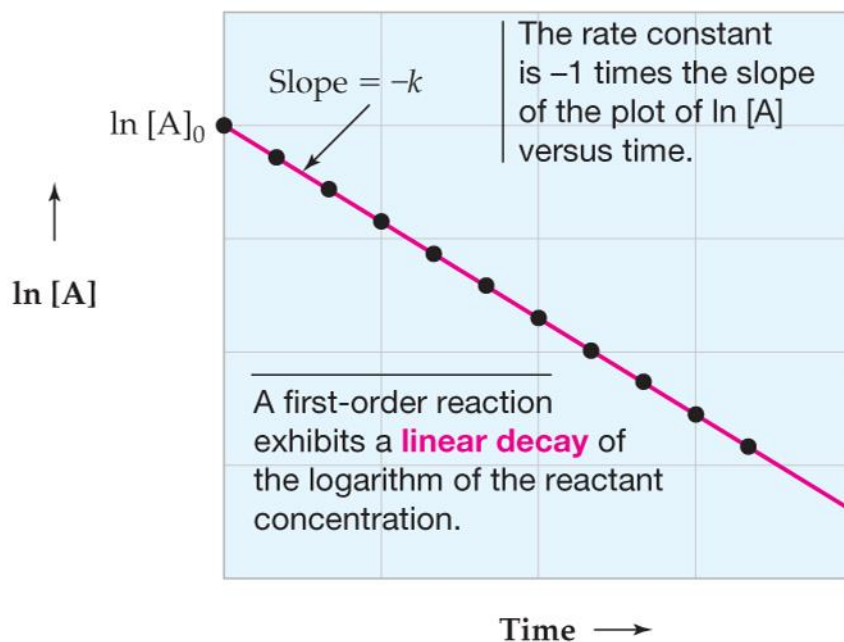
Integrated Rate Law for a First-Order Reaction

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

(a) Reactant concentration versus time



(b) Natural logarithm of reactant concentration versus time

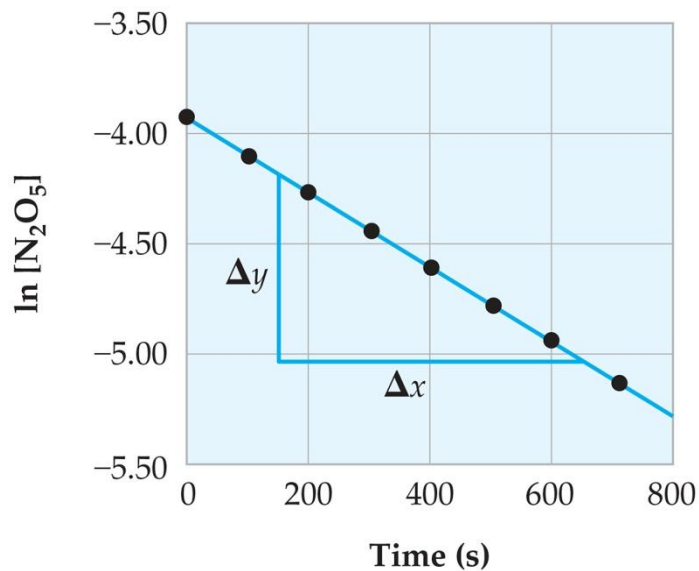


Integrated Rate Law for a First-Order Reaction



$$\text{rate} = k[\text{N}_2\text{O}_5]$$

Time (s)	$[\text{N}_2\text{O}_5]$	$\ln [\text{N}_2\text{O}_5]$
0	0.0200	-3.912
100	0.0169	-4.080
200	0.0142	-4.255
300	0.0120	-4.423
400	0.0101	-4.595
500	0.0086	-4.756
600	0.0072	-4.934
700	0.0061	-5.099



Slope = $-k$

Integrated Rate Law for a First-Order Reaction

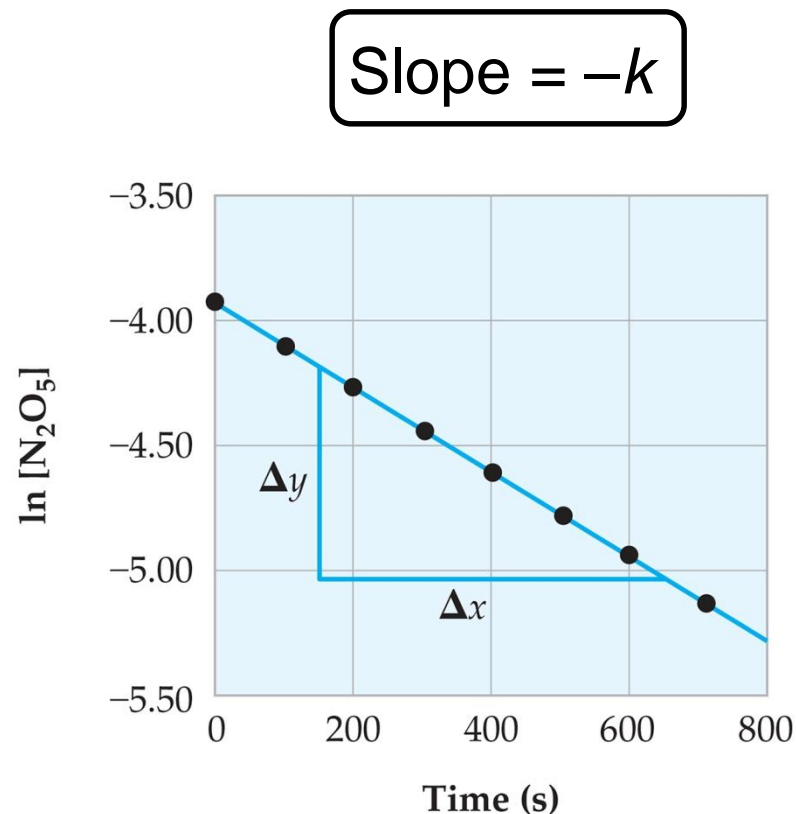


$$\text{rate} = k[\text{N}_2\text{O}_5]$$

Calculate the slope:

$$\frac{-5.02 - (-4.17)}{650 \text{ s} - 150 \text{ s}} = -0.0017 \frac{1}{\text{s}}$$

$$k = 0.0017 \frac{1}{\text{s}}$$



Half-Life of a First-Order Reaction

Half-Life: The time required for the reactant concentration to drop to one-half of its initial value



$$\text{rate} = k[A]$$

$$\ln \left(\frac{[A]_t}{[A]_0} \right) = -kt \left\{ \begin{array}{l} t = t_{1/2} \\ [A]_{t_{1/2}} = \frac{[A]_0}{2} \end{array} \right.$$

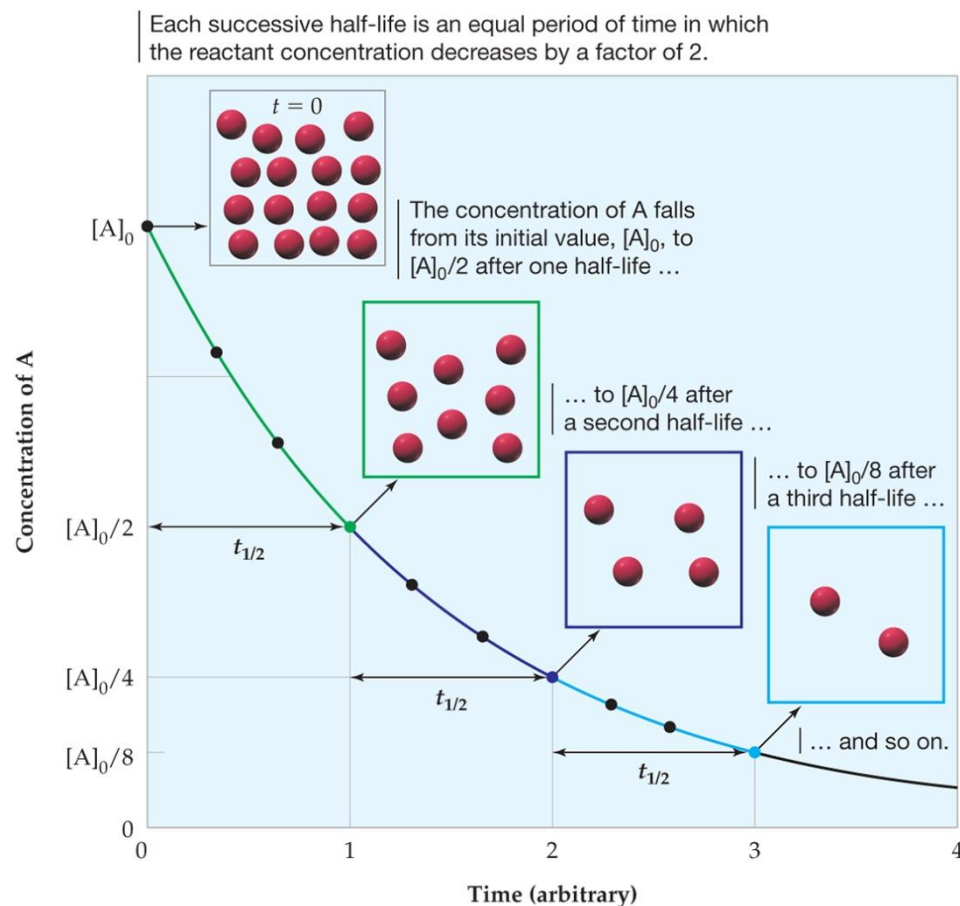
$$\ln \left(\frac{1}{2} \right) = -kt_{1/2} \quad \text{or} \quad t_{1/2} = \frac{0.693}{k}$$

Half-Life of a First-Order Reaction

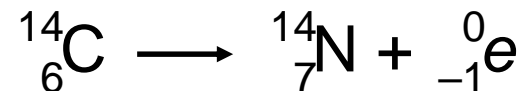
$$t_{1/2} = \frac{0.693}{k}$$

For a first-order reaction, the half-life is independent of the initial concentration.

Each successive half-life is an equal period of time.



Radioactive Decay Rates



$$t_{1/2} = 5715 \text{ y}$$

$$\text{Decay rate} = -\frac{\Delta N}{\Delta t} = kN$$

Second-Order Reactions



$$\text{rate} = k[A]^2 \qquad -\frac{\Delta[A]}{\Delta Dt} = k[A]^2$$

Calculus can be used to derive an **integrated** rate law.

$$\boxed{\begin{array}{l} \frac{1}{[A]_t} = kt + \frac{1}{[A]_0} \\ y = mx + b \end{array}}$$

$[A]_t$ concentration of A at time t

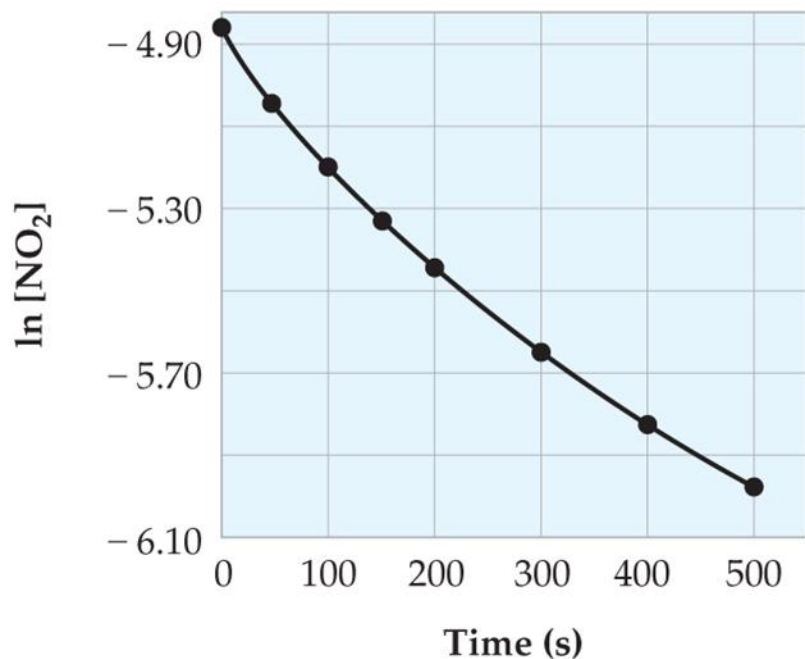
$[A]_0$ initial concentration of A

Second-Order Reactions



Time (s)	[NO ₂]	ln[NO ₂]	1/[NO ₂]
0	8.00×10^{-3}	-4.828	125
50	6.58×10^{-3}	-5.024	152
100	5.59×10^{-3}	-5.187	179
150	4.85×10^{-3}	-5.329	206
200	4.29×10^{-3}	-5.451	233
300	3.48×10^{-3}	-5.661	287
400	2.93×10^{-3}	-5.833	341
500	2.53×10^{-3}	-5.980	395

Second-Order Reactions



Plotting **ln[NO₂]** versus **time** gives a curve and not a straight-line fit.

Therefore, this is **not** a first-order reaction.

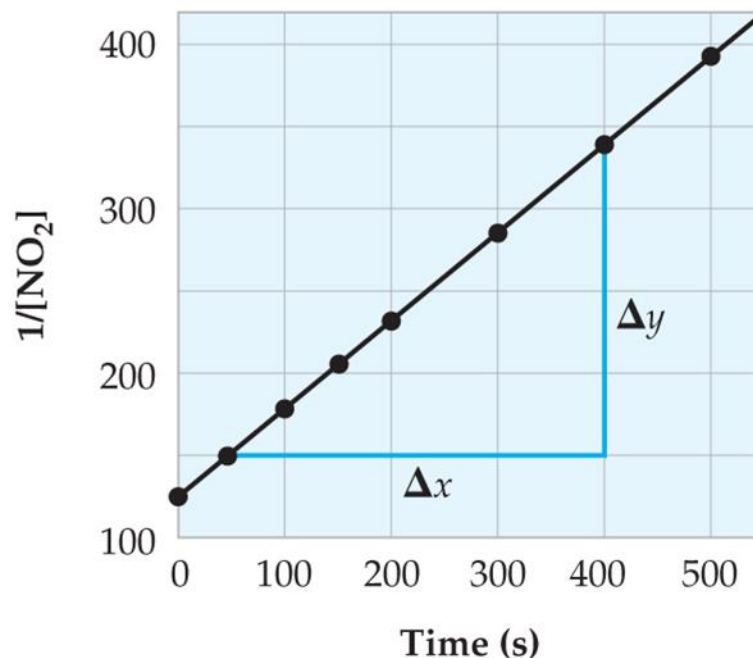
Second-Order Reactions



Plotting $\frac{1}{[\text{NO}_2]}$ versus
time gives a straight-line fit.

Therefore, this is a second-order reaction.

Slope = k



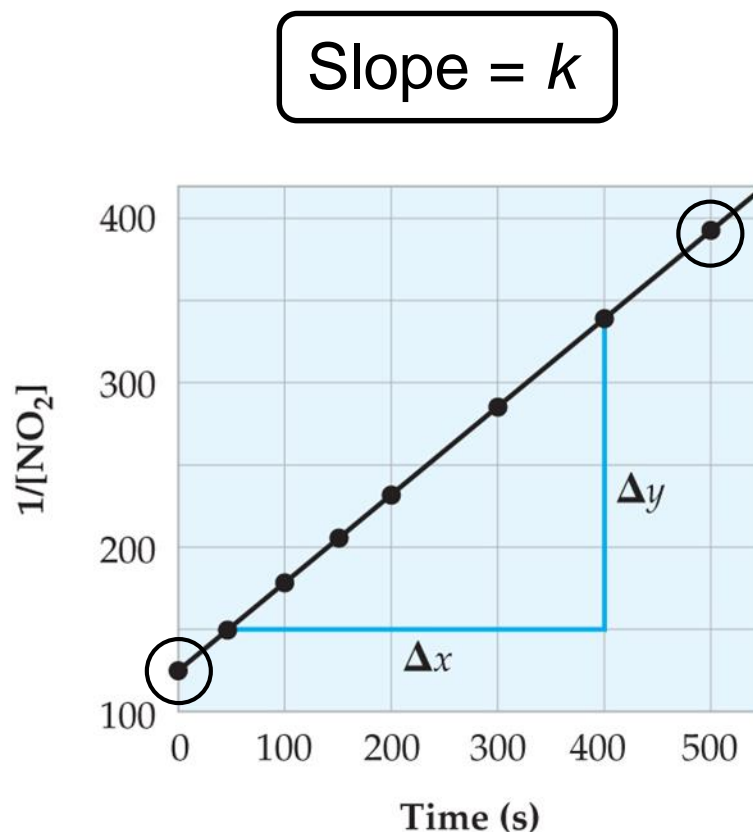
Second-Order Reactions



Calculate the slope:

$$\frac{(340 - 150) \frac{1}{\text{M}}}{400 \text{ s} - 50 \text{ s}} = 0.54 \frac{1}{\text{M s}}$$

$$k = 0.54 \frac{1}{\text{M s}}$$



Second-Order Reactions

Half-life for a second-order reaction



$$\text{rate} = k[A]^2$$

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0} \quad \left\{ \begin{array}{l} t = t_{1/2} \\ [A]_{t_{1/2}} = \frac{[A]_0}{2} \end{array} \right.$$

$$\frac{2}{[A]_0} = kt_{1/2} + \frac{1}{[A]_0}$$

$$t_{1/2} = \frac{1}{k[A]_0}$$

Second-Order Reactions

$$t_{1/2} = \frac{1}{k[A]_0}$$

For a second-order reaction, the half-life is dependent on the initial concentration.

Each successive half-life is twice as long as the preceding one.

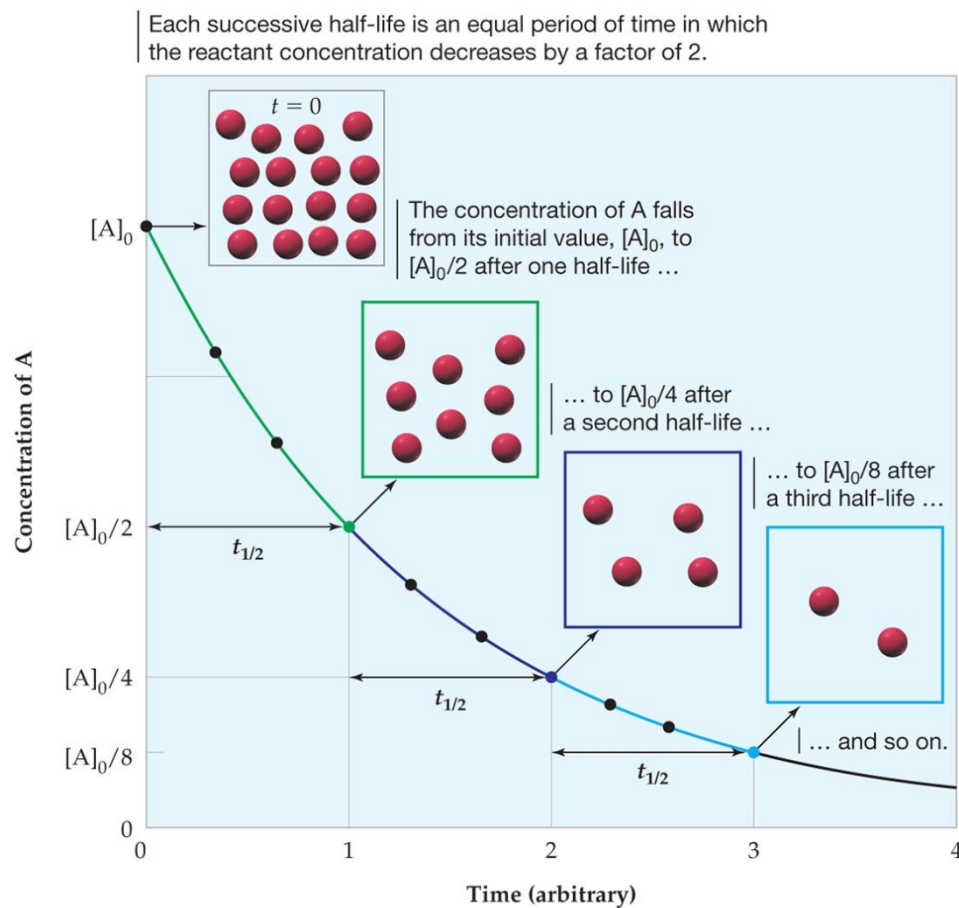
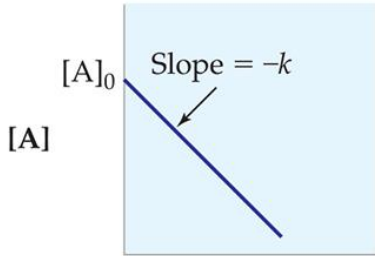
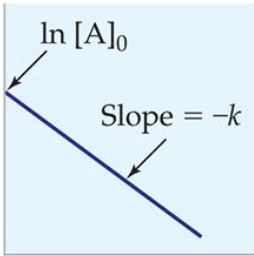
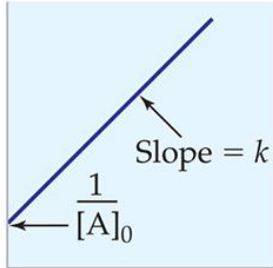


TABLE 13.4 Characteristics of Zeroth-, First-, and Second-Order Reactions of the Type $A \rightarrow \text{Products}$

	Zeroth-Order	First-Order	Second-Order
Rate law	$-\frac{\Delta[A]}{\Delta t} = k$	$-\frac{\Delta[A]}{\Delta t} = k[A]$	$-\frac{\Delta[A]}{\Delta t} = k[A]^2$
Integrated Rate Law	$[A]_t = -kt + [A]_0$	$\ln [A]_t = -kt + \ln [A]_0$	$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$
Linear graph	$[A]$ versus t 	$\ln [A]$ versus t 	$\frac{1}{[A]}$ versus t 
Graphical determination of k	$k = -(\text{Slope})$	$k = -(\text{Slope})$	$k = \text{Slope}$
Half-life	$t_{1/2} = \frac{[A]_0}{2k}$ (not constant)	$t_{1/2} = \frac{0.693}{k}$ (constant)	$t_{1/2} = \frac{1}{k[A]_0}$ (not constant)

Reaction Rates and Temperature: The Arrhenius Equation

Typically, as the temperature increases, the rate of reaction increases.



$$\text{rate} = k[\text{N}_2\text{O}_5]$$

The **rate constant** is dependent on temperature.

Reaction Rates and Temperature: The Arrhenius Equation

Magnesium is inert
in cold water.



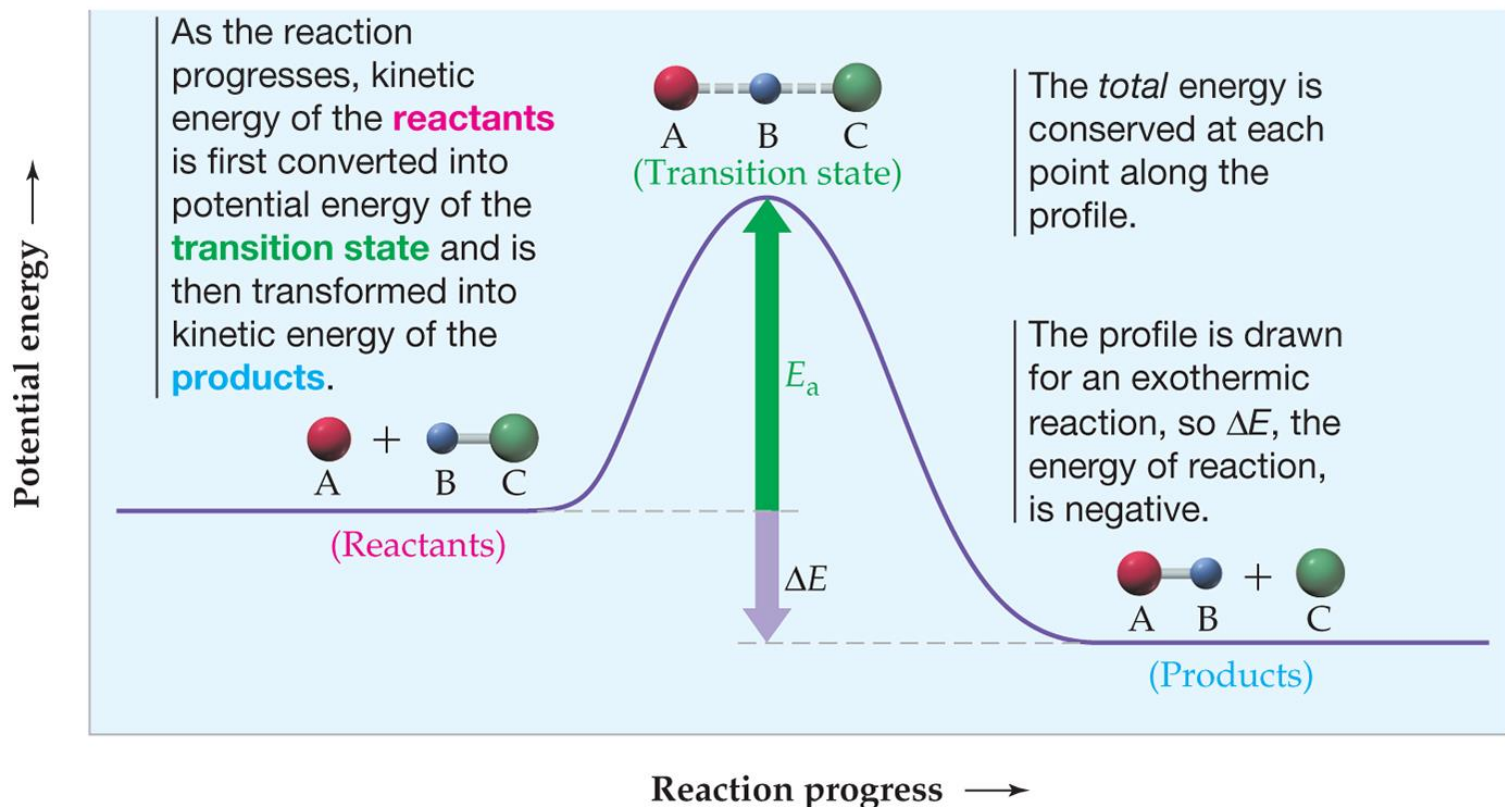
Magnesium reacts
in hot water.



The formation of bubbles of gas and the pink color of phenolphthalein, which indicates an alkaline solution, provide evidence of the reaction.

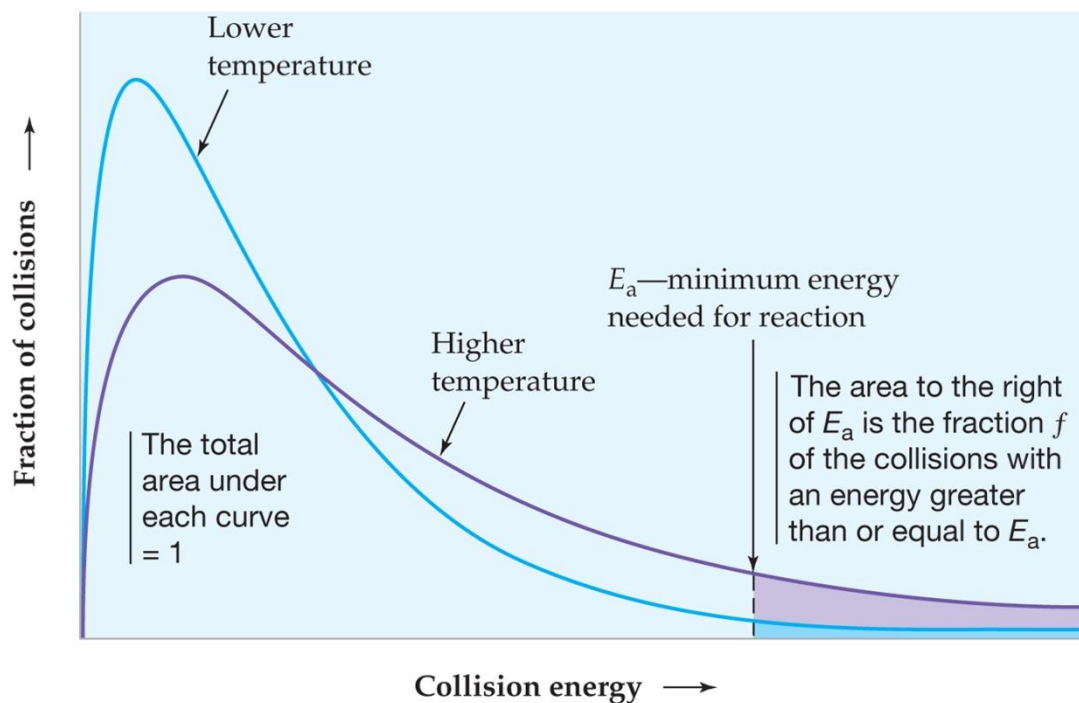
Reaction Rates and Temperature: The Arrhenius Equation

Transition State: The configuration of atoms at the maximum in the potential energy profile. This is also called the activated complex.



Reaction Rates and Temperature: The Arrhenius Equation

Collision Theory: As the average kinetic energy increases, the average molecular speed increases, and thus the collision rate increases.



The fraction of collisions that are sufficiently energetic to result in reaction increases exponentially with increasing temperature.

Reaction Rates and Temperature: The Arrhenius Equation

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

k Rate constant

A Collision frequency factor

E_a Activation energy

R Gas constant

T Temperature (K)

Using the Arrhenius Equation

$$\ln(k) = \ln(A) + \ln\left(e^{-E_a/RT}\right)$$

$$\ln(k) = \ln(A) - \frac{E_a}{RT} \xrightarrow{\text{rearrange the equation}} \ln(k) = \left(\frac{-E_a}{R}\right)\left(\frac{1}{T}\right) + \ln(A)$$

$y = mx + b$

Reaction Mechanisms

Reaction Mechanism: A sequence of reaction steps that describes the pathway from reactants to products

Elementary Reaction (step): A single step in a reaction mechanism

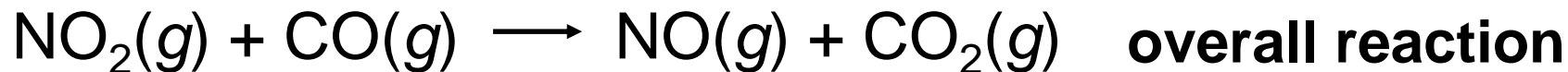
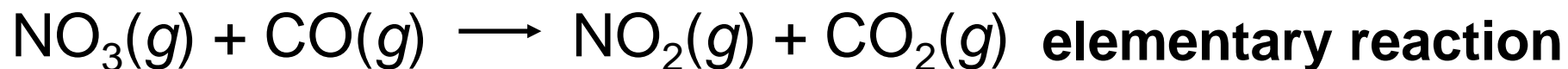
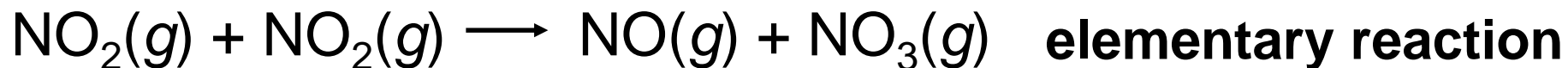
Reaction Mechanisms

An **elementary reaction** describes an individual molecular event.

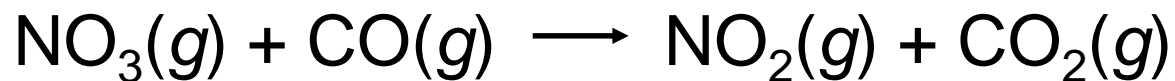
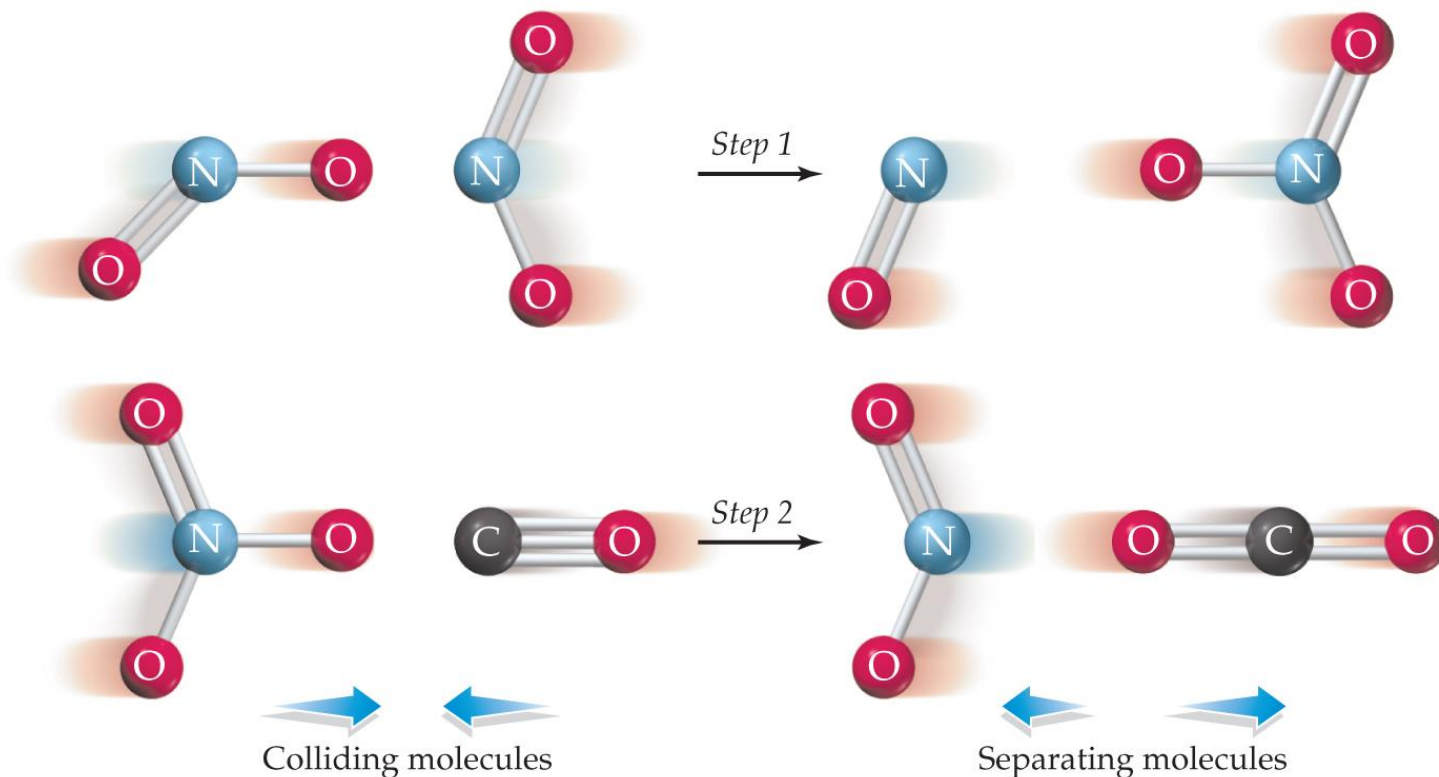
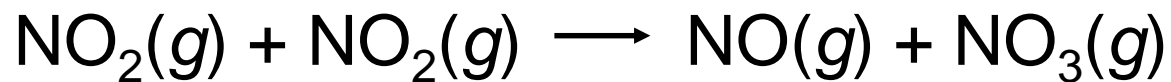
The **overall reaction** describes the reaction stoichiometry and is a summation of elementary reactions.

Reaction Mechanisms

Experimental evidence suggests that the reaction between NO_2 and CO takes place by a two-step mechanism:

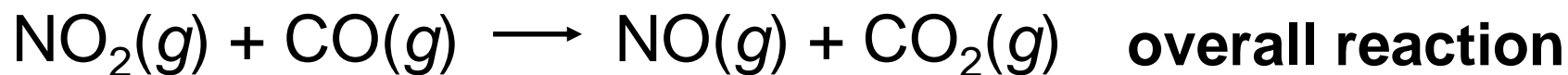
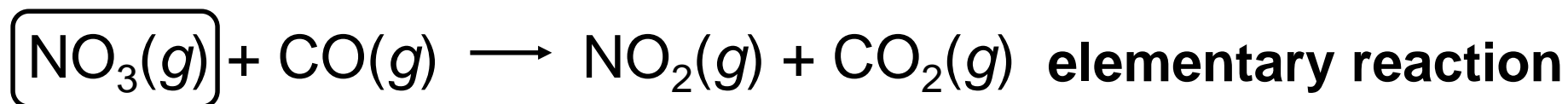
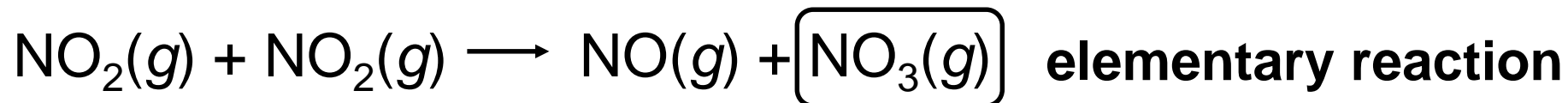


Reaction Mechanisms



Reaction Mechanisms

Experimental evidence suggests that the reaction between NO_2 and CO takes place by a two-step mechanism:



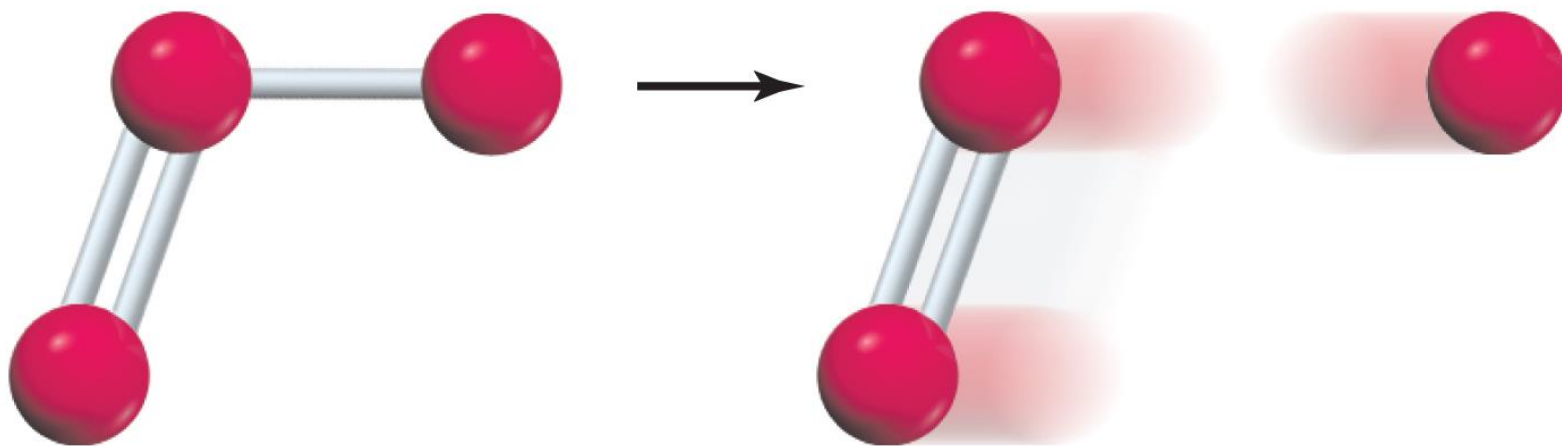
A **reactive intermediate** is formed in one step and consumed in a subsequent step.

Reaction Mechanisms-Molecularity

Molecularity: A classification of an elementary reaction based on the number of molecules (or atoms) on the reactant side of the chemical equation

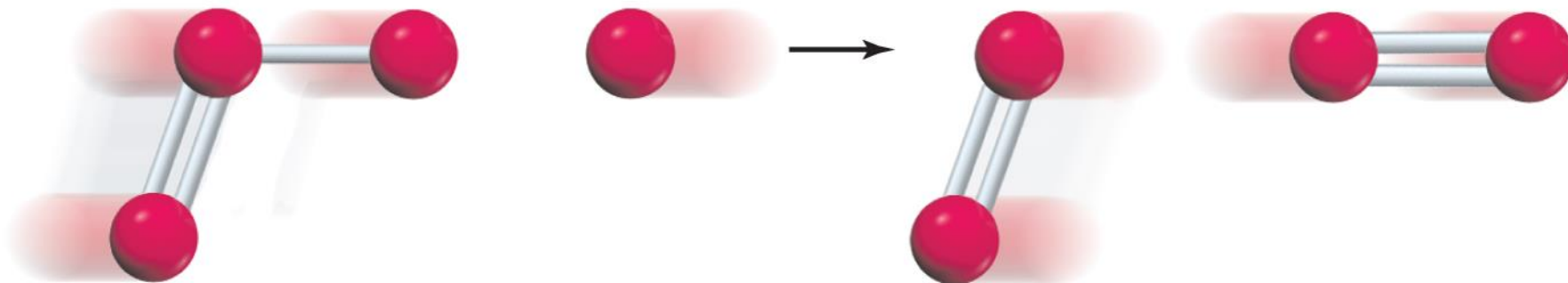
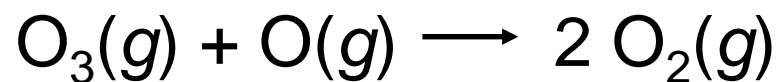
Reaction Mechanisms-Molecularity

Unimolecular Reaction



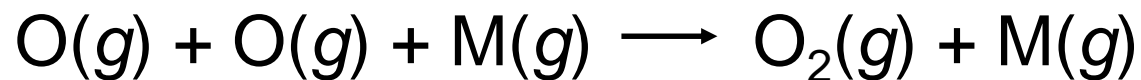
Reaction Mechanisms-Molecularity

Bimolecular Reaction



Reaction Mechanisms-Molecularity

Termolecular Reaction



Rate Laws for Elementary Reactions

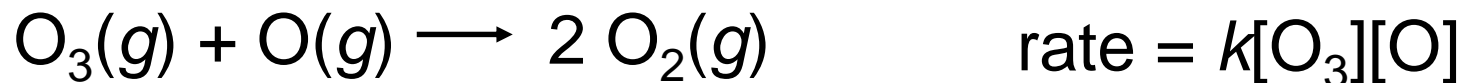
The rate law for an elementary reaction follows directly from its molecularity because an **elementary reaction is an individual molecular event.**

Rate Laws for Elementary Reactions

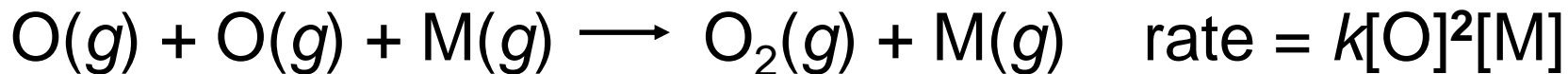
Unimolecular Reaction



Bimolecular Reaction



Termolecular Reaction



Rate Laws for Elementary Reactions

TABLE 13.5 Rate Laws for Elementary Reactions

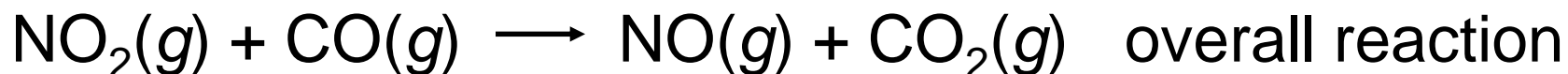
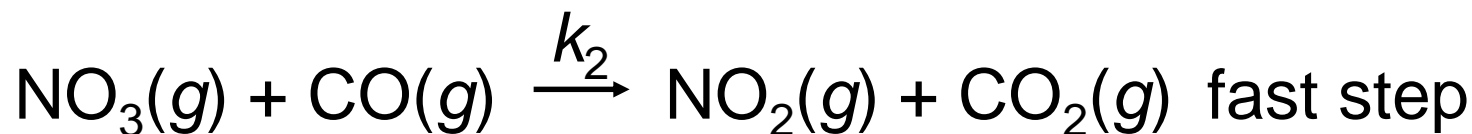
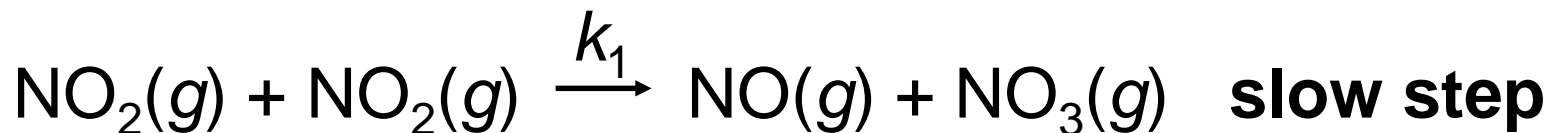
Elementary Reaction	Molecularity	Rate Law
$A \rightarrow \text{Products}$	Unimolecular	$\text{Rate} = k[A]$
$A + A \rightarrow \text{Products}$	Bimolecular	$\text{Rate} = k[A]^2$
$A + B \rightarrow \text{Products}$	Bimolecular	$\text{Rate} = k[A][B]$
$A + A + B \rightarrow \text{Products}$	Termolecular	$\text{Rate} = k[A]^2[B]$
$A + B + C \rightarrow \text{Products}$	Termolecular	$\text{Rate} = k[A][B][C]$

Rate Laws for Overall Reactions

Rate-Determining Step: The slow step in a reaction mechanism since it acts as a bottleneck and limits the rate at which reactants can be converted to products

Rate Laws for Overall Reactions

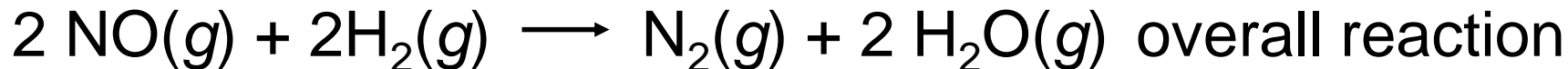
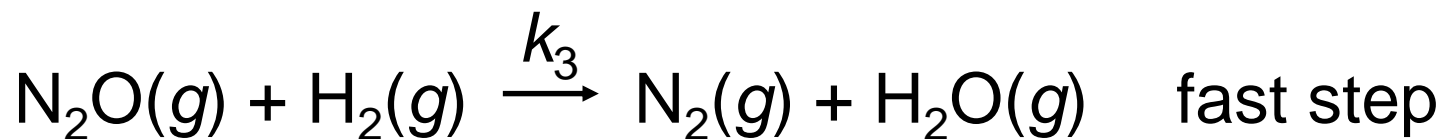
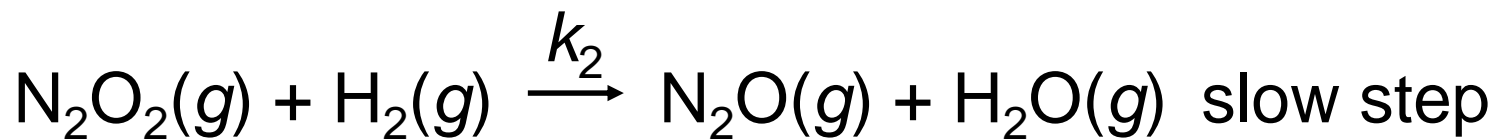
Multistep Reactions with an Initial Slow Step



Based on the slow step: **rate = $k_1[\text{NO}_2]^2$**

Rate Laws for Overall Reactions

Multistep Reactions with an Initial Fast Step



Based on the slow step: **rate = $k_2[\text{N}_2\text{O}_2][\text{H}_2]$**

Rate Laws for Overall Reactions

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

intermediate

First step: $\text{Rate}_{\text{forward}} = k_1[\text{NO}]^2$ $\text{Rate}_{\text{reverse}} = k_{-1}[\text{N}_2\text{O}_2]$

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

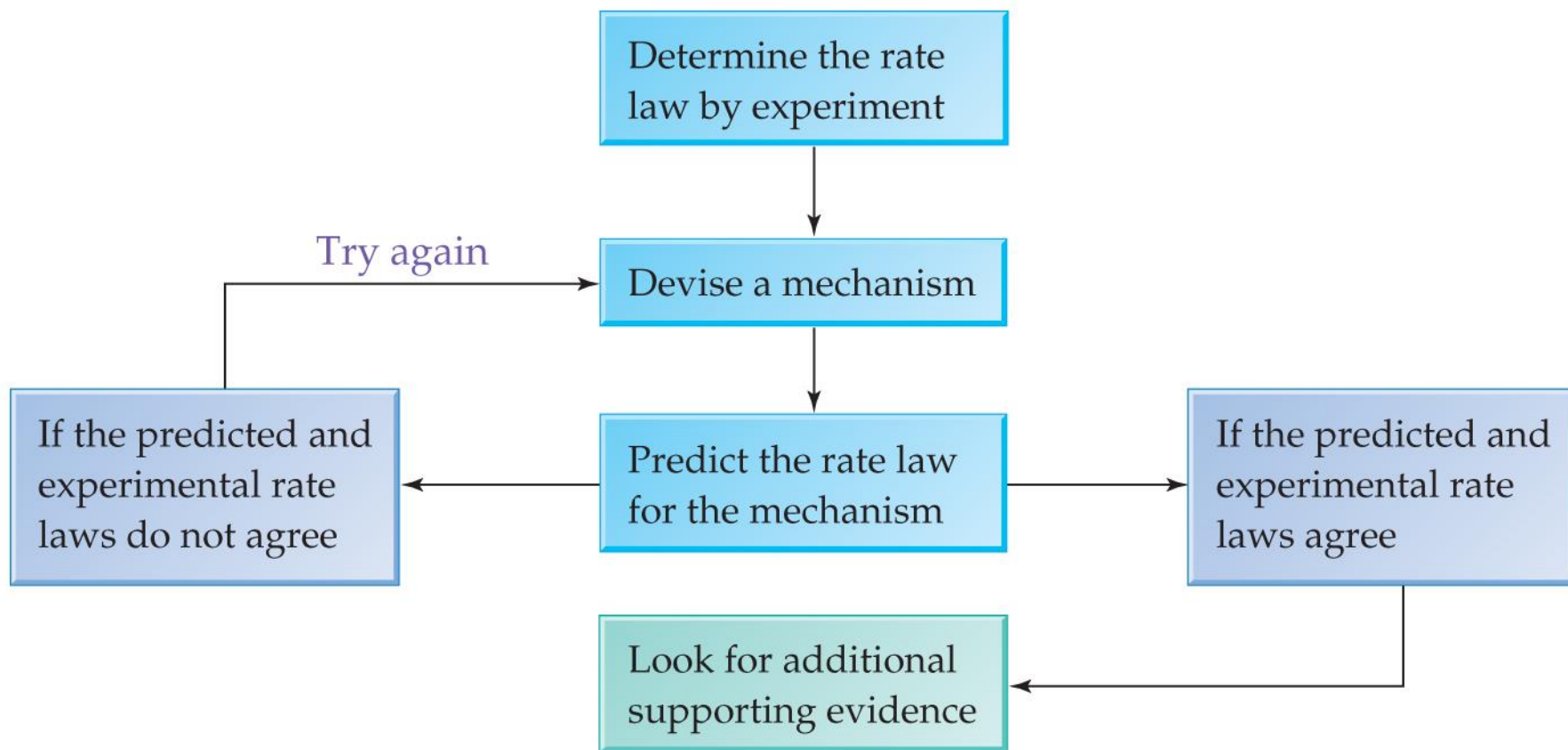
$$[\text{N}_2\text{O}_2] = \frac{k_1}{K_{-1}} [\text{NO}]^2$$

Slow step: $\text{rate} = k_2[\text{N}_2\text{O}_2][\text{H}_2]$

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

Rate Laws for Overall Reactions

Procedure for Studying Reaction Mechanisms

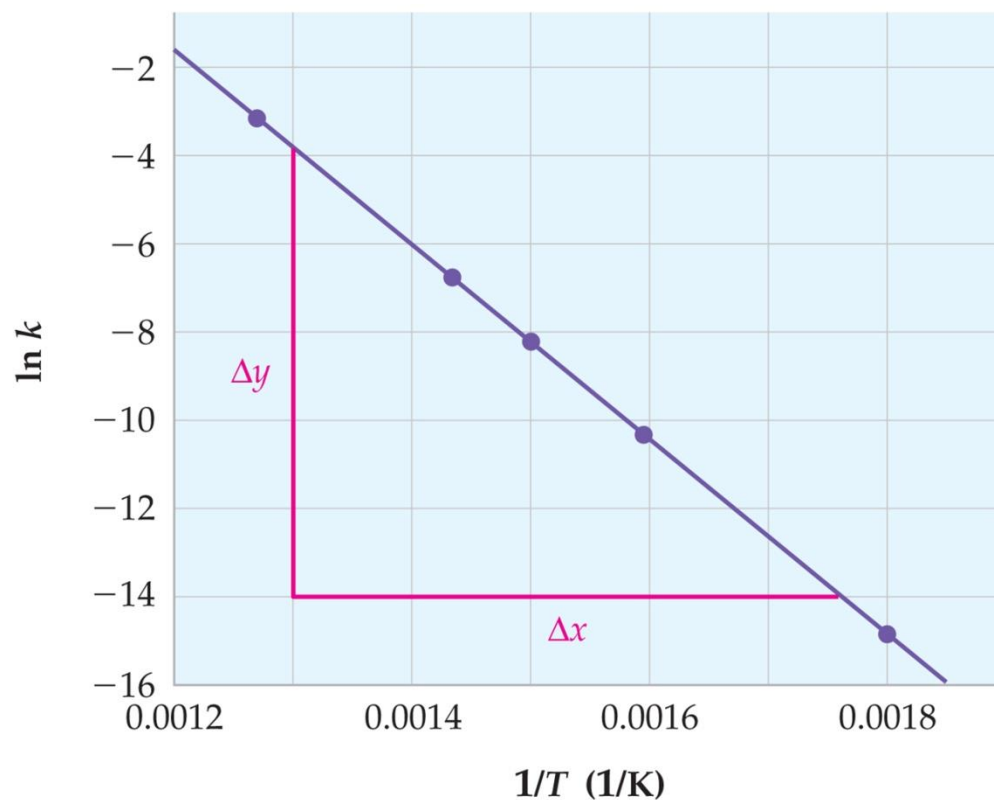


Using the Arrhenius Equation

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

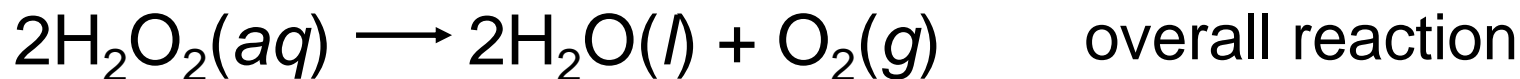
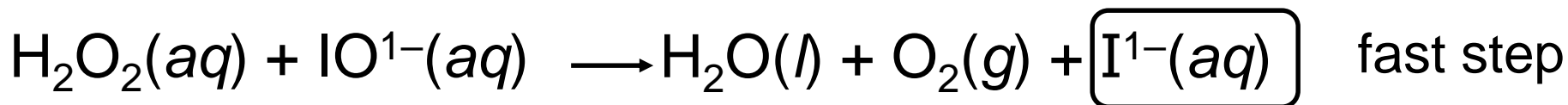
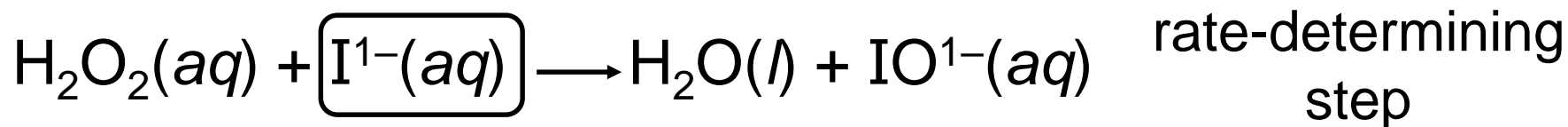
$$\text{Slope} = \frac{-E_a}{R}$$

Plot $\ln(k)$ versus $\frac{1}{T}$



Catalysis

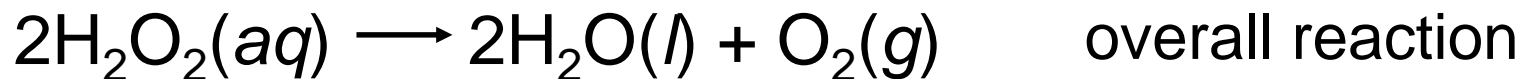
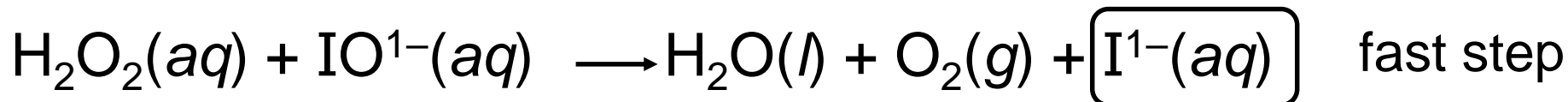
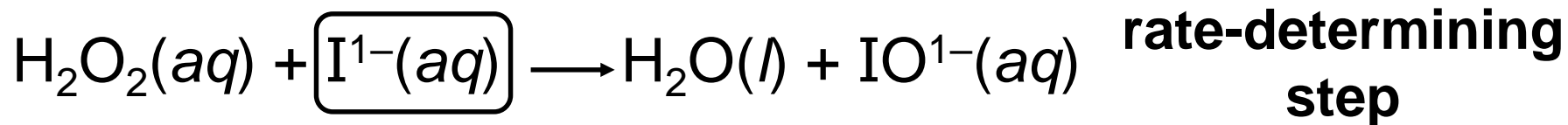
Catalyst: A substance that increases the rate of a reaction without itself being consumed in the reaction. A catalyst is used in one step and regenerated in a later step.



Catalysis

Since the catalyst is involved in the **rate determining step**, it often appears in the rate law.

$$\text{rate} = k[\text{H}_2\text{O}_2][\text{I}^{1-}]$$

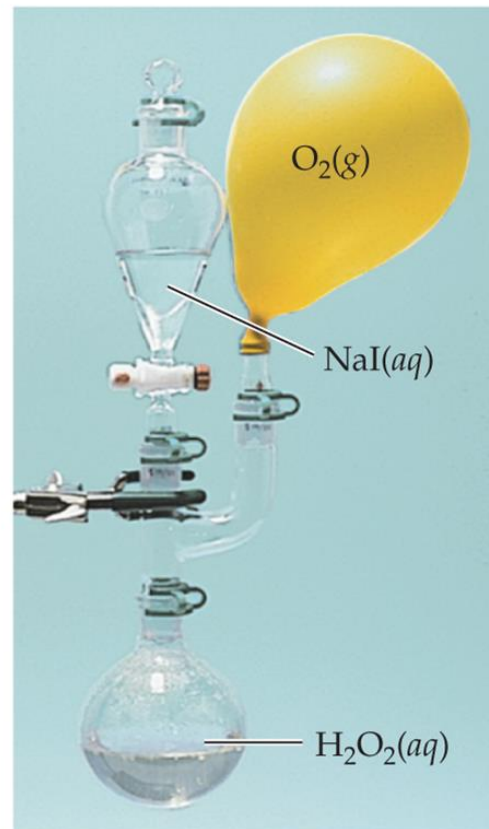


Catalysis

The rate of decomposition of aqueous hydrogen peroxide can be monitored qualitatively by collecting the evolved oxygen gas in a balloon.



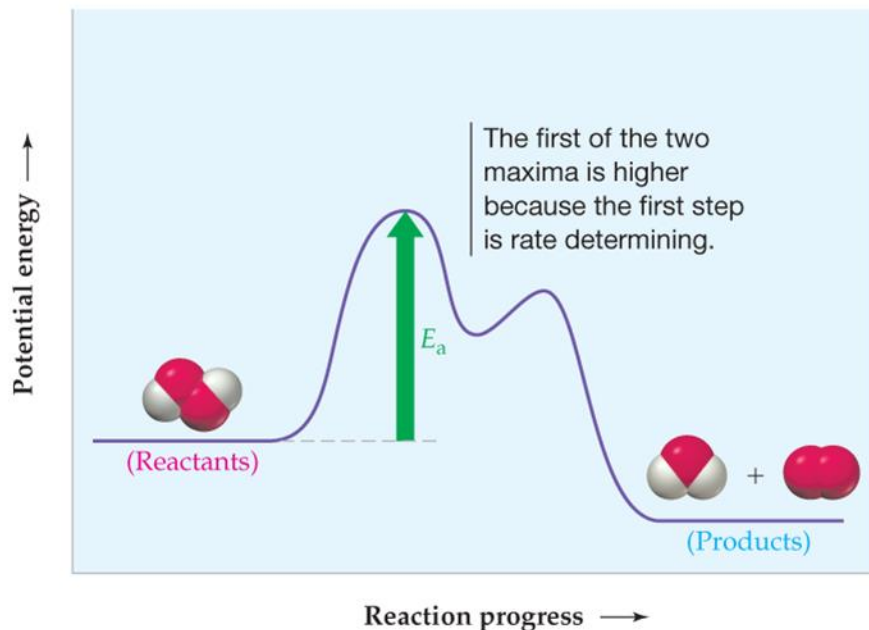
In the absence of a catalyst, little O₂ is produced.



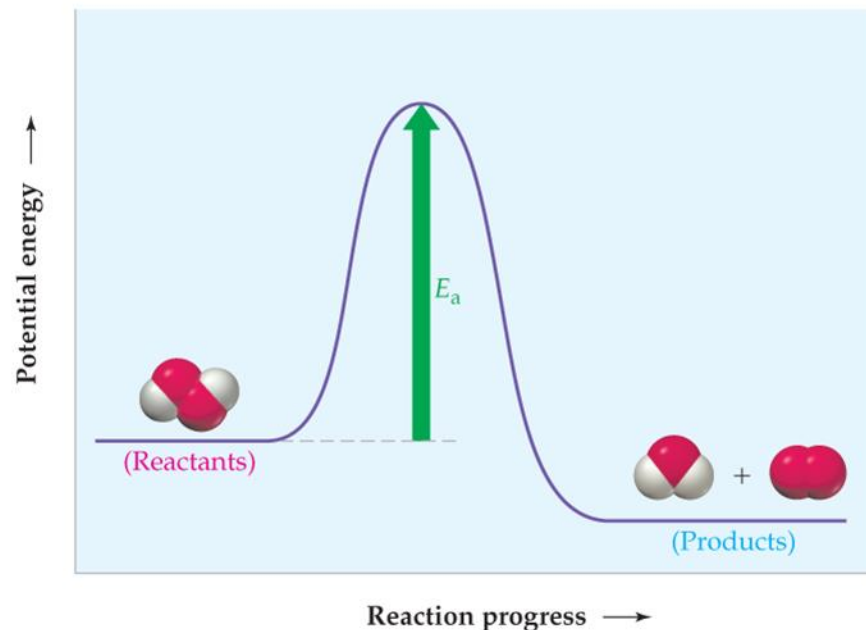
After addition of aqueous sodium iodide, by opening the stopcock to add NaI, the balloon rapidly inflates with O₂.

Catalysis

(a) Catalyzed pathway



(b) Uncatalyzed pathway



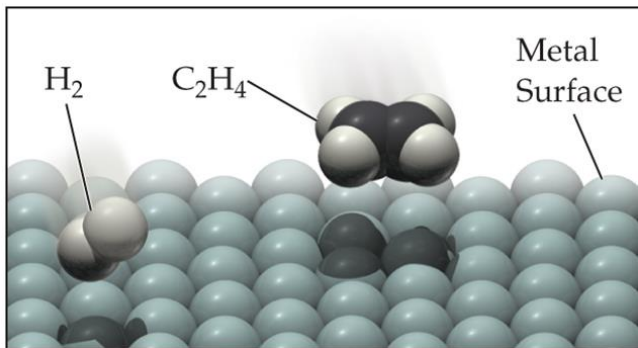
The **activation energy E_a** is lower for the catalyzed pathway. The shape of the barrier for the catalyzed pathway applies to the decomposition of H₂O₂.

Homogeneous and Heterogeneous Catalysts

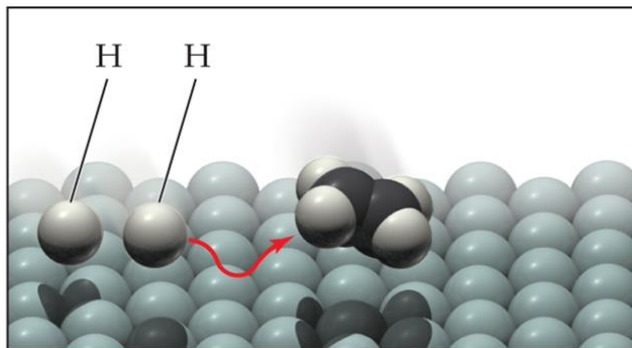
Homogeneous Catalyst: A catalyst that exists in the same phase as the reactants

Heterogeneous Catalyst: A catalyst that exists in a different phase from that of the reactants

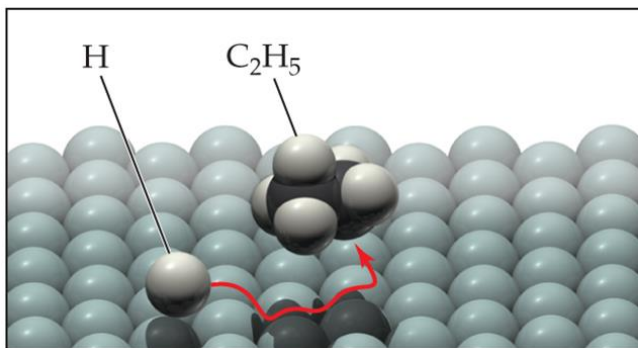
Homogeneous and Heterogeneous Catalysts



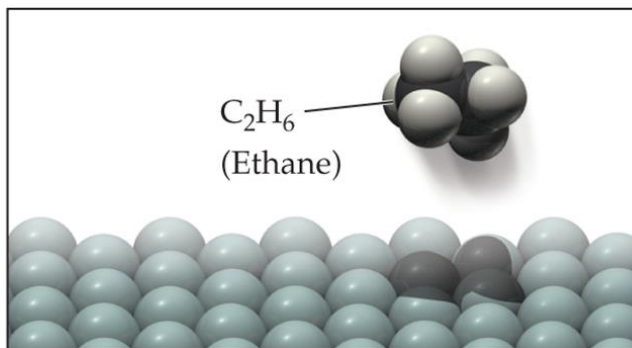
(a) H_2 and C_2H_4 are adsorbed on the metal surface.



(b) The H-H bond breaks as H-metal bonds form, and the H atoms move about on the surface.



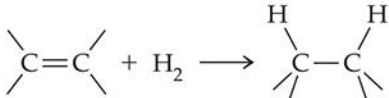
(c) One H atom forms a bond to a C atom of the adsorbed C_2H_4 to give a metal-bonded C_2H_5 group. A second H atom bonds to the C_2H_5 group.



(d) The resulting C_2H_6 molecule is desorbed from the surface.

Homogeneous and Heterogeneous Catalysts

TABLE 13.6 Some Heterogeneous Catalysts Used in Commercially Important Reactions

Reaction	Catalyst	Commercial Process	Product: Commercial Uses
$2 \text{SO}_2 + \text{O}_2 \longrightarrow 2 \text{SO}_3$	Pt or V_2O_5	Intermediate step in the contact process for synthesis of sulfuric acid	H_2SO_4 : Manufacture of fertilizers, chemicals; oil refining
$4 \text{NH}_3 + 5 \text{O}_2 \longrightarrow 4 \text{NO} + 6 \text{H}_2\text{O}$	Pt and Rh	First step in the Ostwald process for synthesis of nitric acid	HNO_3 : Manufacture of explosives, fertilizers, plastics, dyes, lacquers
$\text{N}_2 + 3 \text{H}_2 \longrightarrow 2 \text{NH}_3$	Fe, K_2O , and Al_2O_3	Haber process for synthesis of ammonia	NH_3 : Manufacture of fertilizers, nitric acid
$\text{H}_2\text{O} + \text{CH}_4 \longrightarrow \text{CO} + 3 \text{H}_2$	Ni	Steam-hydrocarbon re-forming process for synthesis of hydrogen	H_2 : Manufacture of ammonia, methanol
$\text{CO} + \text{H}_2\text{O} \longrightarrow \text{CO}_2 + \text{H}_2$	ZnO and CuO	Water-gas shift reaction to improve yield in the synthesis of H_2	H_2 : Manufacture of ammonia, methanol
$\text{CO} + 2 \text{H}_2 \longrightarrow \text{CH}_3\text{OH}$	Cu, ZnO, and Al_2O_3	Industrial synthesis of methanol	CH_3OH : Manufacture of plastics, adhesives, gasoline additives; industrial solvent
 $\text{C}=\text{C} + \text{H}_2 \longrightarrow \text{C}-\text{C}$	Ni, Pd, or Pt	Catalytic hydrogenation of compounds with $\text{C}=\text{C}$ bonds as in conversion of unsaturated vegetable oils to solid fats	Food products: margarine, shortening