

# CHEM 102

# General Chemistry II

# CHEM 101 Review



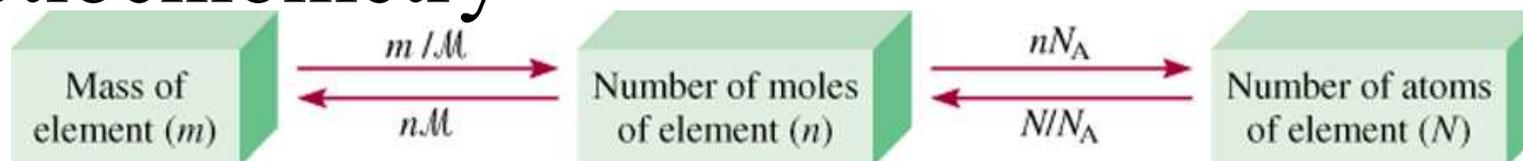
## ➤ Units

Mass, Temperature, Amount, Volume, Dimensional analysis, Density, Uncertainties, and Significant Figures.

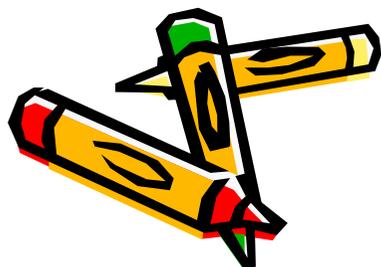
## ➤ Periodic Table, Atoms, and Ions



## ➤ Stoichiometry



## ➤ Concentration units



Mass Percent (%), Mole Fraction ( $X$ ),  
Molarity ( $M$ ), Molality ( $m$ ), Normality ( $N$ )

# CHEM 101 Review

## ➤ Reaction: Balance and Stoichiometry

Types: Precipitation, Acid-base, Redox..



## ➤ Thermochemistry

First law of thermodynamics:  $\Delta E = q + w$

$$q = C\Delta t = ms\Delta t \quad \text{and} \quad w_{exp} = -P\Delta V$$

## ➤ Gas Laws

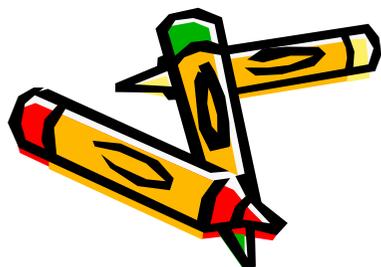
$$P \times V = nRT$$

← Ideal Gas Law

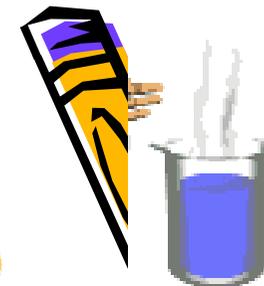
$$[P_{obs} + a(n/V)^2] \times (V - nb) = nRT$$

corrected pressure

corrected volume

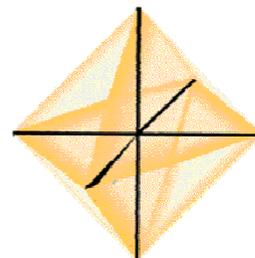


# CHEM 101 Review

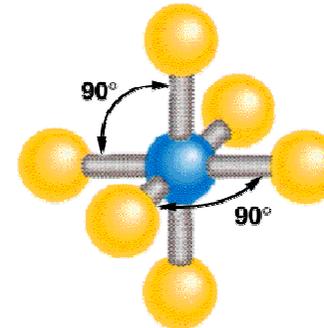


## ➤ Molecules and Geometry

Ionic, Covalent, Lewis Structure,  
VSEPR, Hybridization,  
Molecular Orbital Model

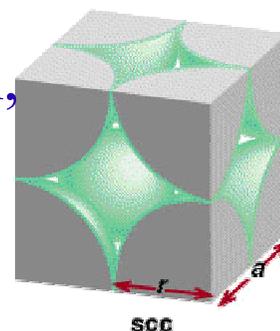


Octahedral

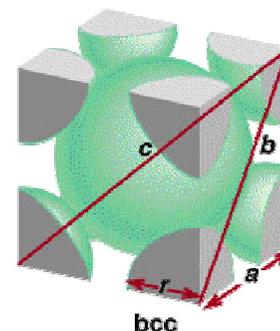


## ➤ Liquid and Solids

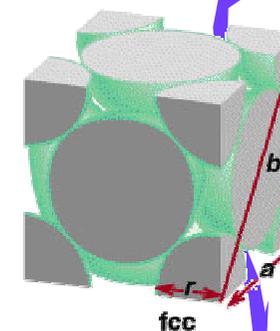
IMF (Dipole-dipole, Dispersion,  
Hydrogen Bonding), Solids  
(Unit cell-SC, BCC, FCC),...



sc



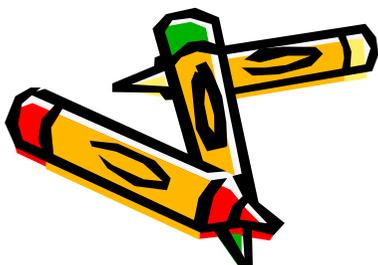
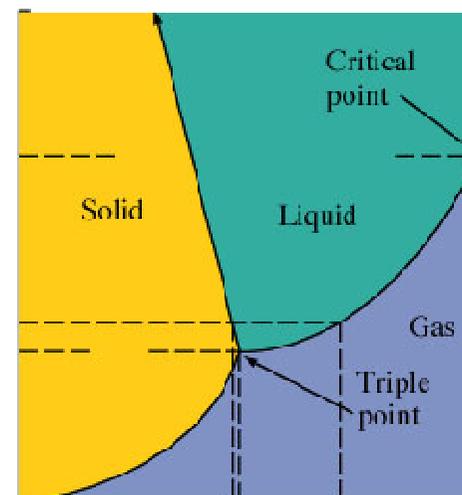
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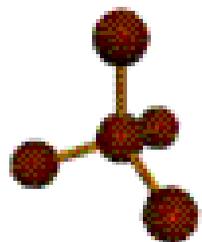


fcc

## ➤ Properties of Solutions

Vapor Pressure Lowering,  
Boiling Point, Elevation  
Freezing Point Depression,  
Osmotic Pressure..





# Chapter 12

# Chemical Kinetics





# Course Outline



- **Chemical Kinetics** (Chap.12)
- **Chemical Equilibrium** (Chap. 13)
- **Acid-Base and Aqueous Equilibria** (Chap. 14 & 15)
- **Spontaneity, Entropy, and Free Energy** (Chap. 16)
- **Electrochemistry** (Chap. 17)
- **The nucleus: A chemist View** (Chap. 18)
- **The Representative Elements: 1A to 8A** (Chap. 19 & 20)
- **Transition Chemistry: Coordination Chemistry** (Chap. 21)
- **Organic Chemistry: Biological Molecules** (Chap. 22)



# Chapter 12 Preview

## Chemical Kinetics



- **Reaction Rates**

Stoichiometry, Conditions, Concentration

- **Rate Laws**

Types (Order), Initial Rate, First Order, Second Order and Zero Order Rxns

- **Reaction Mechanisms and Model of Chemical Kinetics**

- **Catalysis**

Heterogeneous and Homogeneous



# Introduction



- **The main goal of Chemical Kinetics is to understand** the steps by which a reaction take place.
- **The understanding will allow us to find ways to facilitate or inhibit the reactions,**  
e.g., Haber process for Ammonia production requires high temperature (more cost for commercially feasibility), while presence of iron oxide speeds up the reaction.

**Thermodynamics** – does a reaction take place?

**Kinetics** – how fast does a reaction proceed?



# 12.1 Reaction Rates



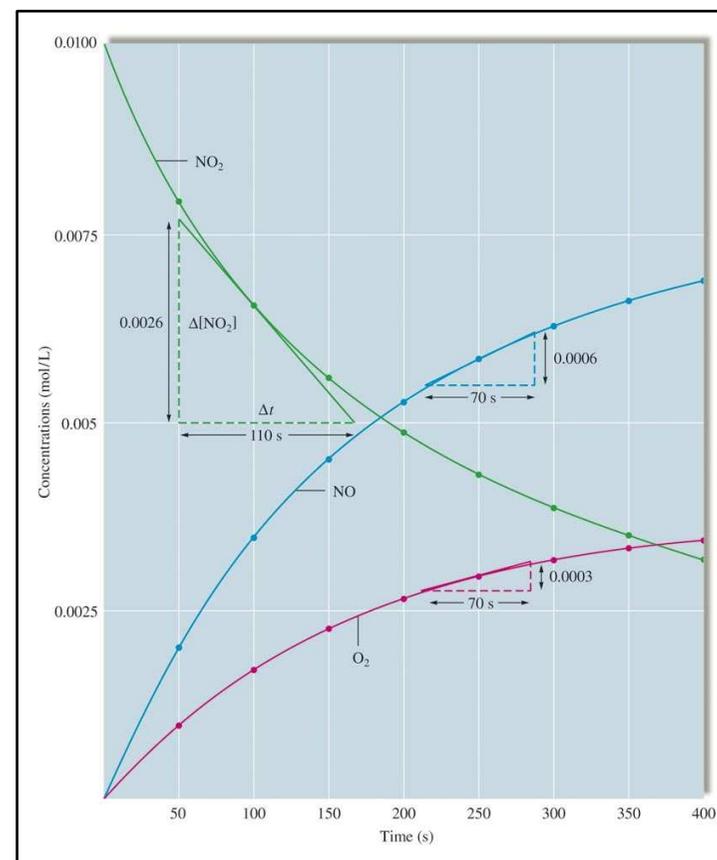
**Reaction rate** is the change in the concentration of a reactant or a product with time ( $M/s$ ).

**Figure 12.1** Following reaction at  $300\text{ }^\circ\text{C}$  and Oxygen are Plotted versus Time



$$\text{Average Rate} = \frac{\Delta M}{\Delta t}$$

Rate of the *disappearance* of  $\text{NO}_2$  is equal in magnitude but opposite in sign to the rate of the *appearance* of  $\text{NO}$  and *half* of the *appearance* of  $\text{O}_2$

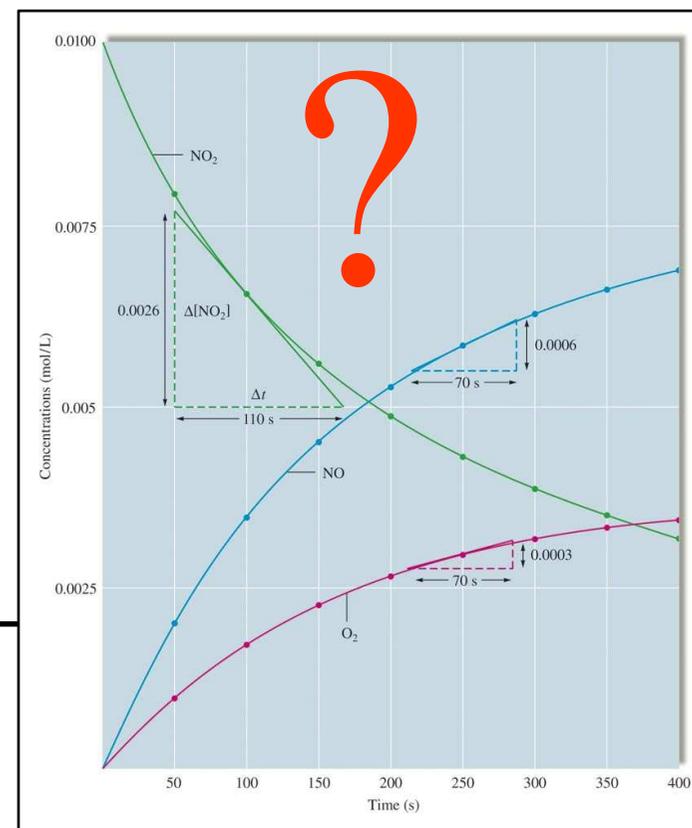
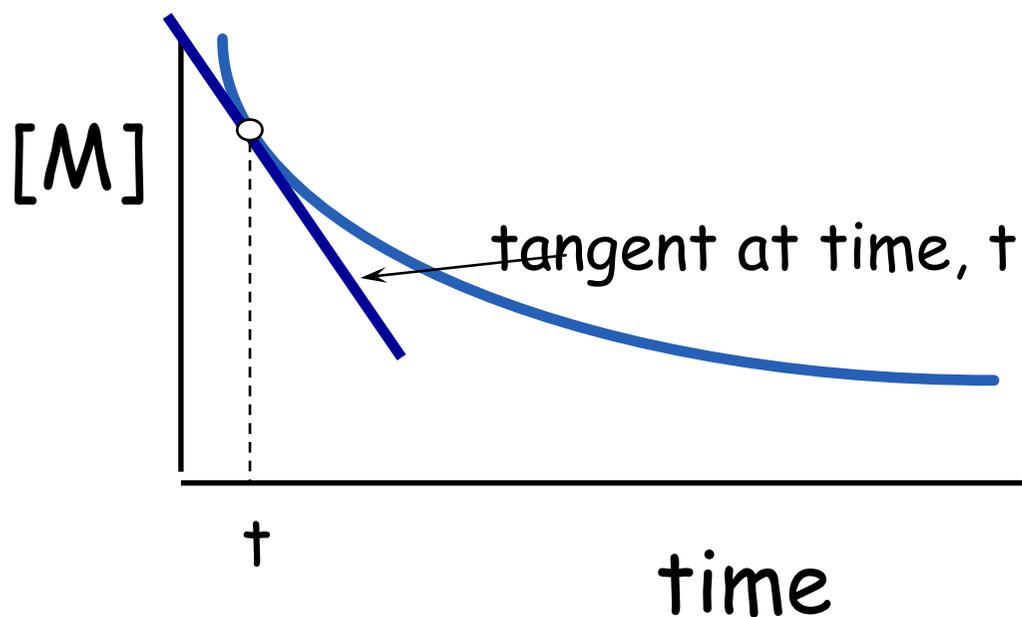




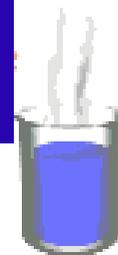
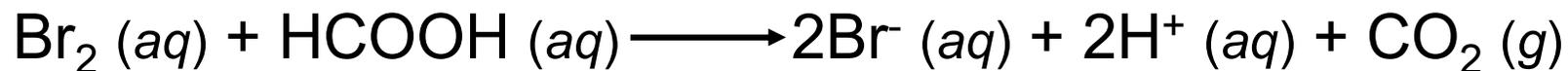
# 12.1 Reaction Rate



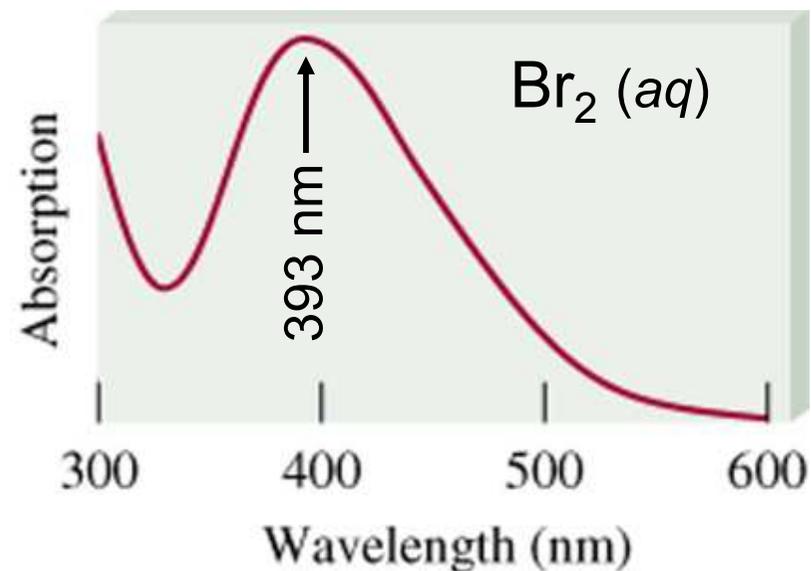
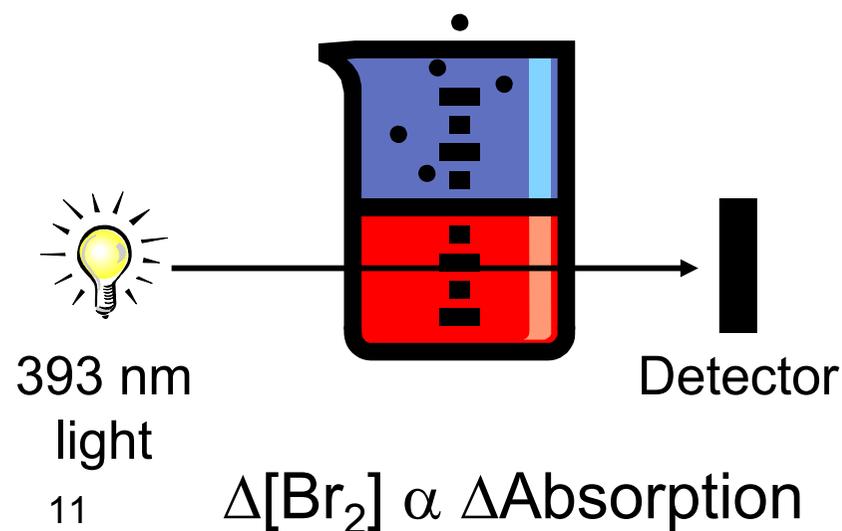
- Average Rate-- $\Delta$  mol (or concentration) over a period of time,  $\Delta t$
- Instantaneous Rate-- slope of the tangent at a specific time,  $t$
- Initial Rate-- instantaneous rate at  $t = 0$



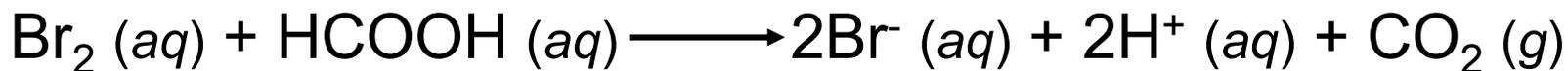
# 12.1 Reaction Rates Examples



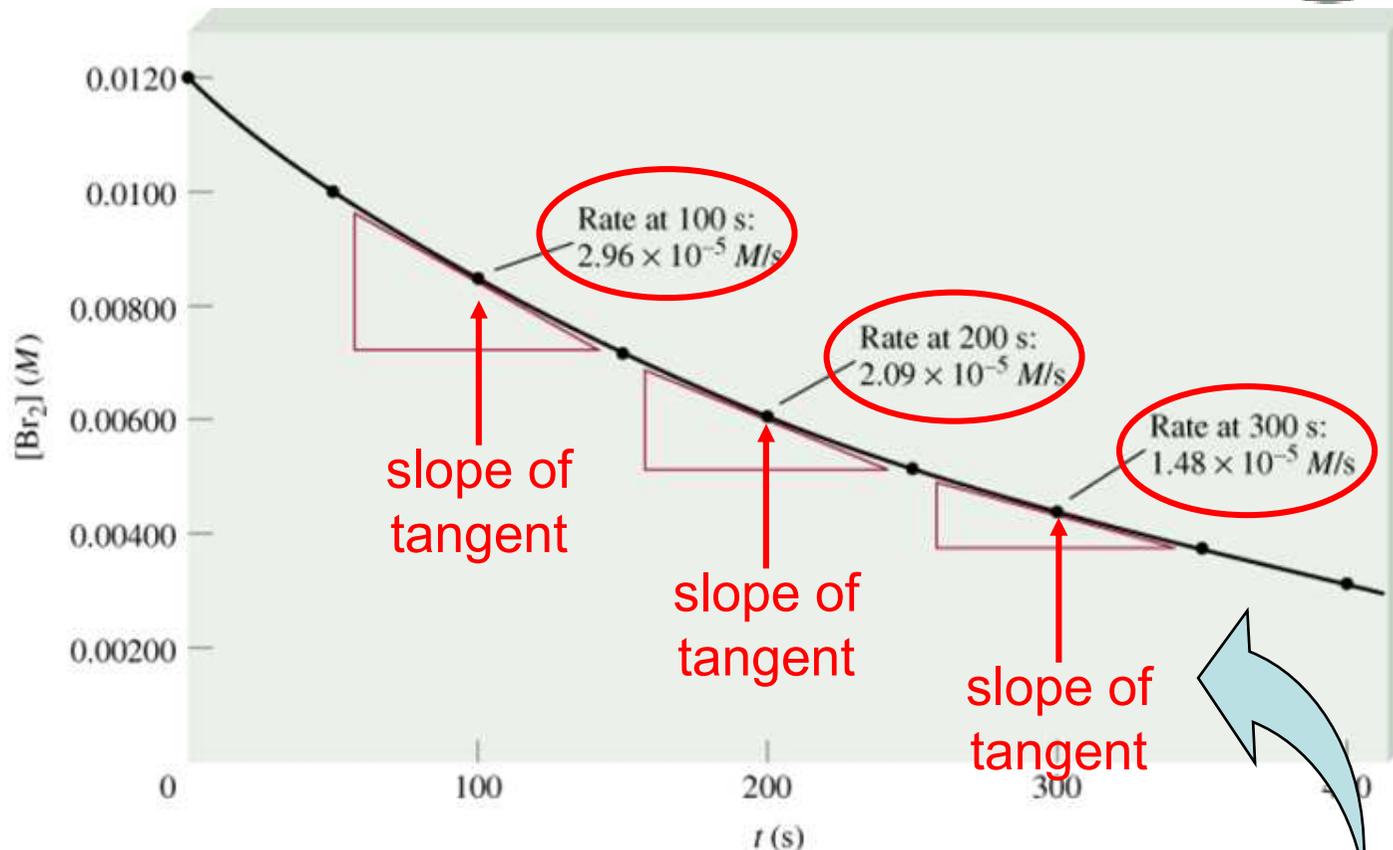
time



# 12.1 Reaction Rates Examples



Time (s)	[Br <sub>2</sub> ] (M)
0.0	0.0120
50.0	0.0101
100.0	0.00846
150.0	0.00710
200.0	0.00596
250.0	0.00500
300.0	0.00420
350.0	0.00353
400.0	0.00296

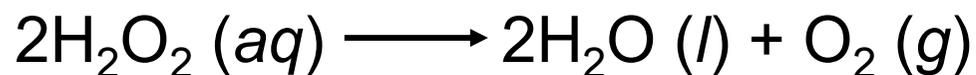
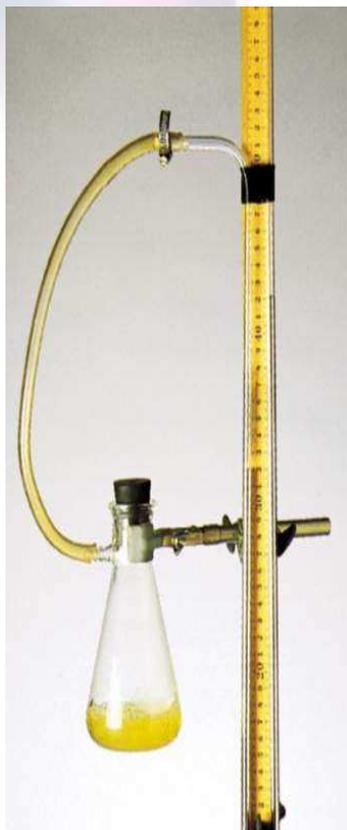


$$\text{Average rate} = -\frac{\Delta[\text{Br}_2]}{\Delta t} = -\frac{[\text{Br}_2]_{\text{final}} - [\text{Br}_2]_{\text{initial}}}{t_{\text{final}} - t_{\text{initial}}}$$

**instantaneous rate** = rate for specific instance in time



# 12.1 Reaction Rates Examples

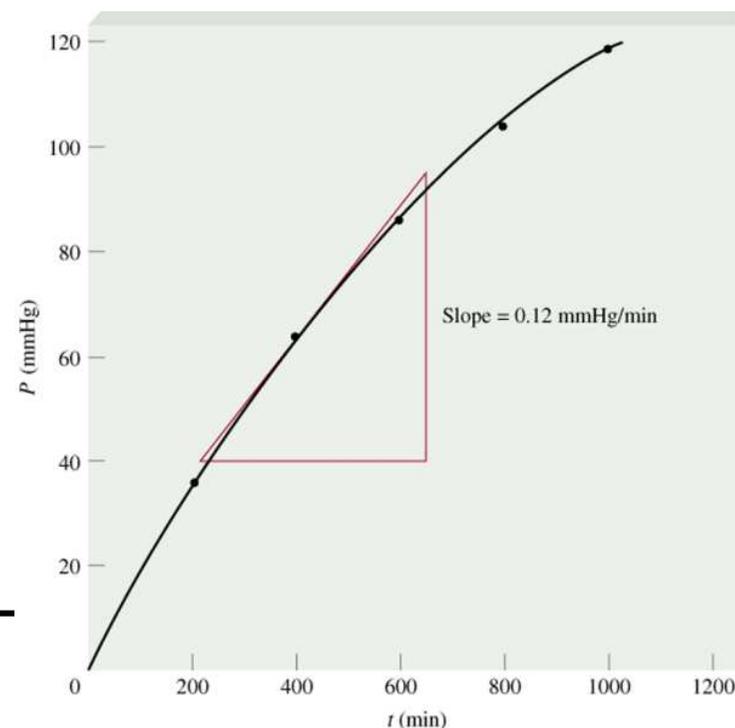


$$PV = nRT$$

$$P = \frac{n}{V} RT = [\text{O}_2]RT$$

$$[\text{O}_2] = \frac{1}{RT} P$$

$$\text{rate} = \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{1}{RT} \frac{\Delta P}{\Delta t}$$



measure  $\Delta P$  over time



# 12.1 Reaction Rates



## Reaction Rates and Stoichiometry

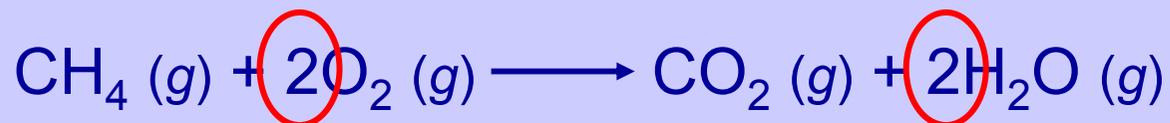


Two moles of A disappear for each mole of B that is formed.

$$\text{rate} = - \frac{1}{2} \frac{\Delta[A]}{\Delta t} \qquad \text{rate} = \frac{\Delta[B]}{\Delta t}$$



Write the rate expression for the following reaction:



$$\text{rate} = - \frac{\Delta[\text{CH}_4]}{\Delta t} = - \frac{1}{2} \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{\Delta[\text{CO}_2]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t}$$



## 12.2 Rate Laws:



The **rate law** expresses the relationship of the rate of a reaction to the rate constant and the concentrations of the reactants raised to some powers.



$$\text{Rate} = k [A]^x [B]^y$$

reaction is **xth order** in A

reaction is **yth order** in B

reaction is **(x + y)th order overall**

**k** is the rate constant for the reaction



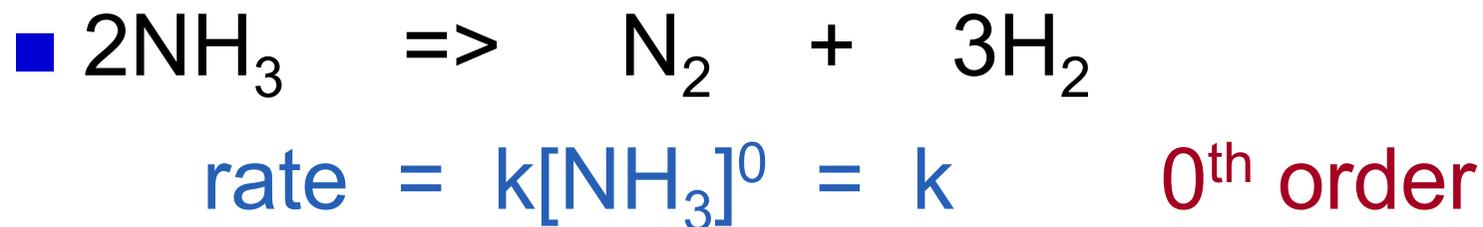
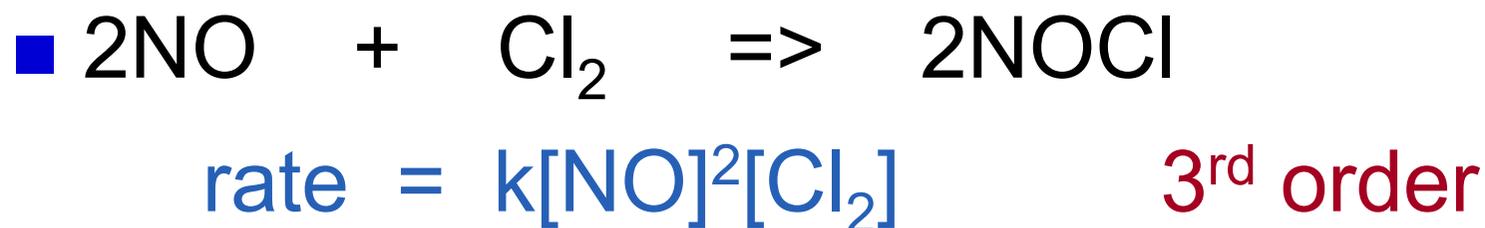
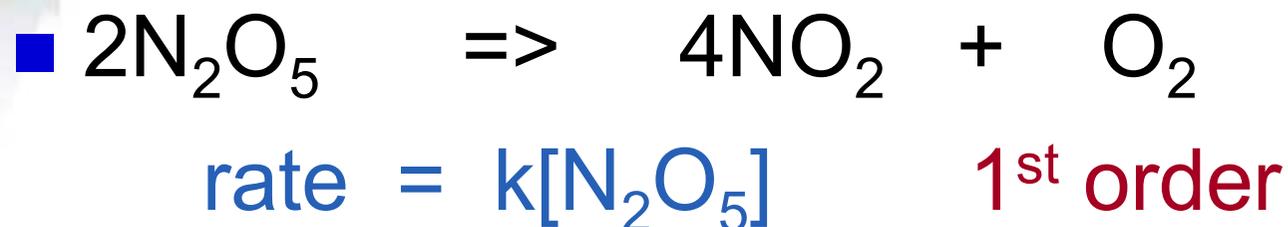
### Types of Rate Laws

A. Differential rate law: how rate depend on concentration.

B. Integrated rate law: how the concentration depend on time.



# Examples:

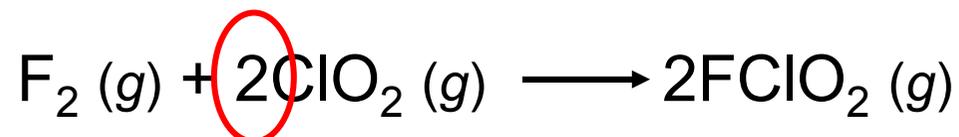
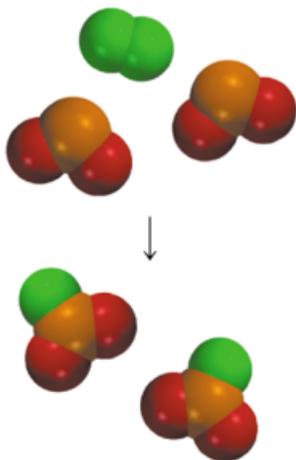




## 12.2 Rate Law Summary



- Rate laws are **always** determined experimentally.
- Reaction order is **always** defined in terms of reactant (not product) concentrations.
- The order of a reactant **is not** related to the stoichiometric coefficient of the reactant in the balanced chemical equation.



$$\text{rate} = k [\text{F}_2][\text{ClO}_2]^1$$



# 12.3 Determination of the Form of Rate Law



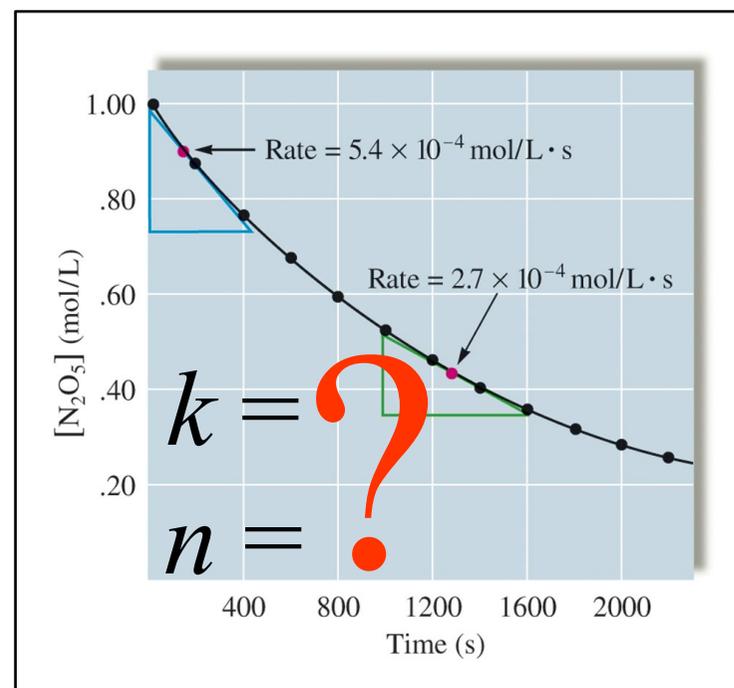
**Figure 12.3 A Plot of the Concentration of  $N_2O_5$  as a Function of Time for the Reaction**



$$rate = - \frac{\Delta[N_2O_5]}{\Delta t} = k[N_2O_5]^n$$

at  $M_1=0.90$  Mol/L ...  $r_1 = 5.4 \times 10^{-4}$  Mol/L.s

at  $M_2=0.45$  Mol/L ...  $r_2 = 2.7 \times 10^{-4}$  Mol/L.s

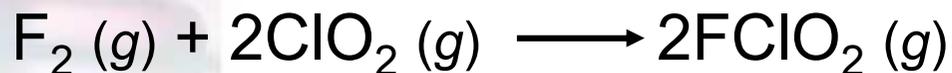


Exponent Value	[conc]	rate
0	double	same
1	double	double
2	double	x 4
3	double	x 8
4	double	x 16

$$n = 1 \quad k = ?$$



# 12.3 Determination of the Form of Rate Law



$$\text{rate} = k [\text{F}_2]^x [\text{ClO}_2]^y$$

$$k =$$

$$x =$$

$$y =$$



## Initial Rates Method

**Table 13.2** Rate Data for the Reaction between  $\text{F}_2$  and  $\text{ClO}_2$

	$[\text{F}_2](M)$	$[\text{ClO}_2](M)$	Initial Rate (M/s)
1.	0.10	0.010	$1.2 \times 10^{-3}$
2.	0.10	0.040	$4.8 \times 10^{-3}$
3.	0.20	0.010	$2.4 \times 10^{-3}$

Double  $[\text{F}_2]$  with  $[\text{ClO}_2]$  constant

Rate doubles

$$x = 1$$

Quadruple  $[\text{ClO}_2]$  with  $[\text{F}_2]$  constant

Rate quadruples

$$y = 1$$

$$\text{rate} = k [\text{F}_2][\text{ClO}_2]$$



Determine the rate law and calculate the rate constant for the following reaction from the following data:



Experiment	$[\text{S}_2\text{O}_8^{2-}]$	$[\text{I}^-]$	Initial Rate (M/s)
1	0.08	0.034	$2.2 \times 10^{-4}$
2	0.08	0.017	$1.1 \times 10^{-4}$
3	0.16	0.017	$2.2 \times 10^{-4}$

$$\text{rate} = k [\text{S}_2\text{O}_8^{2-}]^x [\text{I}^-]^y$$

$$y = 1$$

$$x = 1$$

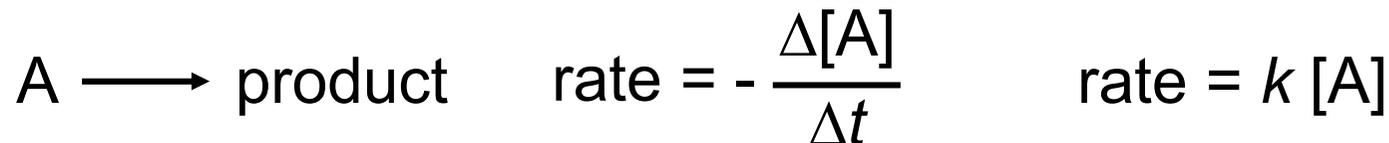
$$\text{rate} = k [\text{S}_2\text{O}_8^{2-}] [\text{I}^-]$$

Double  $[\text{I}^-]$ , rate doubles (experiment 1 & 2)

Double  $[\text{S}_2\text{O}_8^{2-}]$ , rate doubles (experiment 2 & 3)

$$k = \frac{\text{rate}}{[\text{S}_2\text{O}_8^{2-}] [\text{I}^-]} = \frac{2.2 \times 10^{-4} \text{ M/s}}{(0.08 \text{ M})(0.034 \text{ M})} = 0.08/\text{M}\cdot\text{s}$$

# 12.4 The integrated Rate Law



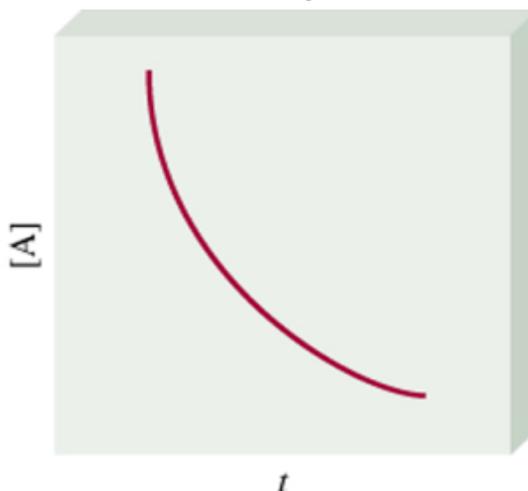
## First-Order Reactions

Unit of rate constant:

$$k = \frac{\text{rate}}{[A]} = \frac{M/s}{M} = 1/s \text{ or } s^{-1}$$

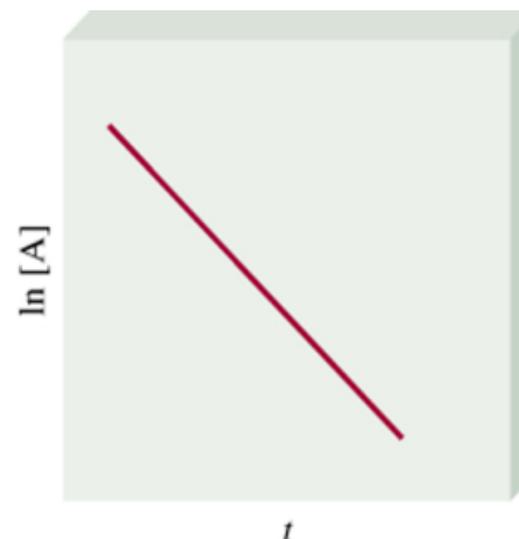
Integrated Form of the first order rate equation:

$$[A] = [A]_0 \exp(-kt)$$



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

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



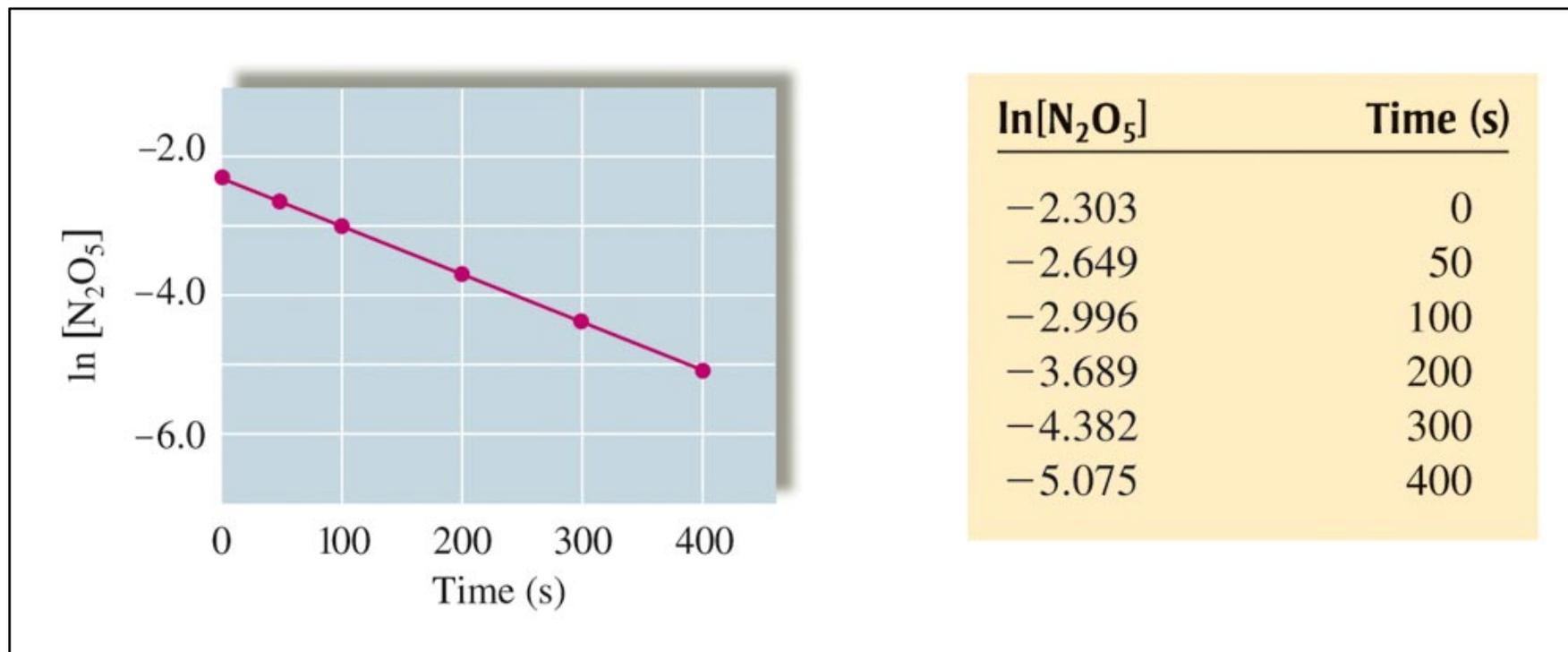
[A] is the concentration of A at any time  $t$

[A]<sub>0</sub> is the concentration of A at time  $t=0$

## First-Order Reactions



Figure 12.4 A Plot of  $\ln[\text{N}_2\text{O}_5]$  versus Time



# First-Order Reactions



The reaction  $2A \longrightarrow B$  is first order in A with a rate constant of  $2.8 \times 10^{-2} \text{ s}^{-1}$  at  $80^\circ\text{C}$ . How long will it take for A to decrease from  $0.88 \text{ M}$  to  $0.14 \text{ M}$  ?

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

$$[A]_0 = 0.88 \text{ M}$$

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

$$[A] = 0.14 \text{ M}$$

$$t = \frac{\ln[A]_0 - \ln[A]}{k} = \frac{\ln \frac{[A]_0}{[A]}}{k} = \frac{\ln \frac{0.88 \cancel{\text{M}}}{0.14 \cancel{\text{M}}}}{2.8 \times 10^{-2} \text{ s}^{-1}} = 66 \text{ s}$$

# Half-life time of the First-Order Reactions

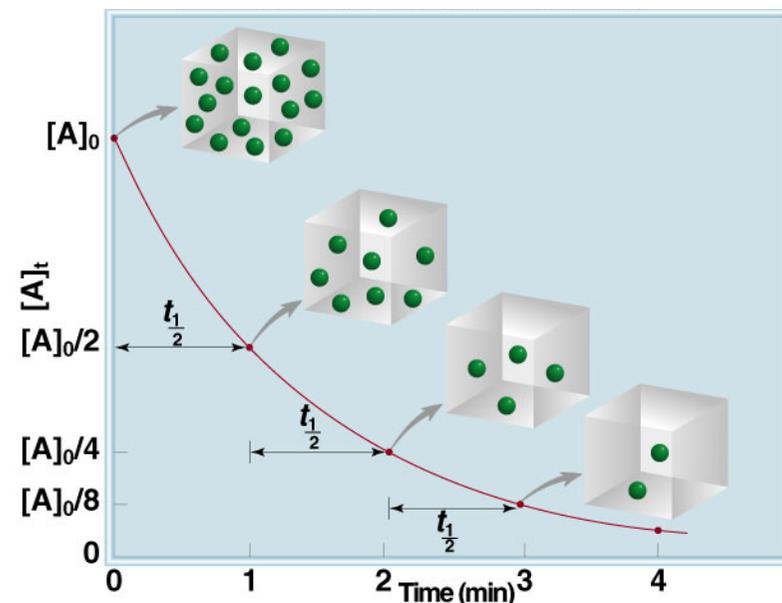


The **half-life**,  $t_{1/2}$ , is the time required for the concentration of a reactant to decrease to half of its initial concentration.

$$t_{1/2} = t \text{ when } [A] = [A]_0/2$$

$$t = \frac{\ln[A]_0 - \ln[A]}{k} = \frac{\ln \frac{[A]_0}{[A]}}{k}$$

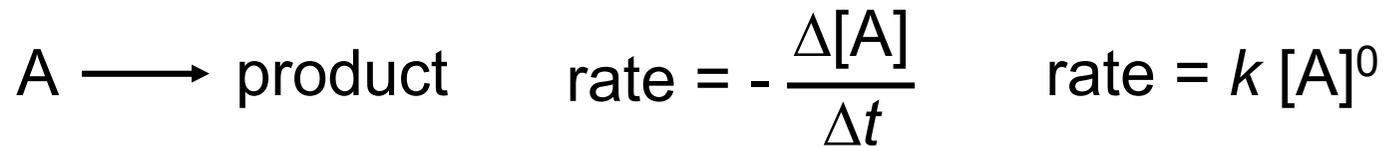
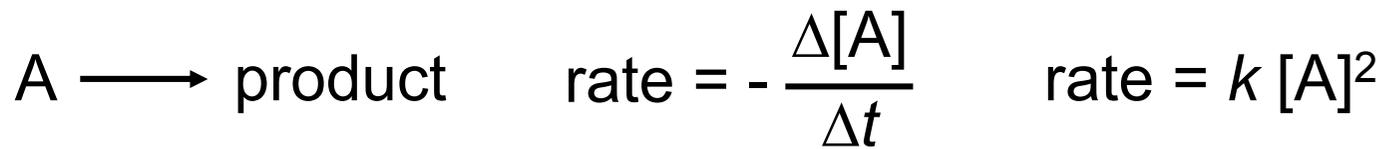
$$t_{1/2} = \frac{\ln \frac{[A]_0}{[A]_0/2}}{k} = \frac{\ln 2}{k} = \frac{0.693}{k}$$



What is the half-life of  $\text{N}_2\text{O}_5$  if it decomposes with a rate constant of  $5.7 \times 10^{-4} \text{ s}^{-1}$ ?

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{5.7 \times 10^{-4} \text{ s}^{-1}} = 1200 \text{ s} = 20 \text{ minutes}$$

How do you know decomposition is first order? units of  $k$  ( $\text{s}^{-1}$ )



## Second-Order Reactions      Zero-Order Reactions

Unit of rate constant

$$k = \frac{\text{rate}}{[A]^2} = \frac{M/s}{M^2} = 1/M \cdot s$$

$$k = \frac{\text{rate}}{[A]^0} = M/s$$

## Integrated Form of the first order rate equation

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

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

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

$$[A] = [A]_0 - kt$$

$$t_{1/2} = t \quad \text{when} \quad [A] = [A]_0/2$$

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

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

# Graphical Methods



## Equation for a Straight Line

$$y = bx + a \quad \left\{ \begin{array}{l} b = \text{slope} \\ a = \text{y intercept} \\ x = \text{time} \end{array} \right.$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{1st order}$$

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0} \quad \text{2nd order}$$

$$[A]_t = kt + [A]_0 \quad \text{Zero order}$$



# 12.5 Rate Laws: A Summary



- The used laws contain reactant concentration only
- Rate types: Differential or Integral rate laws
- The common experimental methods:
  - For differential rate law is the initial concentration method
  - For integral rate law graphical method
- For reactions with several reactants the integral method can be applied only if the concentration of one reactant only can be treated (Pseudo rate law), i.e.

$$\text{Rate} = k[A]^n[B]^m[C]^p$$

To be reduced to

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

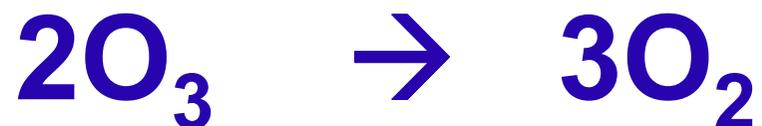
Order	Rate Law	Concentration-Time Equation	Half-Life
0	rate = $k$	$[A] = [A]_0 - kt$	$t_{1/2} = \frac{[A]_0}{2k}$
1	rate = $k [A]$	$\ln[A] = \ln[A]_0 - kt$	$t_{1/2} = \frac{\ln 2}{k}$
2	rate = $k [A]^2$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$	$t_{1/2} = \frac{1}{k[A]_0}$



## 12.6 Reaction Mechanisms



Many Chemical reactions occur by a series of steps called the **Reaction Mechanism**.



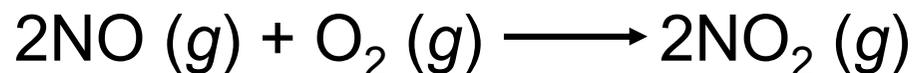
**Intermediates** are species that appear in a reaction mechanism **but not** in the overall balanced equation.

# Reaction Mechanisms

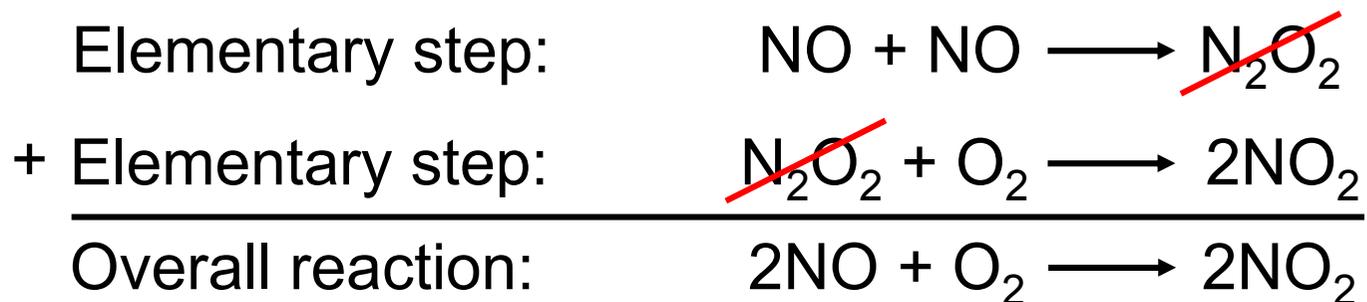


The overall progress of a chemical reaction can be represented at the molecular level by a series of simple ***elementary steps*** or ***elementary reactions***.

The sequence of **elementary steps** that leads to product formation is the ***reaction mechanism***.



$\text{N}_2\text{O}_2$  is detected during the reaction!





# 12.6 Reaction Mechanisms



- Elementary Step

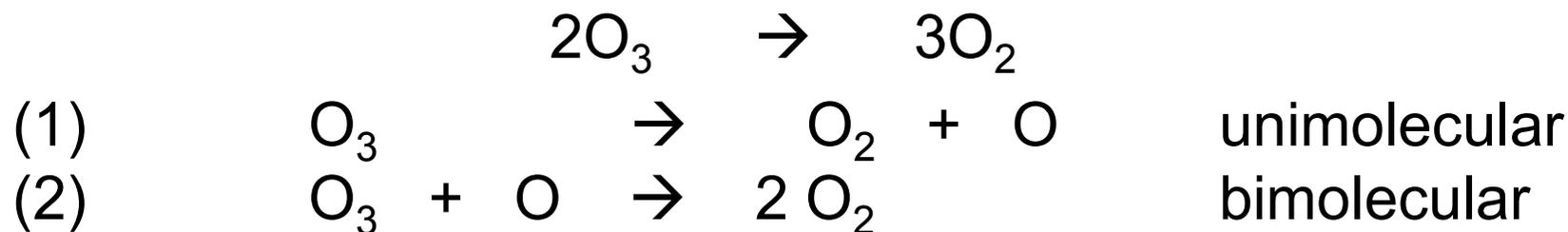
equation describing a single molecular event

- Molecularity

- Unimolecular      rate =  $k[A]$

- Bimolecular      rate =  $k[A][B]$     or    rate =  $k[A]^2$

- Termolecular      rate =  $k[A]^2[B]$





## 12.6 Reaction Mechanisms



The proposed **Reaction Mechanism** will be corrected if it satisfies the following two requirements:

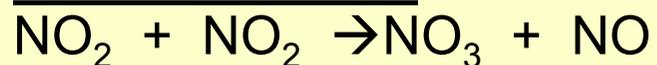
- The sum of the elementary steps give the overall balanced equation.
- The rate of the mechanism agrees the experimental rate law.

For example: Given the following reaction and rate law:



- Does the reaction occur in a single step?
- Which of the following two mechanisms is most likely:

Mechanism A:

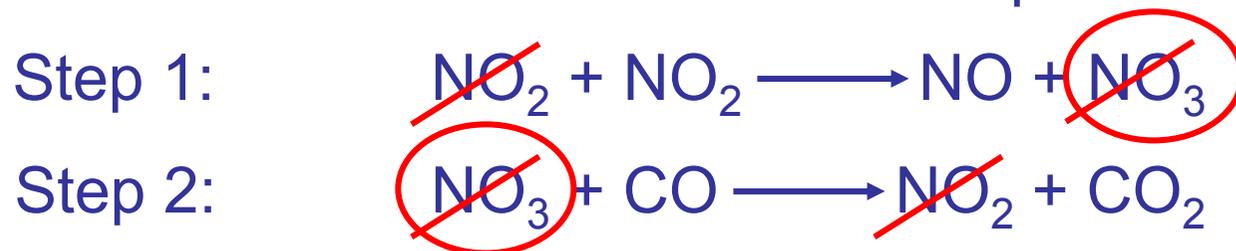


Mechanism B:





The experimental rate law for the reaction between  $\text{NO}_2$  and  $\text{CO}$  to produce  $\text{NO}$  and  $\text{CO}_2$  is  $\text{rate} = k[\text{NO}_2]^2$ . The reaction is believed to occur via two steps:



What is the equation for the overall reaction?



What is the intermediate?



What can you say about the relative rates of steps 1 and 2?

rate =  $k[\text{NO}_2]^2$  is the rate law for step 1 so  
step 1 must be slower than step 2



## 12.7 A Model for Chemical Kinetics



### Temperature Effects

- Rates typically increase with T increase
- Collisions between molecules increase
- Energy of collisions increase
- Even though only a small fraction of collisions lead to reaction
- Minimum Energy necessary for reaction is the Activation Energy



# Molecular Theory

(Collision Theory)



- **Activation Energy varies greatly**

- almost zero to hundreds of kJ
- size of  $E_a$  affects reaction rates

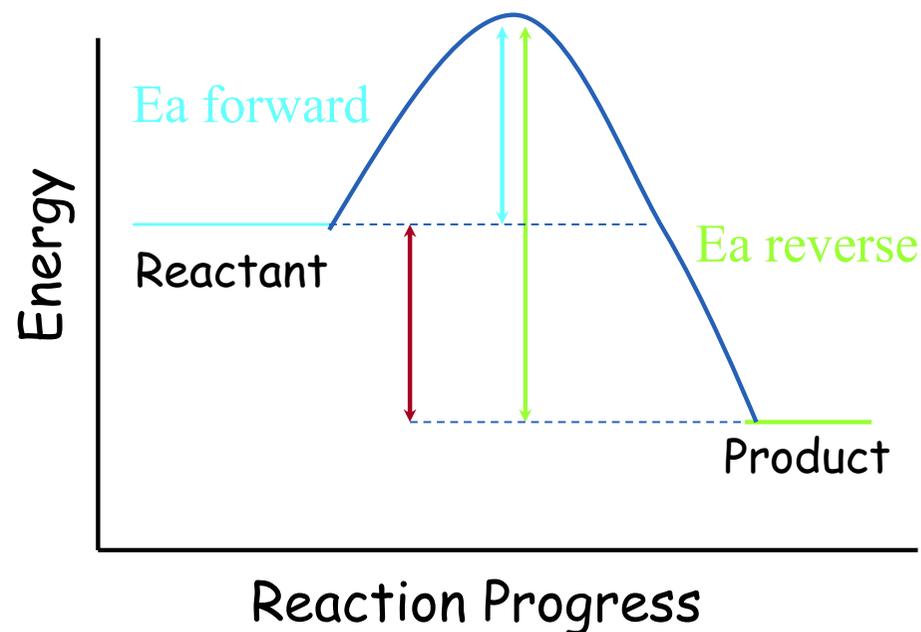
- **Concentration**

- more molecules, more collisions

- **Molecular Orientation**

- collisions must occur “sterically”

## Activation Energy, $E_a$



$$\Delta H_{\text{rxn}} = E_{a \text{ forward}} - E_{a \text{ reverse}}$$



# The Arrhenius Equation



- increase temperature, increase reaction rates
- reaction rates are  $\alpha$  to energy, collisions, temperature & orient

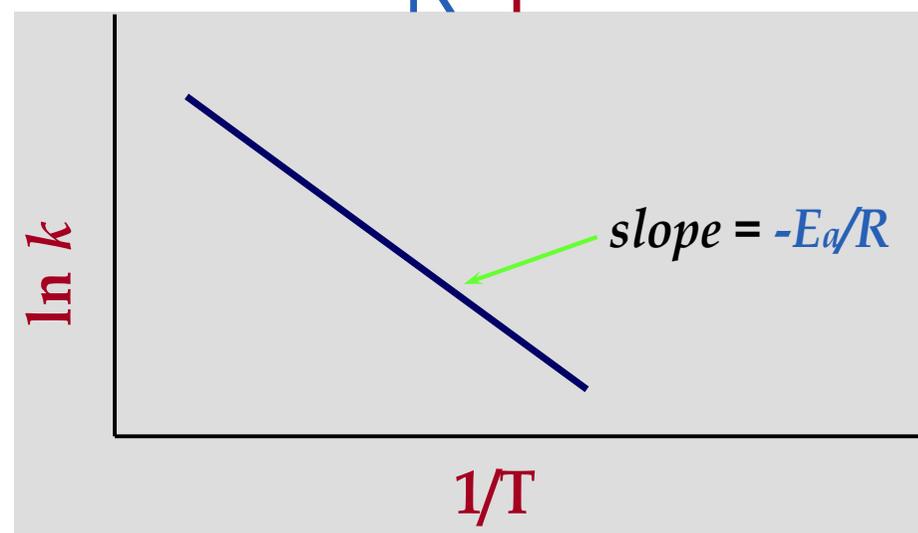
$$k = A e^{-E_a/RT}$$

$k$  = rate constant

$A$  = frequency of collisions

$-E_a/RT$  = fraction of molecules with energy necessary for reaction

$$y = bx + a$$
$$\ln k = \frac{-E_a}{R} \frac{1}{T} + \ln A$$



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

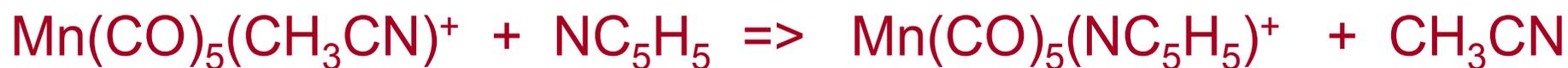


# Problem:



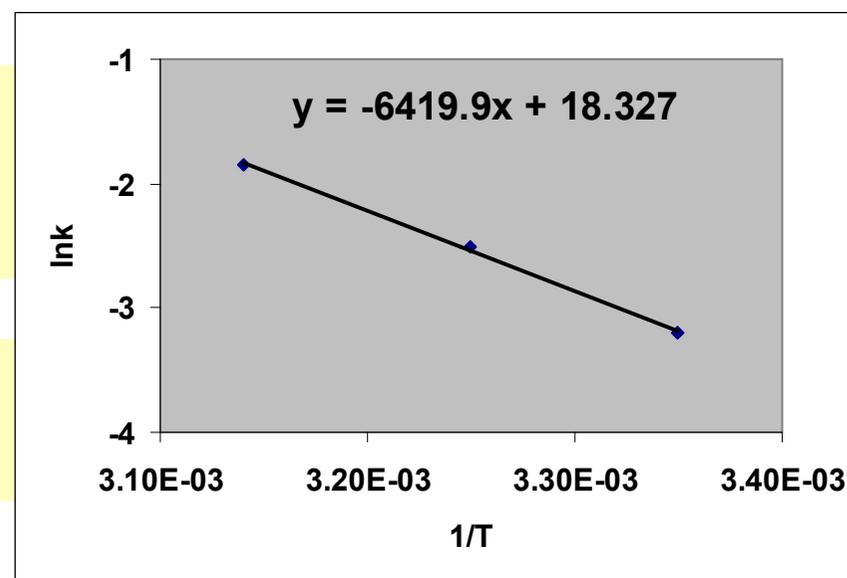
Data for the following reaction are listed in the table. Calculate  $E_a$  graphically, calculate  $A$  and find  $k$  at 311 K.

$\ln k$	$k, \text{min}^{-1}$	$T \text{ (K)}$	$1/T \times 10^{-3}$
-3.20	0.0409	298	3.35
-2.50	0.0818	308	3.25
-1.85	0.157	318	3.14



$$\begin{aligned} \text{slope} &= -6419.9 = -E_a/R \\ E_a &= (-6419.9)(-8.31 \times 10^{-3} \text{ kJ/K mol}) \\ &= 53.3 \text{ kJ} \end{aligned}$$

$$\begin{aligned} y \text{ intercept} &= 18.327 = \ln A \\ A &= 8.0 \times 10^7 \end{aligned}$$





# Problem:



The energy of activation for



is 260 kJ/mol at 800 K and  $k = 0.0315 \text{ sec}^{-1}$

Find  $k$  at 850 K.

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

$$k \text{ at } 850 \text{ K} = 0.314 \text{ sec}^{-1}$$



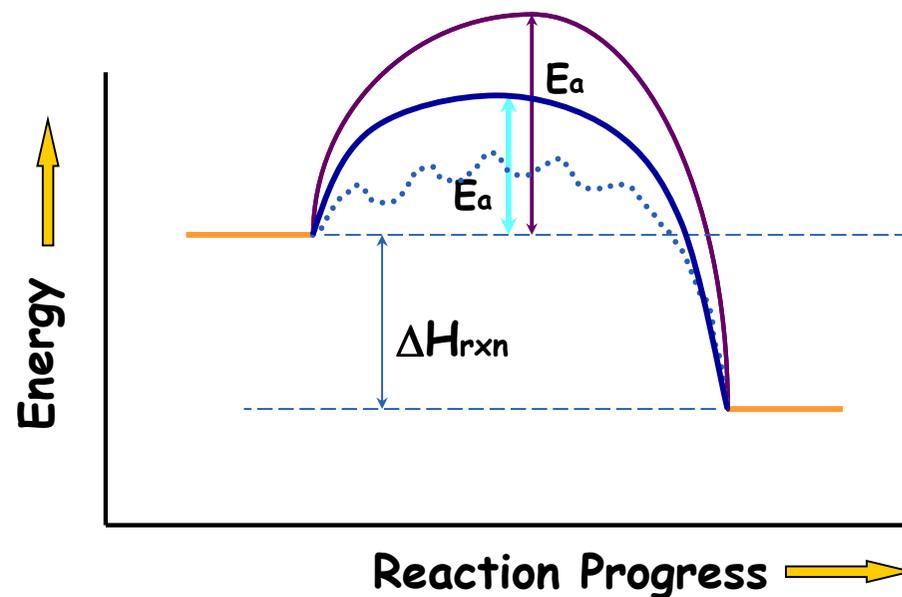
# 12.8 Catalysis



A **catalyst** is a substance that increases the rate of a chemical reaction without itself being consumed.

- homogeneous
- Heterogeneous

Facilitates the progress of a reaction by lowering the overall activation energy



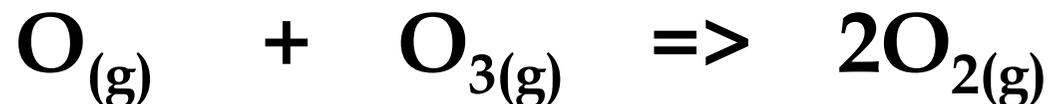
$$k = A \cdot \exp(-E_a/RT) \quad E_a \downarrow \quad k \uparrow \quad \text{rate}_{\text{catalyzed}} > \text{rate}_{\text{uncatalyzed}}$$



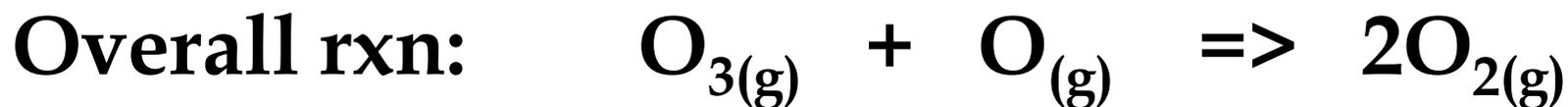
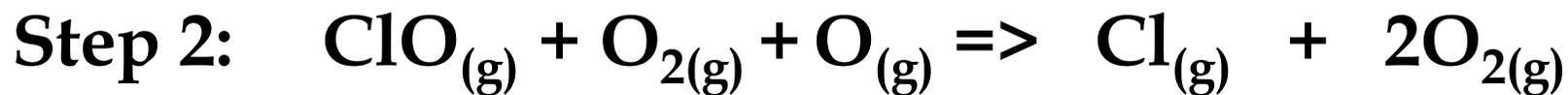
# 12.8 Catalysis



## Uncatalyzed Reaction:



## Catalyzed Reaction:



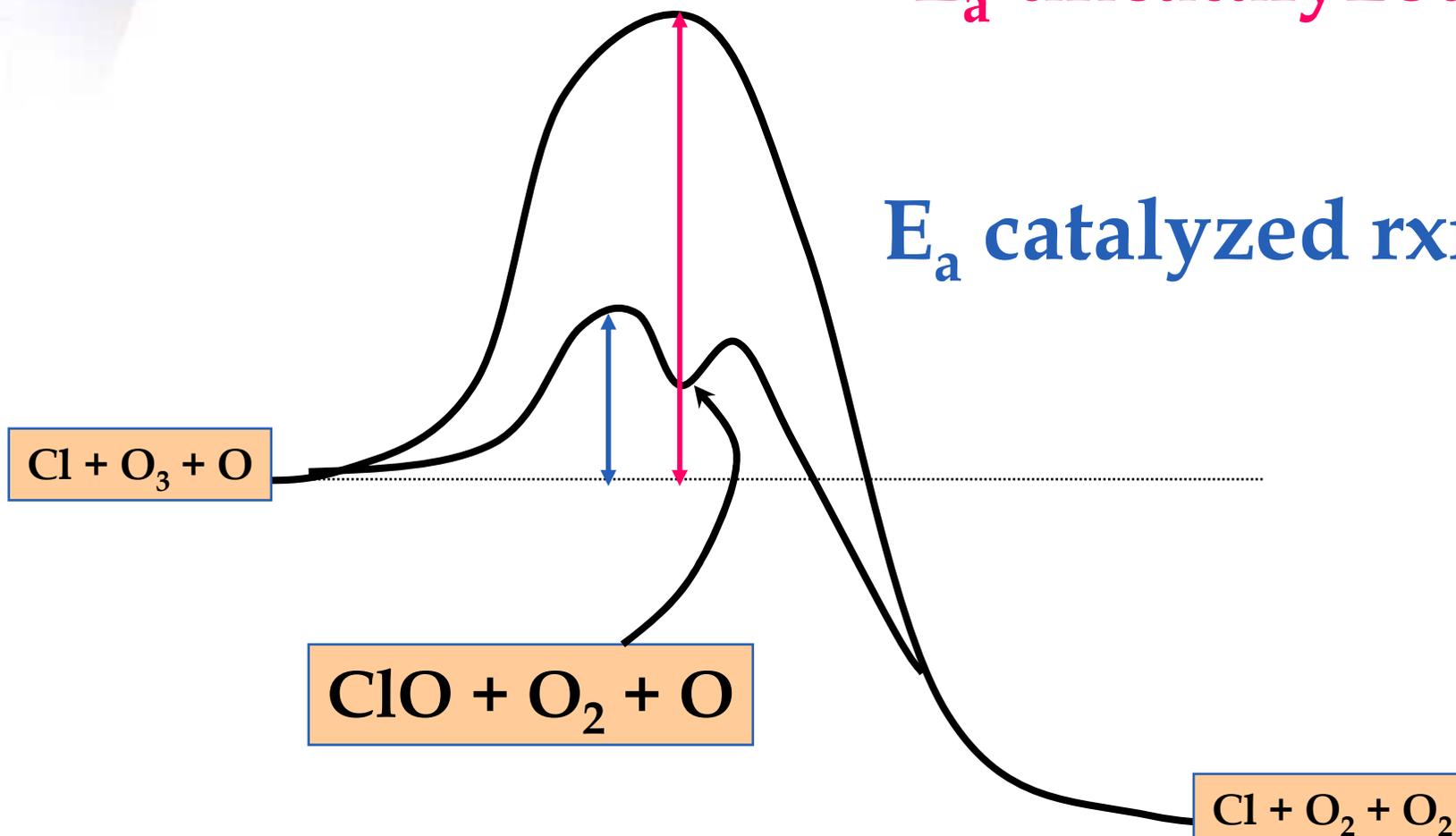


# 12.8 Catalysis



$E_a$  uncatalyzed rxn

$E_a$  catalyzed rxn





## 12.8 Examples



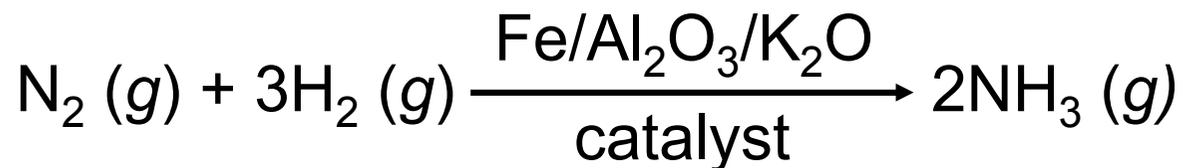
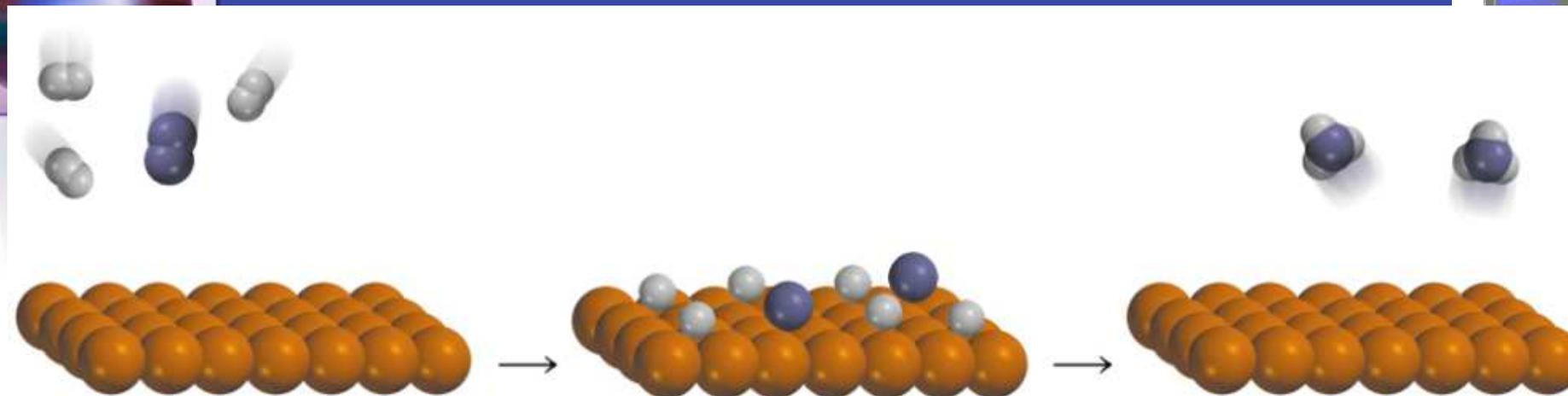
In ***heterogeneous catalysis***, the reactants and the catalysts are in different phases.

- Haber synthesis of ammonia
- Ostwald process for the production of nitric acid
- Catalytic converters

In ***homogeneous catalysis***, the reactants and the catalysts are dispersed in a single phase, usually liquid.

- Acid catalysis
- Base catalysis

# Haber Process

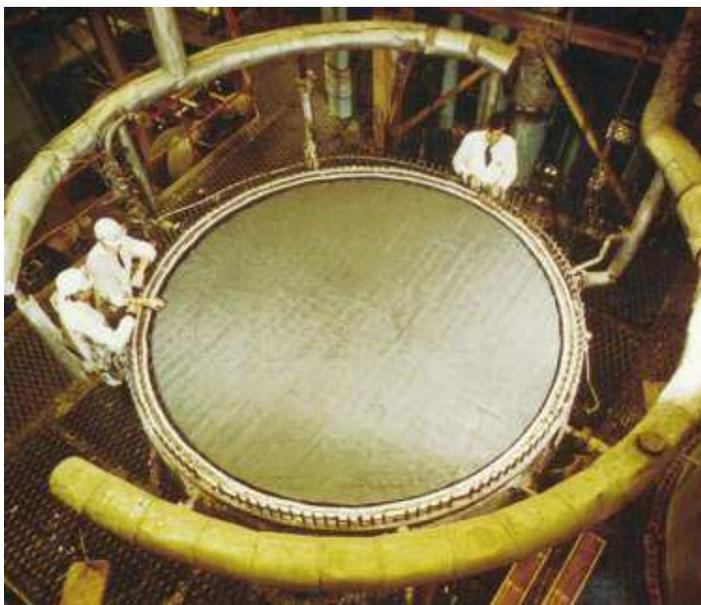
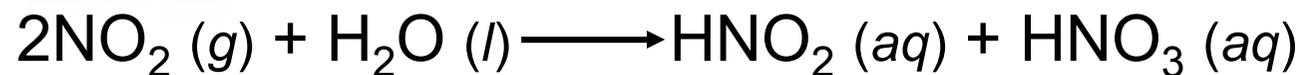
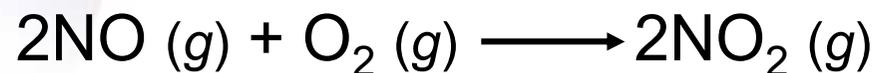


Heterogeneous Catalysis involves:

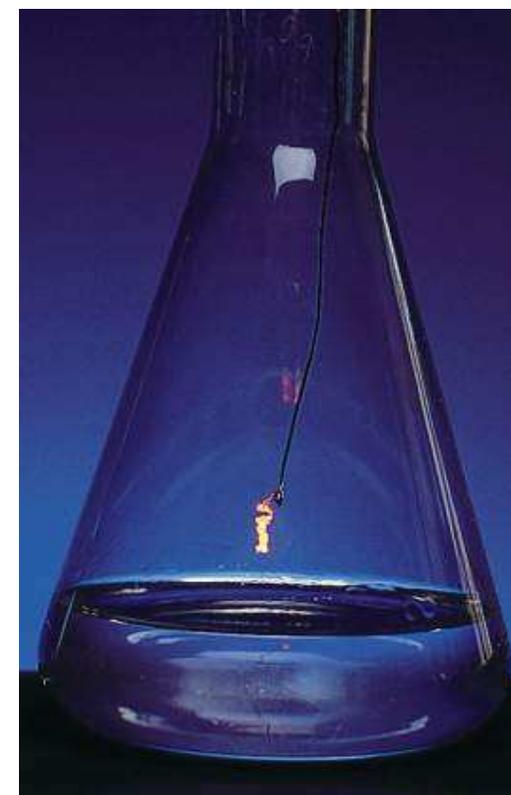
- ✓ Adsorption and activation of reactants
- ✓ Migration of the adsorbed reactants on the surface
- ✓ Reaction of the adsorbed substances
- ✓ Escape, or desorption, of the product



# Ostwald Process

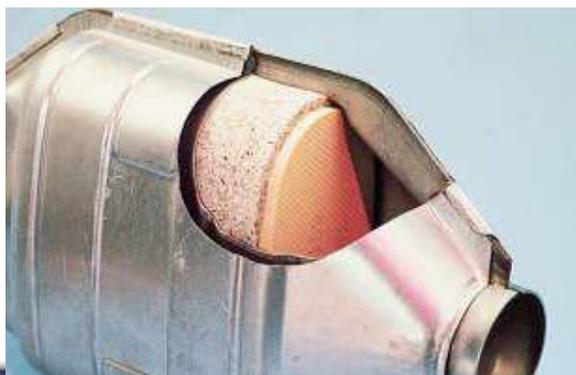
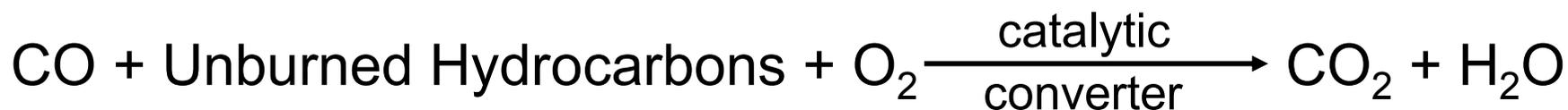
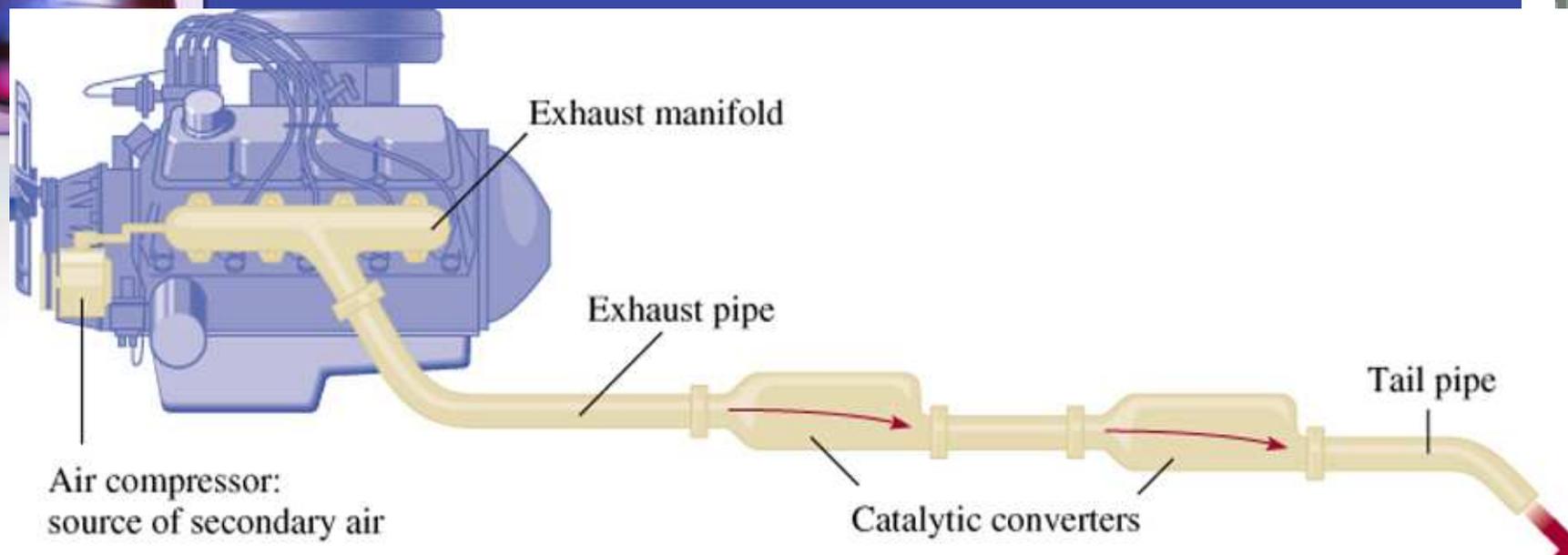


Pt-Rh catalysts used  
in Ostwald process

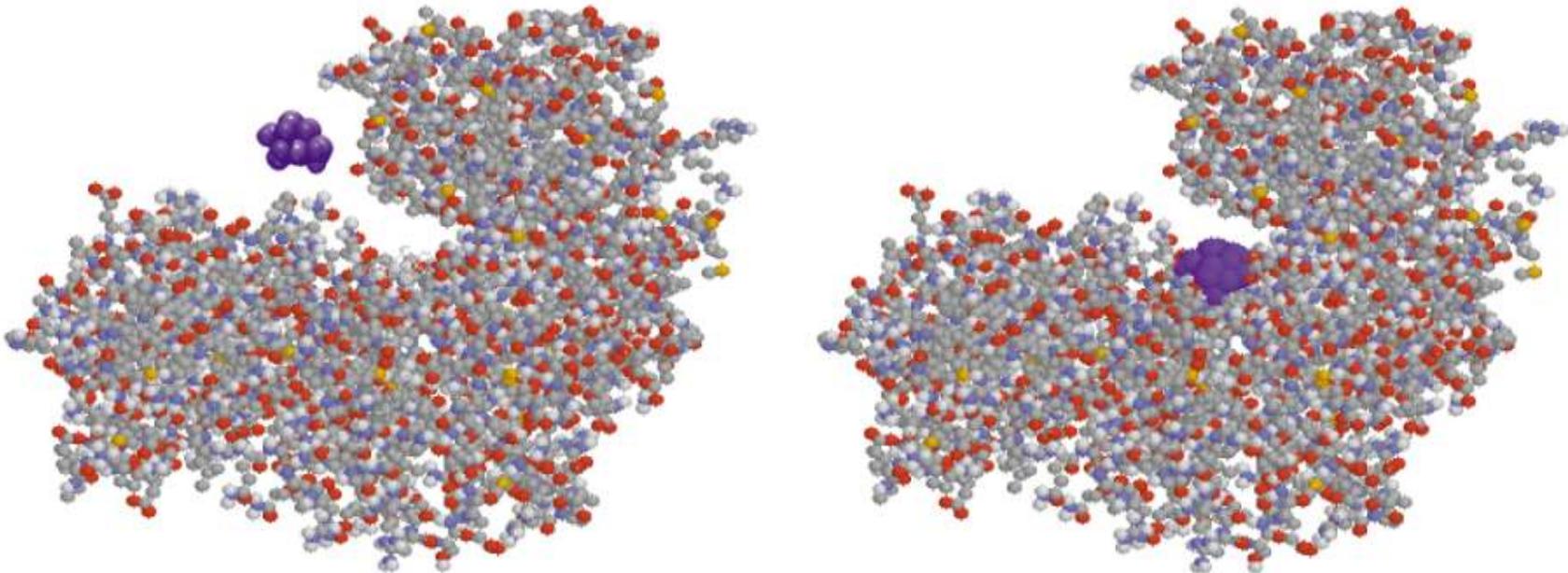
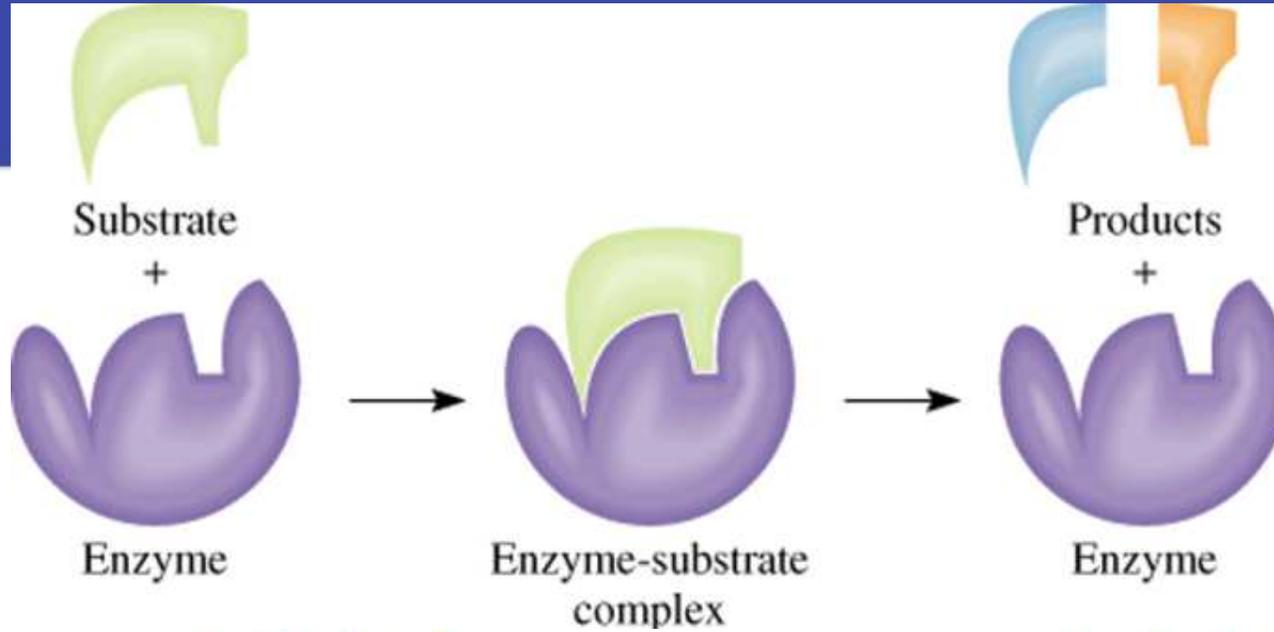


Hot Pt wire  
over  $\text{NH}_3$  solution

# Catalytic Converters



# Enzyme Catalysis





# Review Problems:



- Consider the reaction  $X \rightarrow Y + Z$

Which of the following is a possible rate law?

- a) Rate =  $k[X]$
- b) Rate =  $k[Y]$
- c) Rate =  $k[Y][Z]$
- d) Rate =  $k[X][Y]$
- e) Rate =  $k[Z]$

**ANS:** a)      **SECTION:** 12.2      **LEVEL:** easy



# Review Problems:



- For a reaction in which A and B react to form C, the following initial rate data were obtained:

[A] (mol/L)	[B] (mol/L)	Initial Rate of Formation of C (mol/L × s)
0.10	0.10	1.00
0.10	0.20	4.00
0.20	0.20	8.00

What is the rate law for the reaction?

- a) Rate =  $k[A][B]$
- b) Rate =  $k[A]^2[B]$
- c) Rate =  $k[A][B]^2$
- d) Rate =  $k[A]^2[B]^2$
- e) Rate =  $k[A]^3$

**ANS:**c)

**SECTION:** 12.3

**LEVEL:** medium



# Review Problems:



- A first-order reaction is 45% complete at the end of 35 minutes. What is the length of the half-life of this reaction?
  - a) 41 min
  - b) 39 min
  - c) 30. min
  - d) 27 min
  - e) none of these

**ANS:** a)      **SECTION:** 12.4      **LEVEL:** medium



# Review Problems:



- The reaction  $2\text{NO} \rightarrow \text{N}_2 + \text{O}_2$  has the following rate law:

$$-\Delta[\text{NO}]/\Delta t = 2k[\text{NO}]^2.$$

After a period of  $2.0 \times 10^3$  s, the concentration of NO falls from an initial value of  $2.8 \times 10^{-3}$  mol/L to  $2.0 \times 10^{-3}$  mol/L. What is the rate constant,  $k$ ?

- a)  $7.2 \times 10^{-2} \text{ M}^{-1}/\text{s}$
- b)  $1.7 \times 10^{-4} \text{ M}^{-1}/\text{s}$
- c)  $4.0 \times 10^{-4} \text{ M}^{-1}/\text{s}$
- d)  $4.0 \times 10^{-7} \text{ M}^{-1}/\text{s}$
- e)  $3.6 \times 10^{-2} \text{ M}^{-1}/\text{s}$

**ANS:**e)      **SECTION:** 12.4      **LEVEL:** medium



# Review Problems:



- The reaction  $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$  obeys the rate law  
 $-\Delta[\text{NO}]/\Delta t = k_{\text{obsd}}[\text{NO}]^2[\text{O}_2]$ .

Which of the following mechanisms is consistent with the experimental rate law?

- a)  $\text{NO} + \text{NO} \rightarrow \text{N}_2\text{O}_2$  (slow)  
 $\text{N}_2\text{O}_2 + \text{O}_2 \rightarrow 2\text{NO}_2$  (fast)
- b)  $\text{NO} + \text{O}_2 \rightleftharpoons \text{NO}_3$  (fast equilibrium)  
 $\text{NO}_3 + \text{NO} \rightarrow 2\text{NO}_2$  (slow)
- c)  $2\text{NO} \rightleftharpoons \text{N}_2\text{O}_2$  (fast equilibrium)  
 $\text{N}_2\text{O}_2 \rightarrow \text{NO}_2 + \text{O}$  (slow)  
 $\text{NO} + \text{O} \rightarrow \text{NO}_2$  (fast)
- d)  $\text{O}_2 + \text{O}_2 \rightarrow \text{O}_2 + \text{O}_2$  (slow)  
 $\text{O}_2 + \text{NO} \rightarrow \text{NO}_2 + \text{O}$  (fast)  
 $\text{O} + \text{NO} \rightarrow \text{NO}_2$  (fast)
- e) none of these

**ANS:**b)

**SECTION:** 12.6

**LEVEL:** medium



# Review Problems:



- Use the following information to determine the activation energy for the reaction shown here:



Temperature (K)	Rate Constant (L/mol × s)
1400	0.143
1500	0.659

- a)  $3.2 \times 10^4$  J/mol
- b)  $9.5 \times 10^6$  J/mol
- c)  $2.8 \times 10^4$  J/mol
- d)  $6.8 \times 10^5$  J/mol
- e)  $2.7 \times 10^5$  J/mol

**ANS:**e)

**SECTION:** 12.7

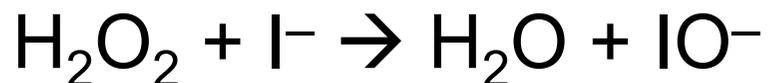
**LEVEL:** easy



# Review Problems:



- The reaction  $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$  has the following mechanism:



The catalyst in the reaction is:

- a)  $\text{H}_2\text{O}$
- b)  $\text{I}^-$
- c)  $\text{H}_2\text{O}_2$
- d)  $\text{IO}^-$

**ANS:**b)  
easy

**SECTION:** 12.6,8

**LEVEL:**