

# **CHAPTER 1: REACTION KINETICS**

**1.1 Reaction Rate**

**1.2 Collision Theory**

**1.3 Factors Affecting Reaction Rate**

A photograph of laboratory glassware. In the center is a clear beaker containing a blue liquid with a single red droplet suspended above it. To the left is a graduated cylinder filled with red liquid. To the right is a larger beaker containing yellow liquid. A red rubber bulb with a glass tube is positioned above the central beaker, with a single drop of red liquid falling from its tip.

# 1.1 REACTION RATE

# Reaction Rate

- Definition: Changes in **concentrations** of reactants (or products) as a function of **time**

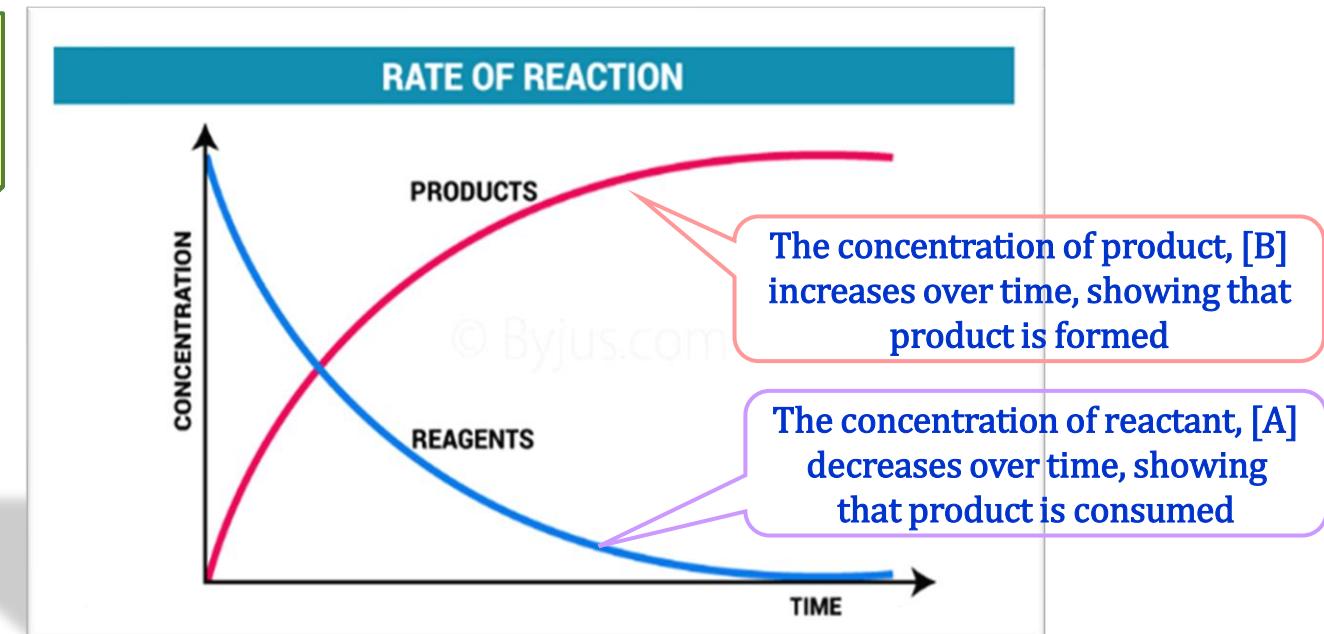
- Consider a reaction :



When reaction occur, the concentration of **A** is **decreased** while the concentration of **B** **increased** with time.

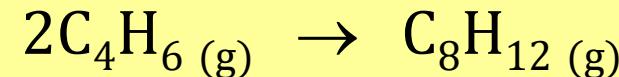
- As reaction proceed:

**[ REACTANT ] ↓ and [ PRODUCT ] ↑**

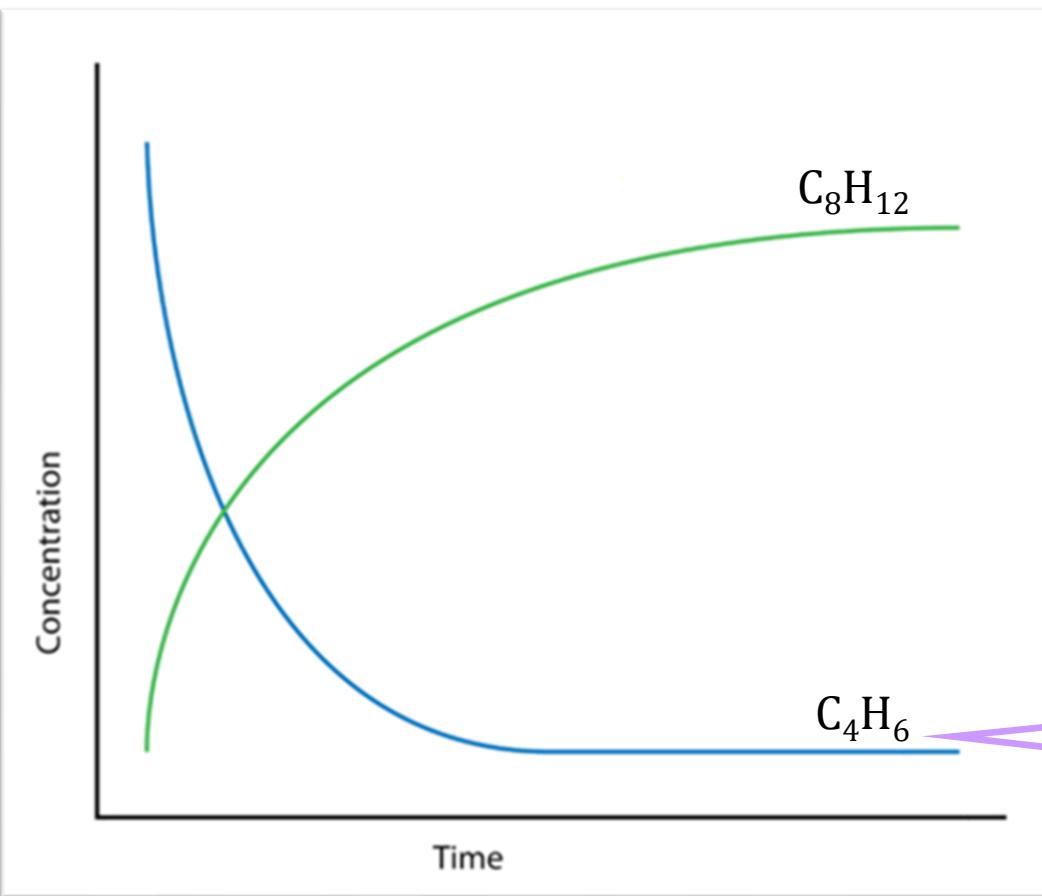


## Graphical interpretation:

Sketch a graph of concentration against time for the following reaction:



Solution:



$$\text{rate A} = \frac{\text{change in mole}}{\text{change in time}}$$

$$\text{rate} = \frac{(\text{final concentration}) - (\text{initial concentration})}{\text{time}}$$

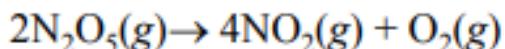
$$\text{rate} = \frac{\Delta(\text{concentration})}{\Delta t}$$

The reactant ( $\text{C}_4\text{H}_6$ ) is decreasing **two times faster** than the formation of product ( $\text{C}_8\text{H}_{12}$ )

## EXAMPLE

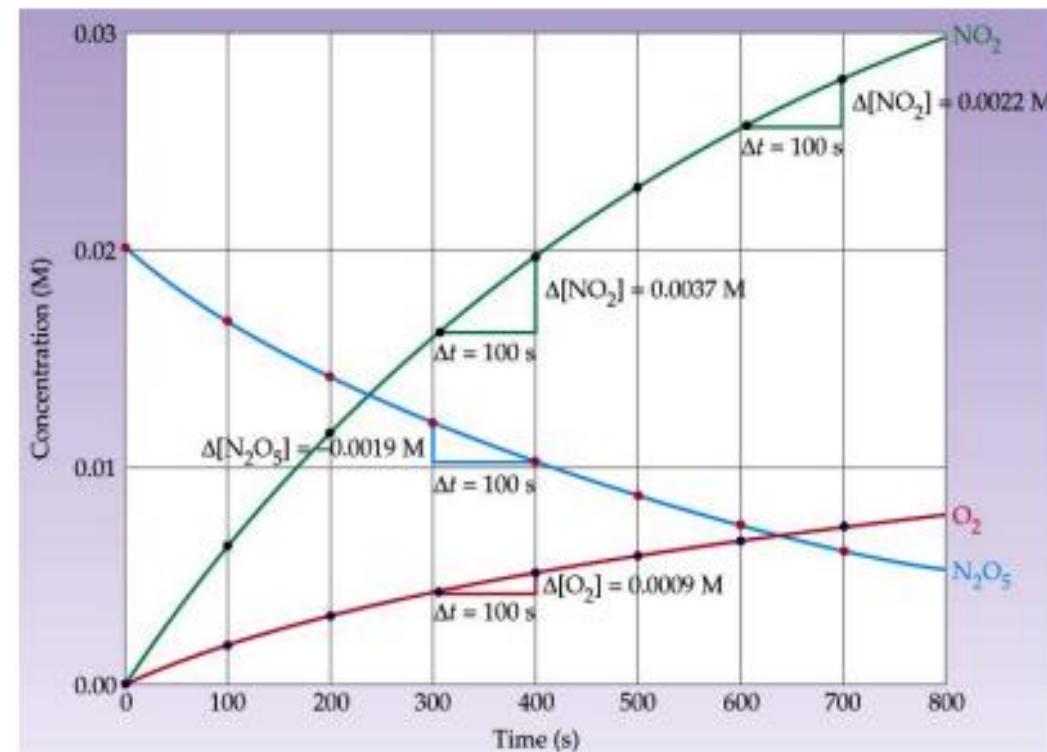
1

Consider the decomposition of  $\text{N}_2\text{O}_5$  to give  $\text{NO}_2$  and  $\text{O}_2$ :



Time (s)	Concentration (M)		
	$\text{N}_2\text{O}_5$	$\text{NO}_2$	$\text{O}_2$
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

reactants decrease with time      products increase with time



From the graph looking at t = 300 to 400 s

$$\text{Rate O}_2 = \frac{0.0009\text{M}}{100\text{s}} = 9 \times 10^{-6} \text{Ms}^{-1}$$

*Why do they differ?*

$$\text{Rate NO}_2 = \frac{0.0037\text{M}}{100\text{s}} = 3.7 \times 10^{-5} \text{Ms}^{-1}$$

*Recall:*

$$\text{Rate N}_2\text{O}_5 = \frac{0.0019\text{M}}{100\text{s}} = 1.9 \times 10^{-5} \text{Ms}^{-1}$$



To compare the rates one must account for the *stoichiometry*.

$$\text{Rate O}_2 = \frac{1}{1} \times 9 \times 10^{-6} \text{Ms}^{-1} = 9 \times 10^{-6} \text{Ms}^{-1}$$

$$\text{Rate NO}_2 = \frac{1}{4} \times 3.7 \times 10^{-5} \text{Ms}^{-1} = 9.2 \times 10^{-6} \text{Ms}^{-1}$$

*Now they agree!*

$$\text{Rate N}_2\text{O}_5 = \frac{1}{2} \times 1.9 \times 10^{-5} \text{Ms}^{-1} = 9.5 \times 10^{-6} \text{Ms}^{-1}$$

# Differential Rate Equation

- Shows the **relationship** between the rate of disappearance of reactants and formation of products.



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

- Negative sign indicates [A] and [B] **decreases** with time.
- Positive sign indicates [C] and [D] **increases** with time.

# KEEP IN MIND



$$\text{Rate} = -\frac{1}{2} \frac{d[A]}{dt}$$

$$\text{Rate} = +\frac{d[B]}{dt}$$

- Two moles of A disappear for each mole of B that formed

# Unit of Reaction Rate



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

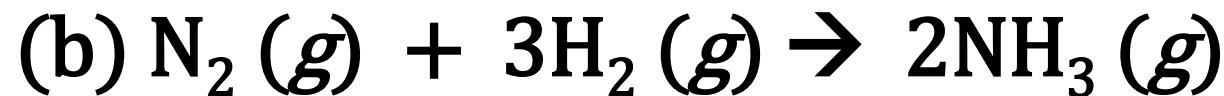
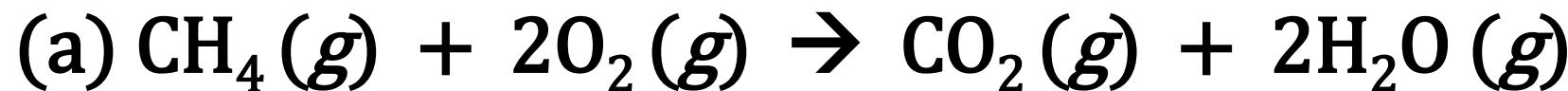
- Unit :** mol L<sup>-1</sup>s<sup>-1</sup> or mol dm<sup>-3</sup> s<sup>-1</sup> or M s<sup>-1</sup>
- Unit of time :** s (seconds), min, hr, day, year
- Example :** mol dm<sup>-3</sup> min<sup>-1</sup>

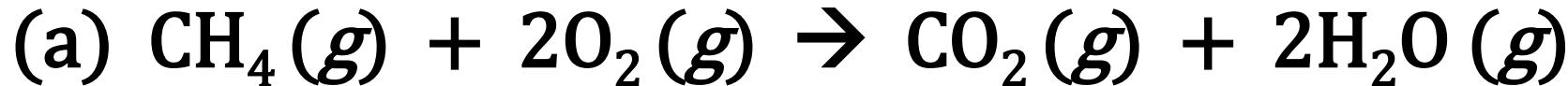


**EXAMPLE**

**2**

Write the differential rate equation for the following reaction.



**EXAMPLE** 2

$$\text{Rate} = -\frac{d[\text{CH}_4]}{dt} = -\frac{1}{2} \frac{d[\text{O}_2]}{dt} = +\frac{d[[\text{CO}_2]]}{dt} = +\frac{1}{2} \frac{d[\text{H}_2\text{O}]}{dt}$$

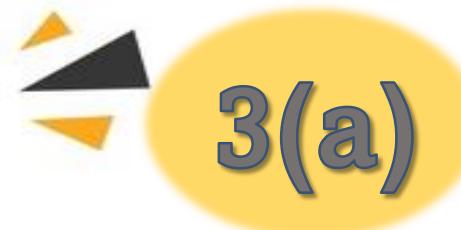


**EXAMPLE****3**

Dinitrogen pentoxide gas decomposes to nitrogen dioxide gas and oxygen gas according to the following balance equation below.



- a) Write the differential rate equation for them dissociation of  $\text{N}_2\text{O}_5$  gas.
- b) If the rate of disappearance of  $\text{N}_2\text{O}_5$  gas is  $1.0 \times 10^2 \text{ Ms}^{-1}$ , what is the rate of formation of  $\text{NO}_2$  gas under the same conditions of temperature and pressure?

**EXAMPLE****3(a)**

- Equation:  $2\text{N}_2\text{O}_5(g) \rightarrow 4\text{NO}_2(g) + \text{O}_2(g)$
- Differential rate equation:

$$(a) \quad \text{Rate} = -\frac{1}{2} \frac{d[\text{N}_2\text{O}_5]}{dt} = +\frac{1}{4} \frac{d[\text{NO}_2]}{dt} = +\frac{d[\text{O}_2]}{dt}$$

**Example 3**

Dinitrogen pentoxide gas decomposes to nitrogen dioxide gas and oxygen gas according to the following balance equation below.



- Write the differential rate equation for them dissociation of  $\text{N}_2\text{O}_5$  gas.
- If the rate of disappearance of  $\text{N}_2\text{O}_5$  gas is  $1.0 \times 10^2 \text{ Ms}^{-1}$ , what is the rate of formation of  $\text{NO}_2$  gas under the same conditions of temperature and pressure?

**EXAMPLE****3(b)**

- Calculate the rate of formation of  $\text{NO}_2$  gas =  $+\frac{d[\text{NO}_2]}{dt}$

**Given**

- Rate of disappearance of  $\text{N}_2\text{O}_5$  gas =  $-\frac{d[\text{N}_2\text{O}_5]}{dt}$   
 $= 1.0 \times 10^2 \text{ Ms}^{-1}$
- From the differential rate equation:

$$+\frac{1}{4} \frac{d[\text{NO}_2]}{dt} = -\frac{1}{2} \frac{d[\text{N}_2\text{O}_5]}{dt}$$

$$\begin{aligned} +\frac{d[\text{NO}_2]}{dt} &= -\frac{4}{2} \frac{d[\text{N}_2\text{O}_5]}{dt} \\ &= 2 \times (1.0 \times 10^2 \text{ Ms}^{-1}) \\ &= 2.0 \times 10^2 \text{ Ms}^{-1} \end{aligned}$$

**Example 3**

Dinitrogen pentoxide gas decomposes to nitrogen dioxide gas and oxygen gas according to the following balance equation below.



- a) Write the differential rate equation for them dissociation of  $\text{N}_2\text{O}_5$  gas.  
b) If the rate of disappearance of  $\text{N}_2\text{O}_5$  gas is  $1.0 \times 10^2 \text{ Ms}^{-1}$ , what is the rate of formation of  $\text{NO}_2$  gas under the same conditions of temperature and pressure?

# EXERCISE

1



Because it has a non-polluting product (water vapour), hydrogen gas is used as fuel aboard the space shuttle,



- Express the rate in terms of changes in  $[\text{H}_2]$ ,  $[\text{O}_2]$  and  $[\text{H}_2\text{O}]$  with time
- When  $[\text{O}_2]$  is decreasing at  $0.23 \text{ mol L}^{-1} \text{ s}^{-1}$ , at what rate is  $[\text{H}_2\text{O}]$  increasing?

Answer: b)  $0.46 \text{ mol L}^{-1} \text{ s}^{-1}$

# Rate Law/Rate equation

- An equation that relates the rate of a reaction to the **rate constant** and the **concentrations of the reactants raised to some powers**

Example:  $aA + bB \rightarrow cC + dD$

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

Where :

$[A]$  = concentration of A

$k$  = rate constant

$m$  = order of reactions with respect to A

$n$  = order of reaction with respect to B

$[B]$  = concentration of B

$m + n$  = overall order of reaction

# Order of Reaction



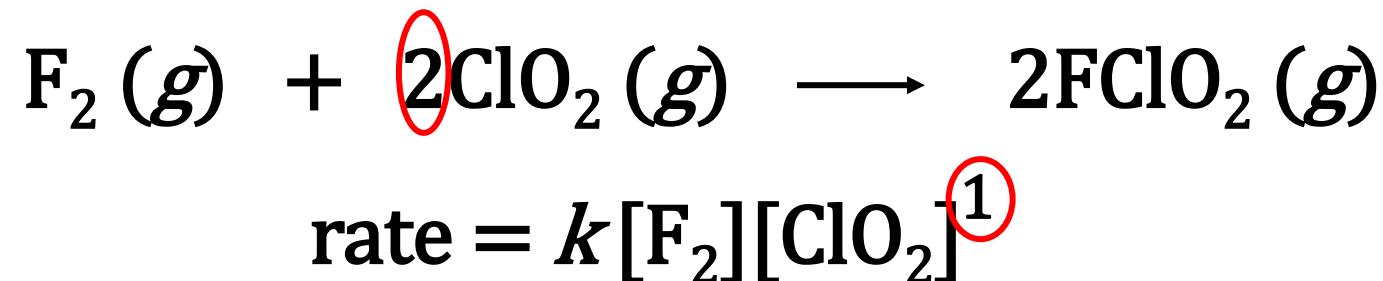
$$\text{Rate} = k[A]^x[B]^y$$

- Exponent of  $x$  (0, 1, 2, 3,  $\frac{1}{2}$ ,  $\frac{3}{4}$  and etc) and  $y$  (0, 1, 2, 3,  $\frac{1}{2}$ ,  $\frac{3}{4}$  and etc)
- Define how the rate affected by reactant concentration
  - The reaction orders cannot be predicted from the stoichiometry of the balanced equation; it must be determined experimentally.

# KEEP IN MIND

- Order of a reactant is **not related** to stoichiometric coefficient of reactant in the balanced chemical **equation**.

EXAMPLE:

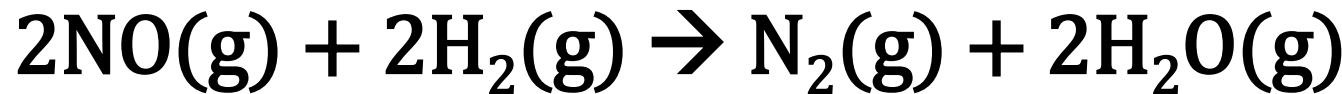


- Can only be determined **experimentally**
- May be **integral** (i.e., 1, 2, 3,...), **zero, fractional, decimal or /and negative**



## EXAMPLE

# 4



$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

The reaction is:

- Second order with respect to NO
- First order with respect to H<sub>2</sub>
- Overall reaction = third order reaction

# EXERCISE

## 2



For each of the following reactions, determine the reaction order with respect to each reactant and the overall order from the given rate law:



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



$$\text{Rate} = k[\text{CH}_3\text{CHO}]^{3/2}$$

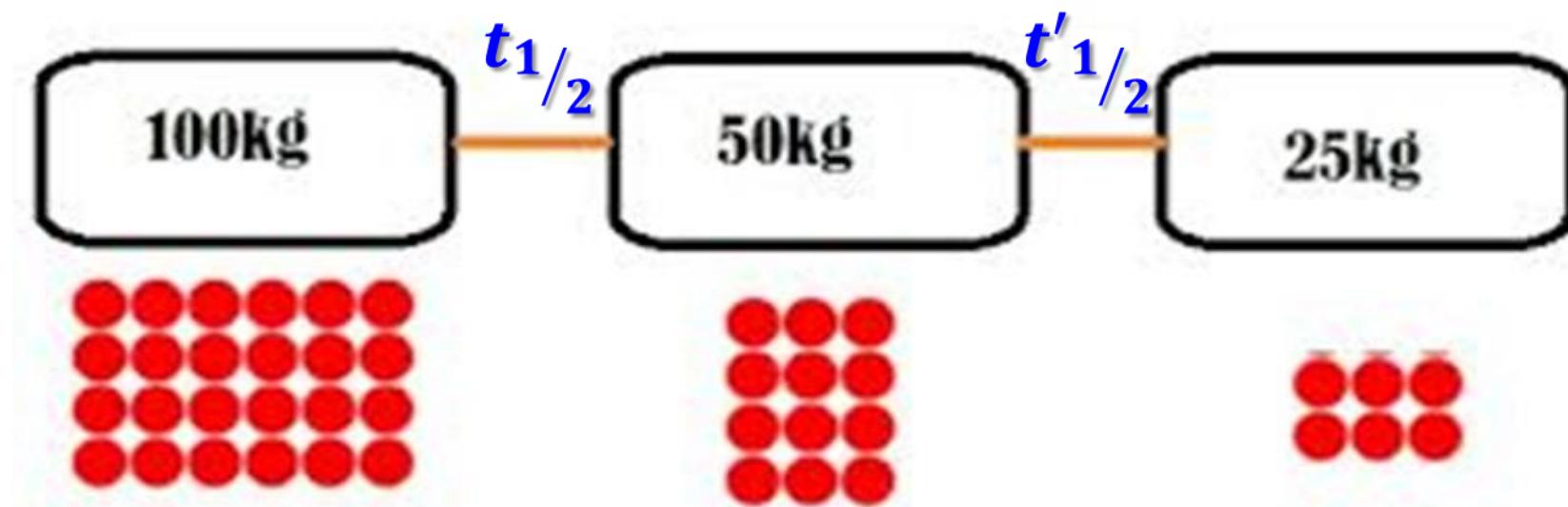


$$\text{Rate} = k[\text{H}_2\text{O}_2][\text{I}^-]$$

- a) second order in NO, first order in O<sub>2</sub>, third order overall   b) 3/2 order in CH<sub>3</sub>CHO, 3/2 order overall  
c) first order in H<sub>2</sub>O<sub>2</sub>, first order in I<sup>-</sup>, zero order in H<sup>+</sup>, second order overall

# Half-life( $t_{1/2}$ )

- Time required for the concentration of a reactant to decrease to half of its initial concentration



# $t_{1/2}$ for Zero Order Reaction

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

- $t_{1/2}$  directly proportional to the initial concentration of the reactant

# $t_{1/2}$ for First Order Reaction

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

- $t_{1/2}$  is independent of the initial concentration of the reactant.

# $t_{1/2}$ for Second Order Reaction

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

- ☐  $t_{1/2}$  inversely proportional to the initial concentration of the reactant.

**EXAMPLE**5

Determine the half-life of substance A if the rate constant is  $0.00624 \text{ M s}^{-1}$  and its original concentration is  $1.2 \text{ M}$ .

**EXAMPLE**5

- The unit of  $k = \text{M s}^{-1}$  ➔ Zero order reaction

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

$$t_{\frac{1}{2}} = \frac{1.2 \text{ M}}{2(0.00624 \text{ M s}^{-1})}$$

$$t_{\frac{1}{2}} = 96.15 \text{ s}$$

**Example 5** $\text{A} \rightarrow \text{products}$ 

Determine the half-life of substance A if the rate constant is  $0.00624 \text{ M s}^{-1}$  and its original concentration is  $1.2 \text{ M}$ .

# EXERCISE

3



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

$$t_{\frac{1}{2}} = 20.27 \text{ min}$$

# Rate Law for Zero Order Reaction

➤ If the reaction is;  $A \rightarrow \text{Products}$

$$\begin{aligned}\text{Rate} &= k[A]^0 \\ \text{Rate} &= k\end{aligned}$$

- Zero order in A
- Zero order overall
- Quite common in metal-catalyzed and biochemical processes
- Rate is **not dependent** on the concentration of the reactant

# Rate Law for Zero Order Reaction

□ If  $[A]$  is DOUBLE ➡ rate is NOT AFFECTED

$$[A] = [2A]$$

$$\text{Rate} = k[2A]^0$$

$$\text{Rate} = k$$

# Rate Law for First Order Reaction

➤ If the reaction is;  $A \rightarrow \text{Products}$

➤ Rate Law:  $\text{Rate} = k[A]$

- First order in A
- First order overall
- Example: Radioactive reaction  
(uranium, plutonium etc)

- Rate is dependent on the concentration of the reactant.
- Doubling  $[A]$ , the rate will increase by a factor of 2 / increase 2-fold / double.

# Rate Law for First Order Reaction

- If  $[A]$  is DOUBLE → rate is DOUBLE

$$[A] = [2A]$$

$$\text{Rate} = k[2A]$$

$$\text{Rate} = 2k[A]$$

# Rate Law for Second Order Reaction

- If the reaction is;  $A \rightarrow \text{Products}$

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

- Second order in A
- Second order overall

Rate is dependent on the concentration of the reactant.

Doubling  $[A]$ , the rate will increase by a  
factor of 4 / 4-fold

# Rate Law for Second Order Reaction

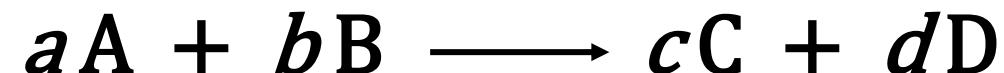
- If  $[A]$  is DOUBLE  rate is QUADRUPLES

$$[A] = [2A]$$

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

$$\text{Rate} = 4k[A]^2$$

# Rate Constant (*k*)



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

- Proportionality constant for a given reaction and temperature
- Does not change as reaction proceeds



## EXAMPLE

# 6

You can always  
determine  
the unit of  $k$   
**mathematically**

- What is the **unit of  $k$**  for a **third order reaction**?

$$\text{Rate} = k[A]^3$$

$$\frac{M}{s} = k[M]^3$$

$$k = M^{-2}s^{-1}$$

# EXERCISE

## 4



$A \rightarrow$  products

Without consulting textbooks, give the units of the rate constants for reactions with the following orders:  
*(Time: in s)*

- (a) First order
- (b) Second order
- (c) Third order
- (d) 5/2 order

a)  $k = s^{-1}$  b)  $k = M^{-1} s^{-1}$  c)  $k = M^{-2} s^{-1}$  d)  $k = M^{-3/2} s^{-1}$

# Overall Reaction Order From The Unit of Rate Constant, $k$

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

Overall Reaction Order	Units of $k$ ( $t$ in seconds)
0	$\text{Mol L}^{-1} \text{s}^{-1}$ or $\text{M s}^{-1}$
1	$\text{s}^{-1}$
2	$\text{L mol}^{-1} \text{s}^{-1}$ or $\text{M}^{-1} \text{s}^{-1}$

# **Integrated Rate Law**

- The relationship between the concentration of the reactants and time.
- May be used to determine the rate constant and half-life.

# Integrated Rate Law : Zero Order Reaction

If the reaction is;  $A \rightarrow \text{Products}$

$$[A]_o - [A]_t = kt \quad \text{OR} \quad [A]_t = [A]_o - kt$$

Where:

$[A]_o$  = Initial concentration

$[A]_t$  = Concentration at particular time

$t$  = Time of reaction

$k$  = Rate constant

# Integrated Rate Law : First Order Reaction

If the reaction is;  $A \rightarrow \text{Products}$

$$\ln \frac{[A]_o}{[A]_t} = kt \quad \text{OR} \quad \ln[A]_o - \ln[A]_t = kt \quad \text{OR} \quad \ln[A]_t = \ln[A]_o - kt$$

Where:

$[A]_o$  = Initial concentration

$[A]_t$  = Concentration at particular time

$t$  = Time of reaction

$k$  = Rate constant

# Integrated Rate Law : Second Order Reaction

If the reaction is;  $A \rightarrow \text{Products}$

$$\frac{1}{[A]_t} - \frac{1}{[A]_o} = kt \quad \text{OR} \quad \frac{1}{[A]_t} = kt + \frac{1}{[A]_o}$$

Where:

$[A]_o$  = Initial concentration

$[A]_t$  = Concentration at particular time

t = Time of reaction

k = Rate constant

**EXAMPLE****7**

A reaction starts with an initial concentration of reactant of 1.00 M and has a rate constant of  $1.65 \times 10^{-5} \text{ M s}^{-1}$ . What is the concentration of the reactant after 100 minute ?

**EXAMPLE****7**

The reaction is zero order reaction;  $[A]_t - [A]_o = -kt$

$$t = 100 \text{ minute} = 6000 \text{ s}$$

$$[A]_t = -kt + [A]_o$$

$$[A]_t = -(1.65 \times 10^{-5} \text{ M s}^{-1})(6000 \text{ s}) + 1.00 \text{ M}$$

$$[A]_t = 0.901 \text{ M}$$

Example 7

$A \rightarrow \text{products}$

A reaction starts with an initial concentration of reactant of 1.00 M and has a rate constant of  $1.65 \times 10^{-5} \text{ M s}^{-1}$ . What is the concentration of the reactant after 100 minute ?

**EXAMPLE****8**

$\text{N}_2\text{O}_5$  decomposes into  $\text{N}_2\text{O}_4$  and  $\text{O}_2$ .



At  $45^\circ\text{C}$ ,  $k = 6.22 \times 10^{-4} \text{ s}^{-1}$ . If the initial concentration of  $\text{N}_2\text{O}_5$  in a carbon tetrachloride solution at  $45^\circ\text{C}$  is  $0.500 \text{ M}$ , what will the concentration be after exactly one hour?

**EXAMPLE**8

- Equation:  $2\text{N}_2\text{O}_5(g) \rightarrow 2\text{N}_2\text{O}_4(g) + \text{O}_2(g)$
- The reaction is first-order overall;

$$\ln \frac{[\text{N}_2\text{O}_5]_0}{[\text{N}_2\text{O}_5]_t} = kt$$

- Given;  $k = 6.22 \times 10^{-4} \text{ s}^{-1}$   
 $t = 1 \text{ hr} \times 3600 \text{ s/hr} = 3600 \text{ s}$   
 $[\text{N}_2\text{O}_5]_0 = 0.500 \text{ M}$   
 $[\text{N}_2\text{O}_5]_t = ?$

**Example 8**

$\text{N}_2\text{O}_5$  decomposes into  $\text{N}_2\text{O}_4$  and  $\text{O}_2$ .



At  $45^\circ\text{C}$ ,  $k = 6.22 \times 10^{-4} \text{ s}^{-1}$ .

If the initial concentration of  $\text{N}_2\text{O}_5$  in a carbon tetrachloride solution at  $45^\circ\text{C}$  is  $0.500 \text{ M}$ , what will the concentration be after exactly one hour?

**EXAMPLE**

8



$$\ln \frac{[N_2O_5]_0}{[N_2O_5]_t} = kt$$

$$\ln \frac{0.500}{[N_2O_5]_t} = 6.22 \times 10^{-4} \text{ s}^{-1} \times 3600 \text{ s}$$

$$\frac{0.500}{[N_2O_5]_t} = 9.3858$$

$$[N_2O_5]_t = 0.0533 M$$

**Example 8**

$N_2O_5$  decomposes into  $N_2O_4$  and  $O_2$ .



At  $45^\circ\text{C}$ ,  $k = 6.22 \times 10^{-4} \text{ s}^{-1}$ .

If the initial concentration of  $N_2O_5$  in a carbon tetrachloride solution at  $45^\circ\text{C}$  is  $0.500 M$ , what will the concentration be after exactly one hour?

# EXERCISE

5



Iodine atoms combine to form molecular iodine in the gas phase



The rate constant of the reaction is  $7.0 \times 10^9 \text{ M}^{-1} \text{ s}^{-1}$  at  $23^\circ\text{C}$ . If the initial concentration of I was 0.086 M, calculate the concentration after 2.0 min.

$$[I]_t = 1.190 \times 10^{-12} \text{ M}$$

# **METHODS TO DETERMINE ORDER OF REACTION**



- Initial Rate Method
- Half-life Graph Method
- Linear Graph Method

# 1. Initial Rate Method

Use concentration measurement to find initial rate

Use initial rates from several experiments to find reaction orders

Use these values to calculate rate constant

**EXAMPLE****9**

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



Experiment	[F <sub>2</sub> ]	[ClO <sub>2</sub> ]	Initial Rate (M/s)
1	0.10	0.010	1.2 x 10 <sup>-3</sup>
2	0.10	0.040	4.8 x 10 <sup>-3</sup>
3	0.20	0.010	2.4 x 10 <sup>-3</sup>

**EXAMPLE**

9



- Rate law:

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

- ☞ Find order in  $\text{F}_2$  with  $[\text{ClO}_2]$  constant:

- Compare experiment 3 and 1

Experiment	$[\text{F}_2]$	$[\text{ClO}_2]$	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	$2.4 \times 10^{-3}$

$$\frac{\text{Rate 3}}{\text{Rate 1}} = \frac{k[\text{F}_2]^x[\text{ClO}_2]^y}{k[\text{F}_2]^x[\text{ClO}_2]^y} \quad \Rightarrow \quad \frac{2.4 \times 10^{-3}}{1.2 \times 10^{-3}} = \frac{k(0.2)^x(0.01)^y}{k(0.1)^x(0.01)^y}$$

$$\frac{2.4 \times 10^{-3}}{1.2 \times 10^{-3}} = \frac{k(0.2)^x}{k(0.1)^x}$$

$$(2.0)^1 = (2.0)^x$$

$$x = 1$$

- ☞ The reaction is first order with respect to  $\text{F}_2$

**EXAMPLE**

9



- ☞ Find order in  $\text{ClO}_2$  with  $[\text{F}_2]$  constant:
- Compare experiment 2 and 1

$$\frac{\text{Rate 3}}{\text{Rate 1}} = \frac{k[\text{F}_2]^x[\text{ClO}_2]^y}{k[\text{F}_2]^x[\text{ClO}_2]^y}$$

$$\frac{4.8 \times 10^{-3}}{1.2 \times 10^{-3}} = \frac{k(0.1)^x(0.04)^y}{k(0.1)^x(0.01)^y}$$



Experiment	$[\text{F}_2]$	$[\text{ClO}_2]$	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	$2.4 \times 10^{-3}$

$$\frac{4.8 \times 10^{-3}}{1.2 \times 10^{-3}} = \frac{k(0.04)^y}{k(0.01)^y}$$

$$(4.0)^1 = (4.0)^x$$

$$y = 1$$

- ☞ The reaction is first order with respect to  $\text{ClO}_2$

**EXAMPLE**

9

ANSWERS

☞ Rate =  $k[F_2]^x[ClO_2]^y$

$$x = 1, y = 1$$

$$\text{rate} = k[F_2][ClO_2]$$

☞ Use any experiment data (example exp. 1) to find  $k$

$$1.2 \times 10^{-3} \text{ M s}^{-1} = k(0.1 \text{ M})(0.01 \text{ M})$$

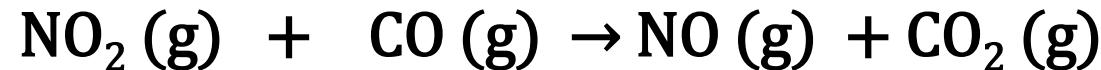
$$k = 1.2 \text{ M}^{-1} \text{ s}^{-1}$$

# EXERCISE

# 7



Consider the reaction between nitrogen dioxide,  $\text{NO}_2$  and carbon monoxide,  $\text{CO}$  :



The initial rate of the reaction is measured at several different concentrations of the reactants with the following results :

Experiment	[ $\text{NO}_2$ ] / M	[ $\text{CO}$ ] / M	Initial rate (Ms <sup>-1</sup> )
1	0.10	0.10	0.0021
2	0.20	0.10	0.0082
3	0.20	0.20	0.0083
4	0.40	0.10	0.033

From the data, determine the rate law and rate constant for the reaction.

$x=2, y=0$

## 2. Half-life Graph Method

The **reaction order** also can be determined by using the half-life equation !

How ?

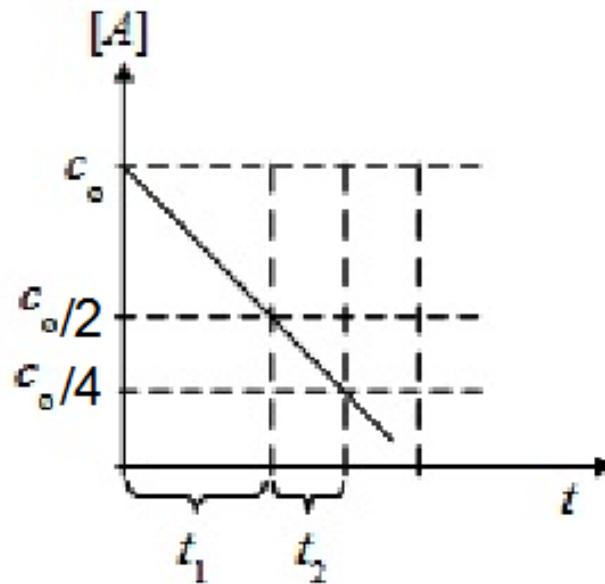
By **plotting graph** concentration against time !  
☞ look for the relationship between 1<sup>st</sup> half-life with 2<sup>nd</sup> half-life

Why ?

Let' discuss !

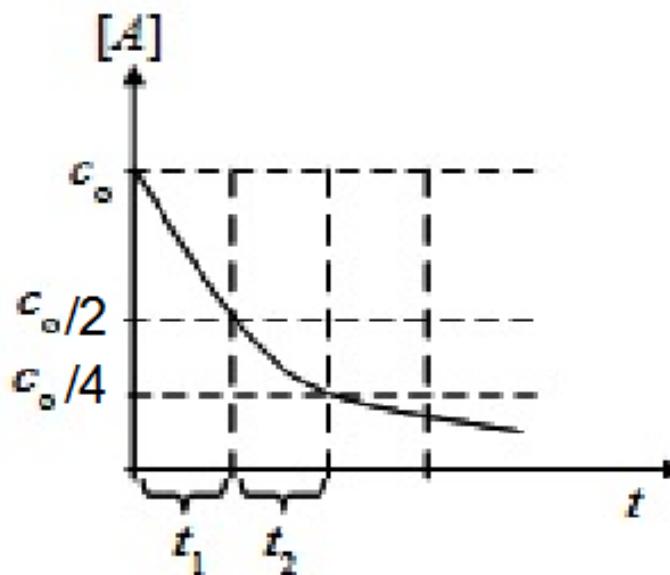
## Comparing the first and the second half-life

**zero-order**



$$t_1 = 2t_2$$

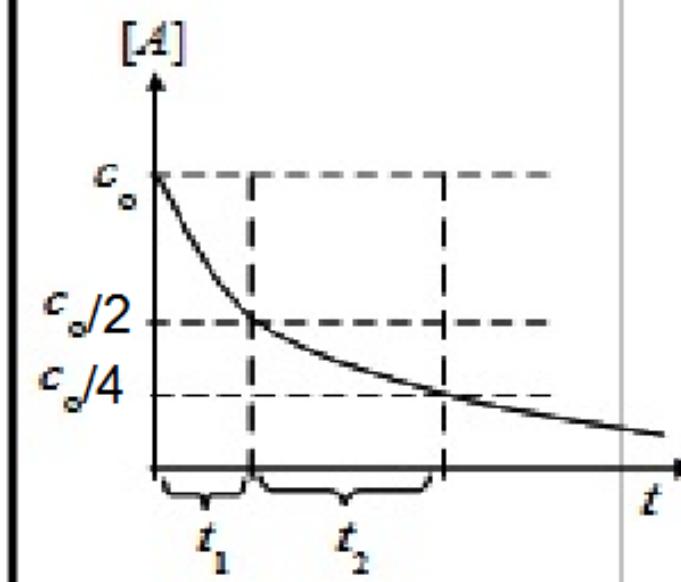
**1<sup>st</sup>-order**



$$t_1 = t_2$$

\*A straight line is obtained

**2<sup>nd</sup>-order**



$$2t_1 = t_2$$

\*A curve is obtained

\*A curve is obtained

**EXAMPLE**10

From the following data determine the order of the reaction, half-life and the rate constant.

TIME (min)	[ C <sub>2</sub> H <sub>5</sub> I ] (M)
0	1.00
25	0.78
50	0.61
75	0.47
100	0.37
150	0.22
200	0.14
250	0.08



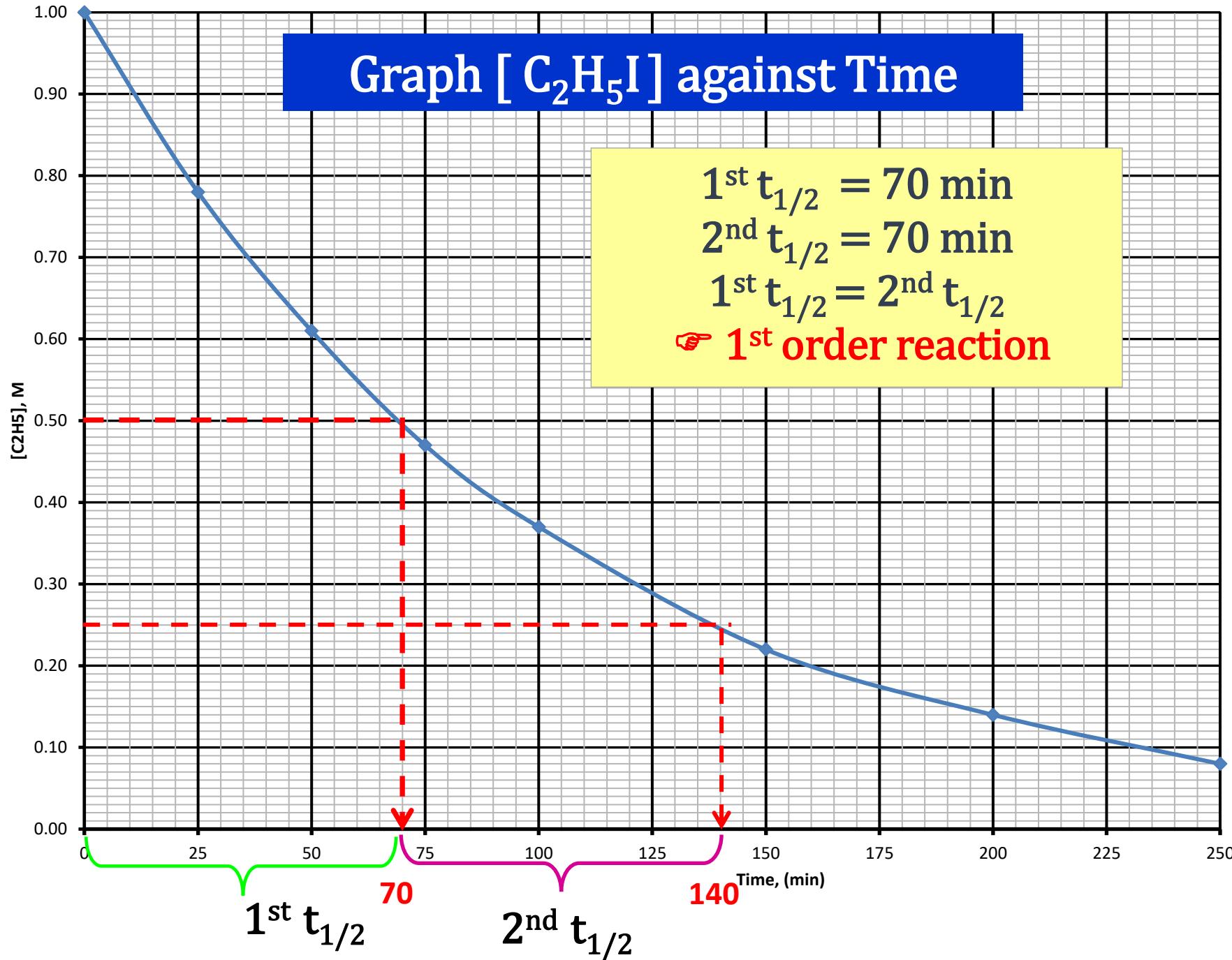
**EXAMPLE**



**10**



- Plot a graph  $[C_2H_5I]$  vs time
- Determine the 1<sup>st</sup> half-life and 2<sup>nd</sup> half-life from the graph
- Determine the order of the reaction
- Determine the rate constant,  $k$



**EXAMPLE****10**

**1<sup>st</sup> half-life = 2<sup>nd</sup> half-life,**

$$t_{1/2} \text{ (II)} = t_{1/2} \text{ (I)}$$

☞ **First order reaction**

**Half-life = 70 minute (from the graph)**

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

$$k = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{70 \text{ min}} = 9.9 \times 10^{-3} \text{ min}^{-1}$$

### 3. Linear Graph Method

The **reaction order** also can be determined by using the **integrated rate law** or **rate law**!

How ?

By **plotting the graph** !  
☞ look for the **linear graph**

Which graph ?  
Which one ???

I don't know. Just  
**try-and-error!**



Order

Rate Law

## Concentration-Time Equation (Integrated Rate Law)

0

rate =  $k$

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

1

rate =  $k[A]$

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

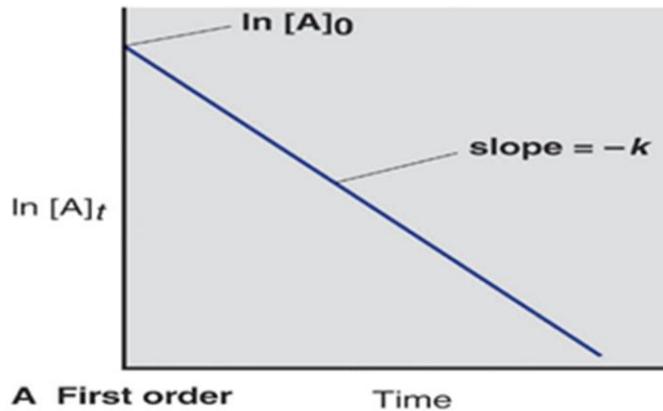
2

rate =  $k[A]^2$

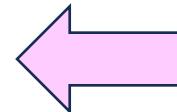
$$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$$

$$y = mx + c$$

$$y = mx + c$$

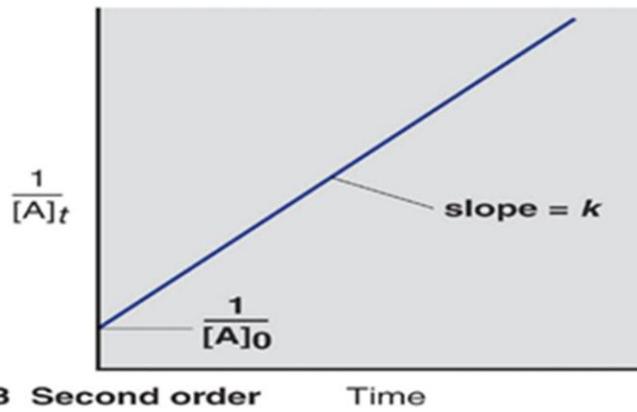


A First order

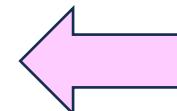


Plot  $\ln[\text{reactant}]$  vs. time:

- ☞ Linear graph obtained
- ☞ First order reaction

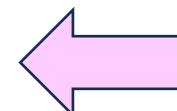
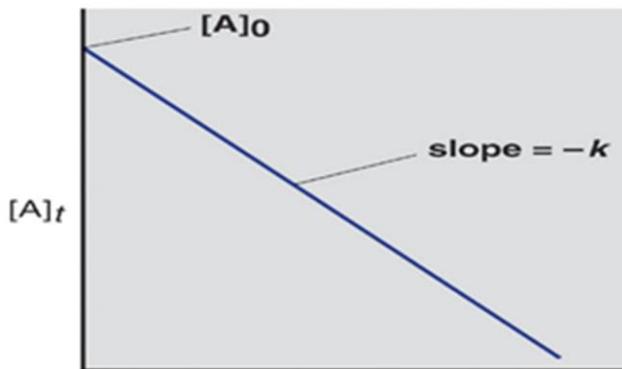


B Second order



Plot  $1/[\text{reactant}]$  vs. time:

- ☞ Linear graph obtained
- ☞ Second order reaction



Plot  $[\text{reactant}]$  vs. time:

- ☞ Linear graph obtained
- ☞ Zero order reaction

**EXAMPLE****11**

Ethyl iodide ( $\text{C}_2\text{H}_5\text{I}$ ) decomposes at a certain temperature in the gas phase as follows:



From the following data determine the order of the reaction and the rate constant.

TIME (min)	[ $\text{C}_2\text{H}_5\text{I}$ ] (M)
0	0.36
15	0.30
30	0.25
48	0.19
75	0.13

**EXAMPLE**

11



TIME (min)	[ C <sub>2</sub> H <sub>5</sub> I ] (M)	ln[C <sub>2</sub> H <sub>5</sub> I]	1/[C <sub>2</sub> H <sub>5</sub> I]
0	0.36	- 1.02	2.78
15	0.30	- 1.20	3.33
30	0.25	- 1.39	4.00
48	0.19	- 1.66	5.26
75	0.13	- 2.04	7.69

**EXAMPLE****11**

□ Plot  $[C_2H_5I]$  vs. time:

If a linear graph obtained ➔ **Zero Order Reaction**

□ Plot  $\ln[C_2H_5I]$  vs. time:

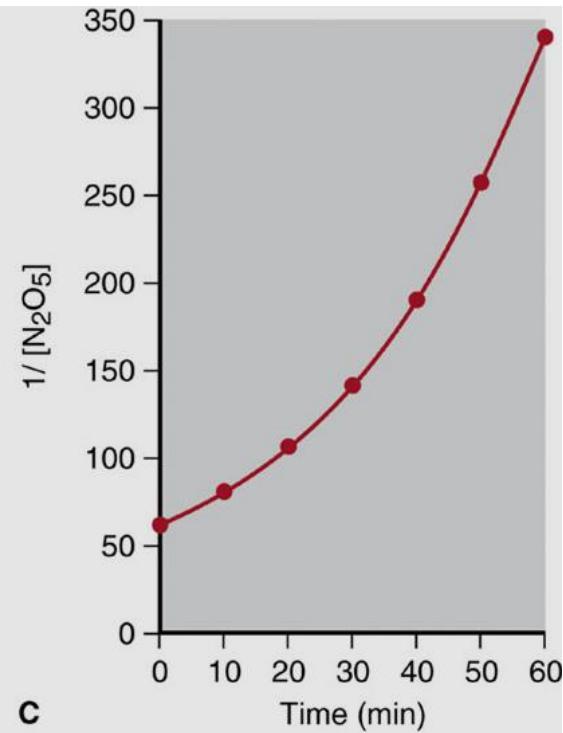
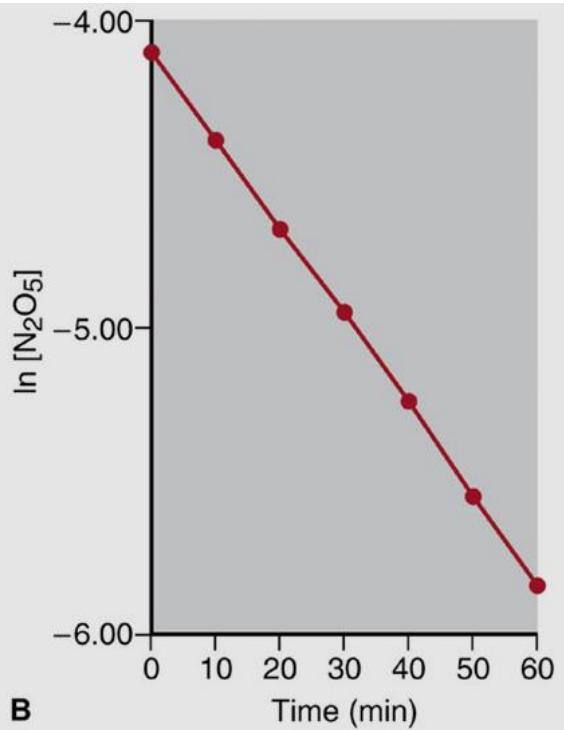
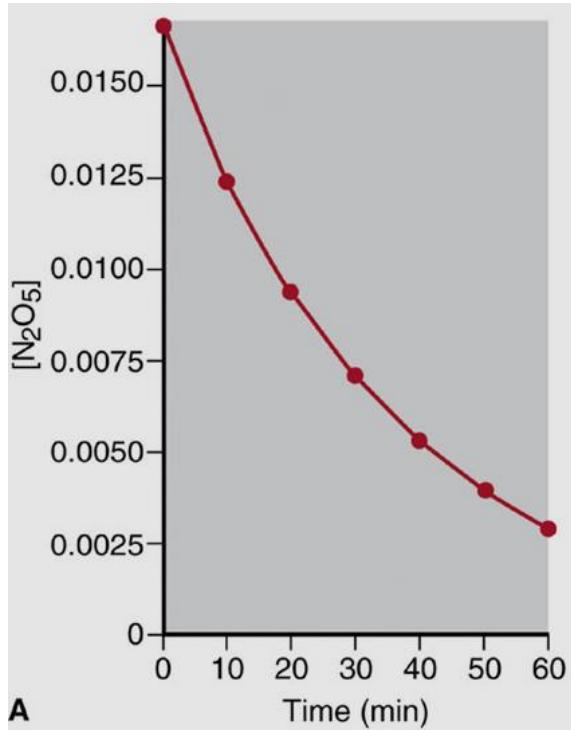
If a linear graph obtained ➔ **First Order Reaction**

□ Plot  $1/[C_2H_5I]$  vs. time:

If a linear graph obtained ➔ **Second Order Reaction**

## EXAMPLE

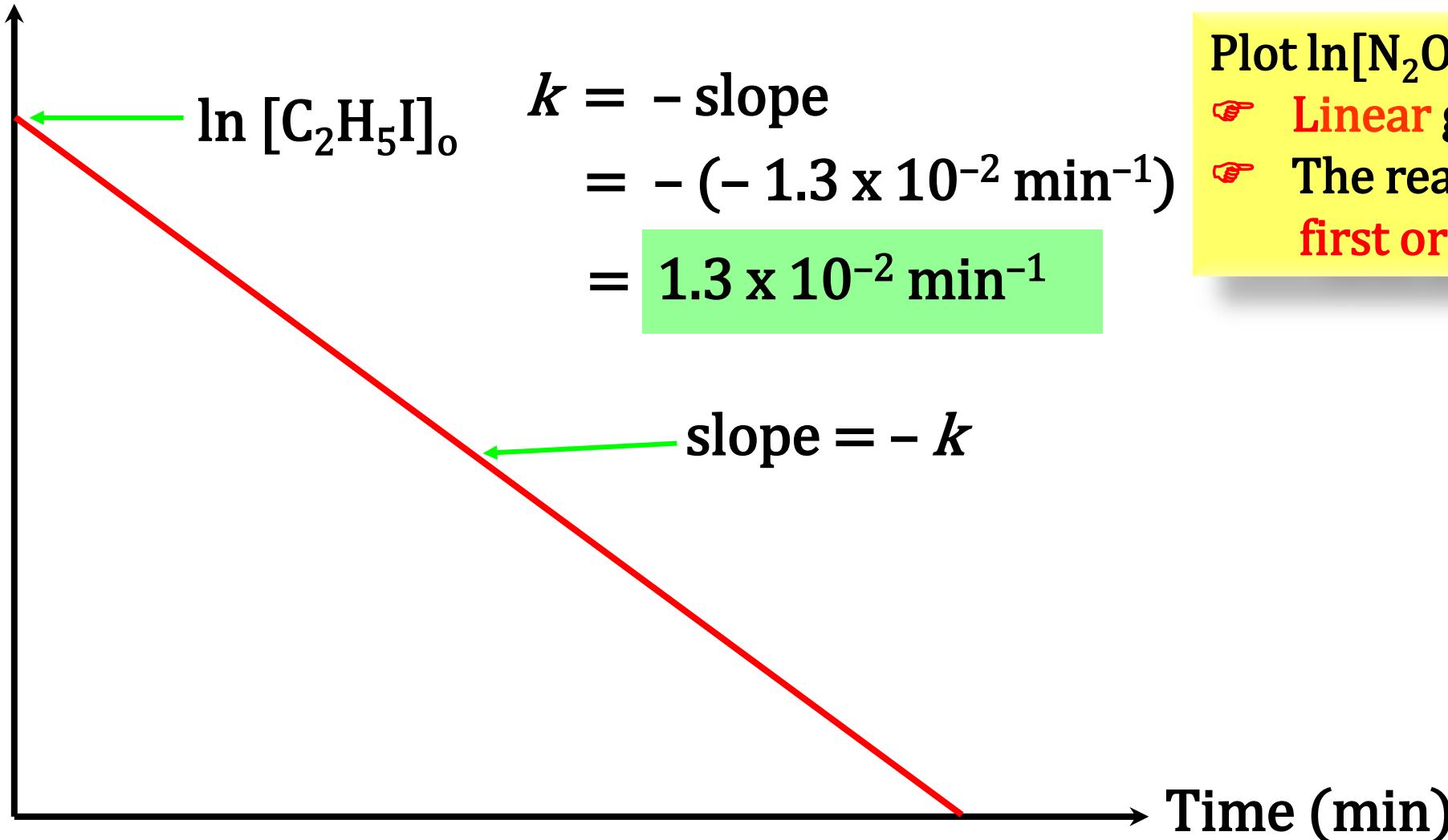
11



Time (min)	$[\text{N}_2\text{O}_5]$	$\ln [\text{N}_2\text{O}_5]$	$1/[\text{N}_2\text{O}_5]$
0	0.0165	-4.104	60.6
10	0.0124	-4.390	80.6
20	0.0093	-4.68	$1.1 \times 10^2$
30	0.0071	-4.95	$1.4 \times 10^2$
40	0.0053	-5.24	$1.9 \times 10^2$
50	0.0039	-5.55	$2.6 \times 10^2$
60	0.0029	-5.84	$3.4 \times 10^2$

**EXAMPLE**

11

**ANSWERS** $\ln [C_2H_5I]$ 

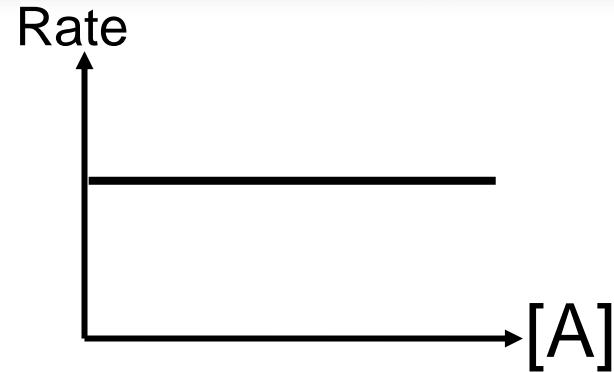
Plot  $\ln[N_2O_5]$  vs time:

- ☞ Linear graph obtained
- ☞ The reaction is first order in  $N_2O_5$

# Graph for Zero Order Reaction

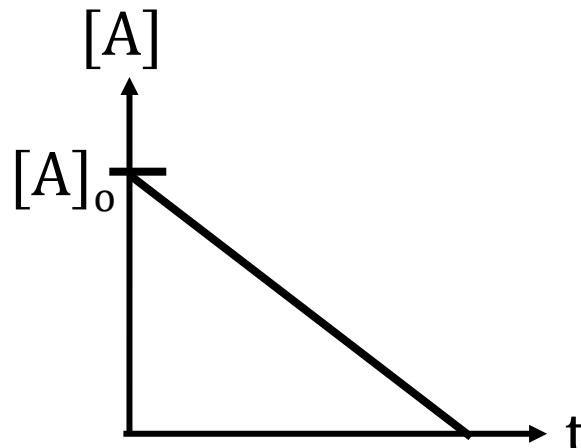
(a) From Rate Law:

$$\text{Rate} = k[A]^0$$
$$\text{Rate} = k$$

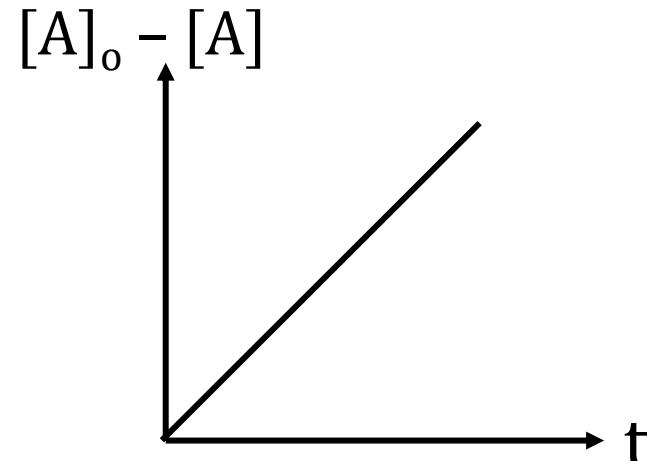


(b) From Integrated Rate Law:

(i)  $[A]_t = [A]_o - kt$



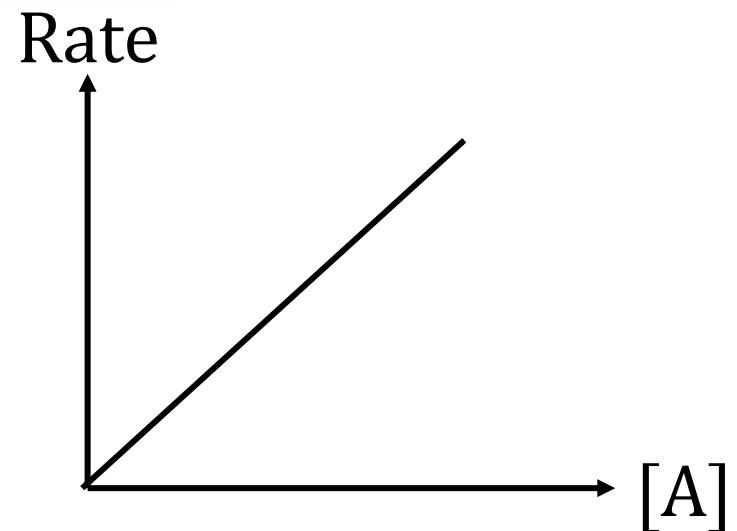
(ii)  $[A]_o - [A]_t = kt$



# Graph for First Order Reaction

(a) From Rate Law:

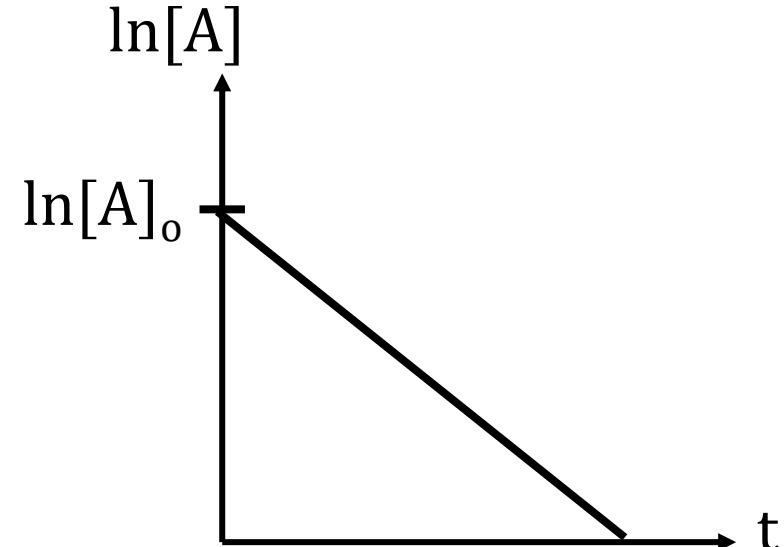
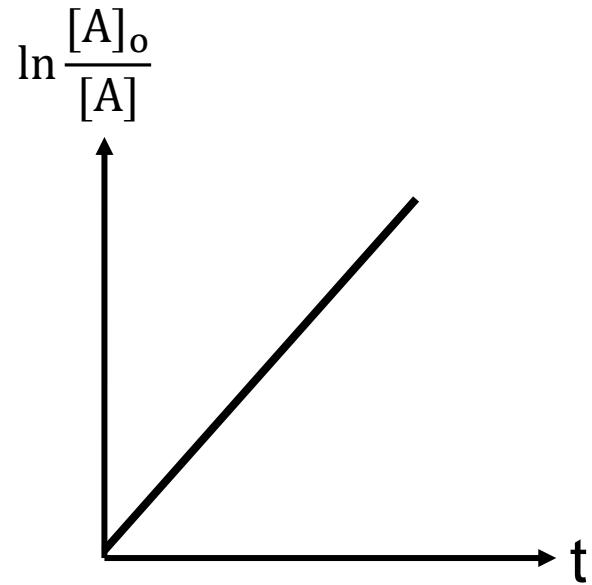
$$\text{Rate} = k[A]$$



## (b) From Integrated Rate Law:

$$(i) \ln \frac{[A]_o}{[A]_t} = kt$$

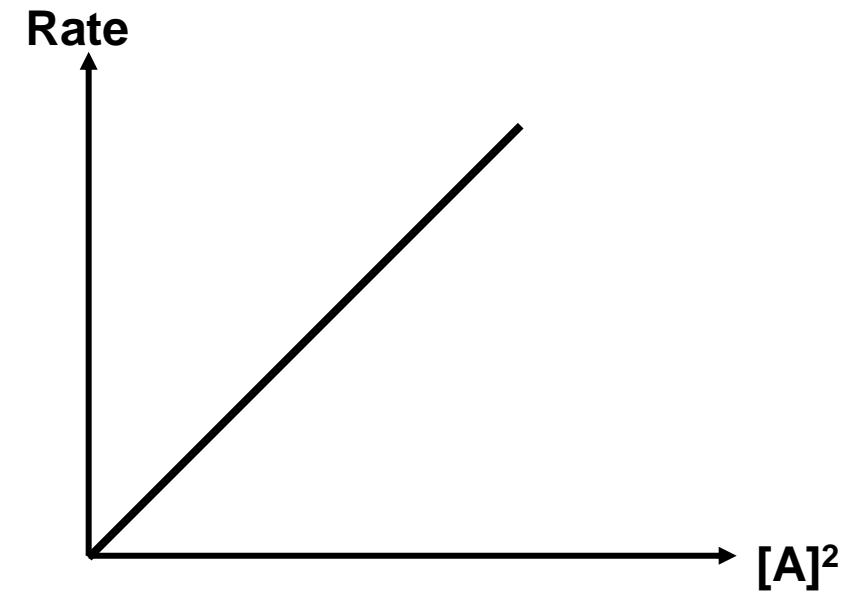
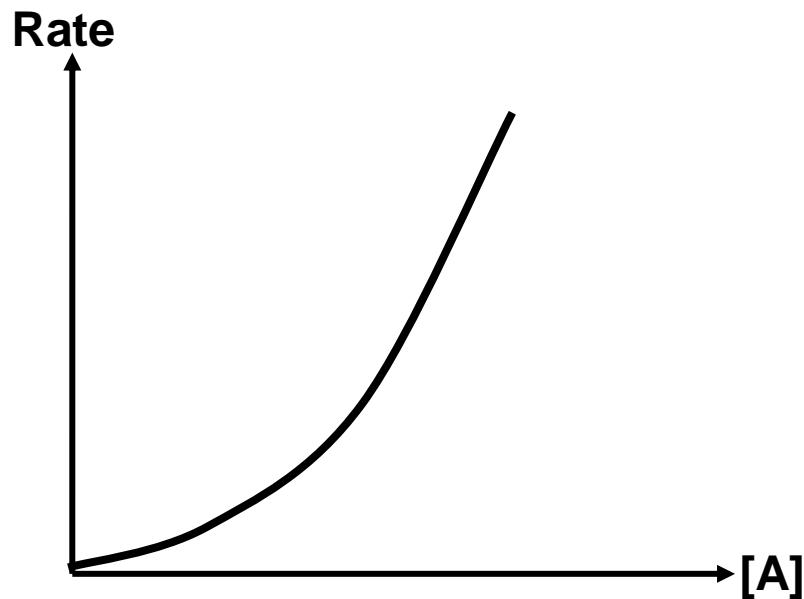
$$(ii) \ln[A]_t = \ln[A]_o - kt$$



# Graph for Second Order Reaction

(a) From Rate Law:

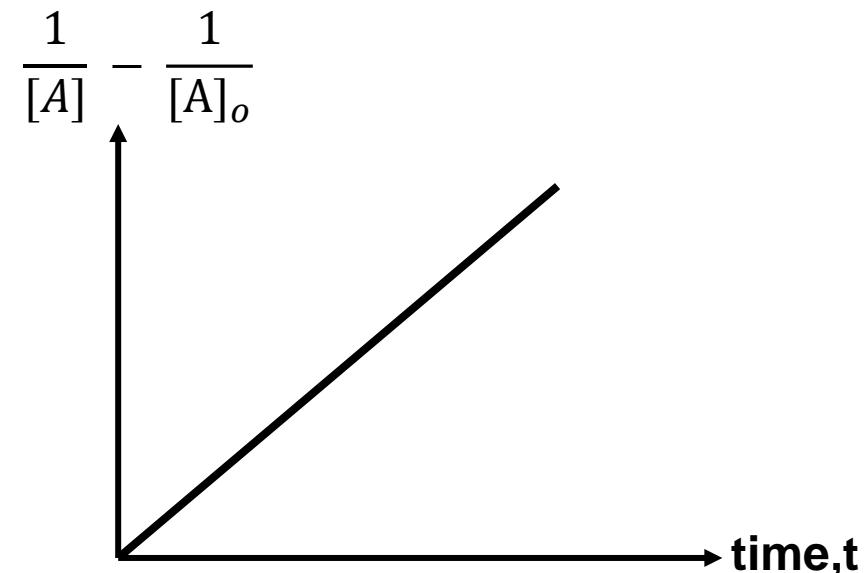
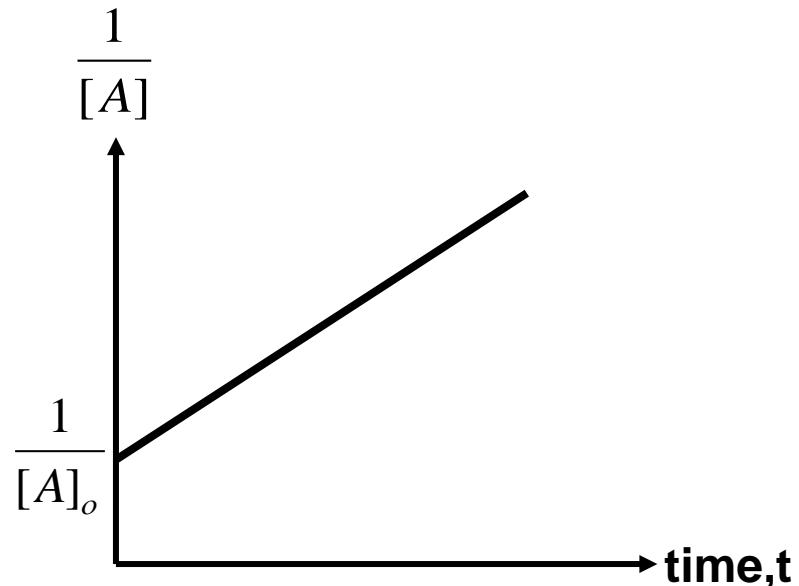
$$\text{Rate} = k[A]^2$$



## (b) From Integrated Rate Law:

$$(i) \frac{1}{[A]_t} = kt + \frac{1}{[A]_o}$$

$$(ii) \frac{1}{[A]_t} - \frac{1}{[A]_o} = kt$$





## Integrated Rate Law (Concentration-Time)

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

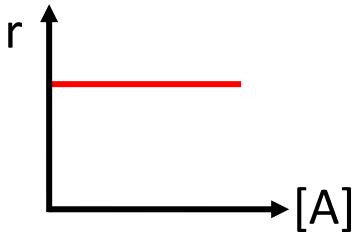


### Zero order

$A \rightarrow \text{product}$

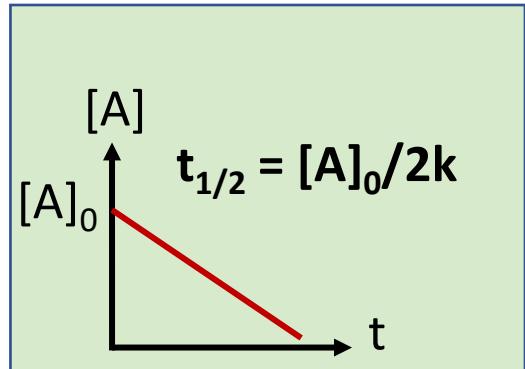
$$\text{Rate} = k[A]^0 \text{ or Rate} = k$$

$$\text{Unit } k = \text{M s}^{-1}$$



Integrated rate law

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

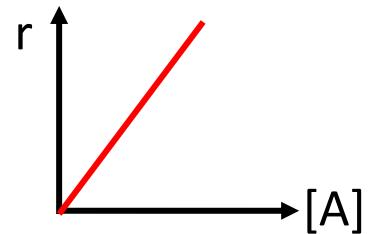


### 1<sup>st</sup> order

$A \rightarrow \text{product}$

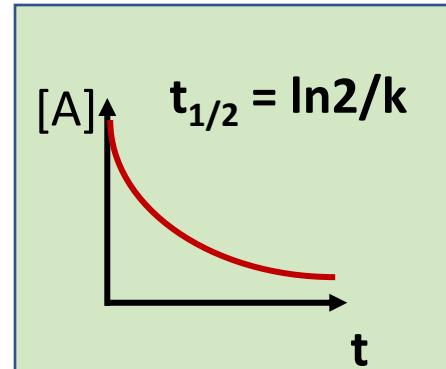
$$\text{Rate} = k[A]^1$$

$$\text{Unit } k = \text{s}^{-1}$$



Integrated rate law

$$\ln([A]_0 / [A]) = kt$$

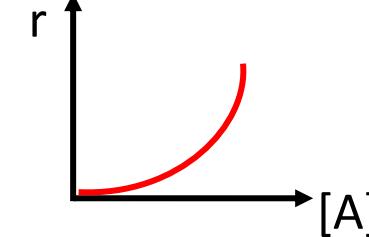


### 2<sup>nd</sup> order

$A \rightarrow \text{product}$

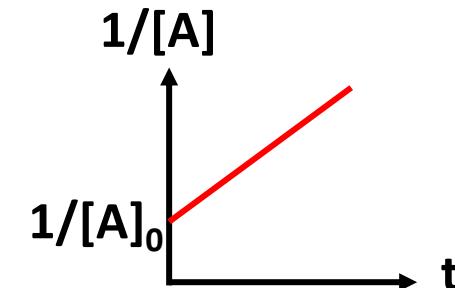
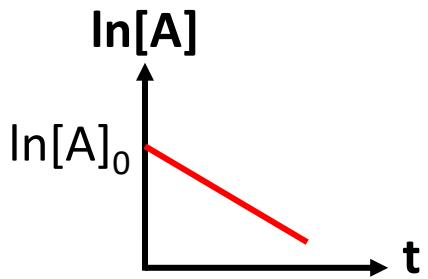
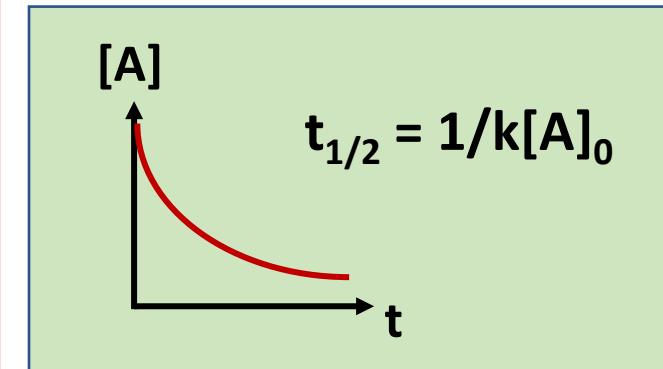
$$\text{Rate} = k[A]^2$$

$$\text{Unit } k = \text{M}^{-1} \text{s}^{-1}$$



Integrated rate law

$$1/[A] - 1/[A]_0 = kt$$



## **1.2 COLLISION THEORY**

# Collision Theory

- The collision theory is based on these ideas:
  - i) Molecules of **reactants must collide** in order to form products.
  - ii) Products are formed only when **effective collisions** occur between molecules.

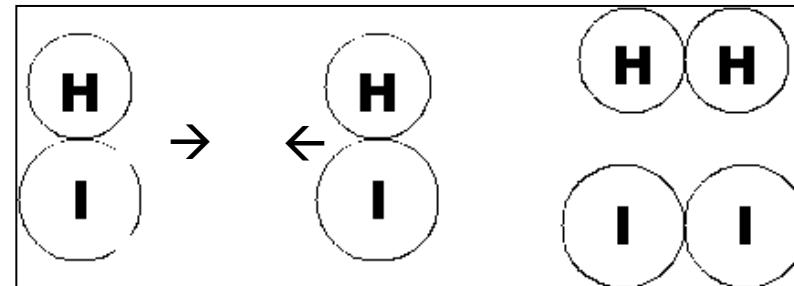
# Effective Collision

- Effective collisions require the followings:
  - i) The colliding molecules must have a total kinetic energy equal to or greater than the **activation energy,  $E_a$ .**
  - ii) **Collisions occur at the correct orientation.**

# Orientation of Collision

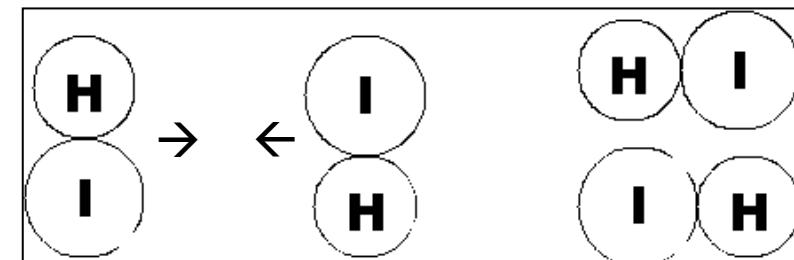
- The correct orientation means that the **collision of molecules occur at the correct angle that favors the formation of products.**

Example:



Effective Collision

☞ Reaction occurs



Ineffective Collision

☞ No reaction occurs

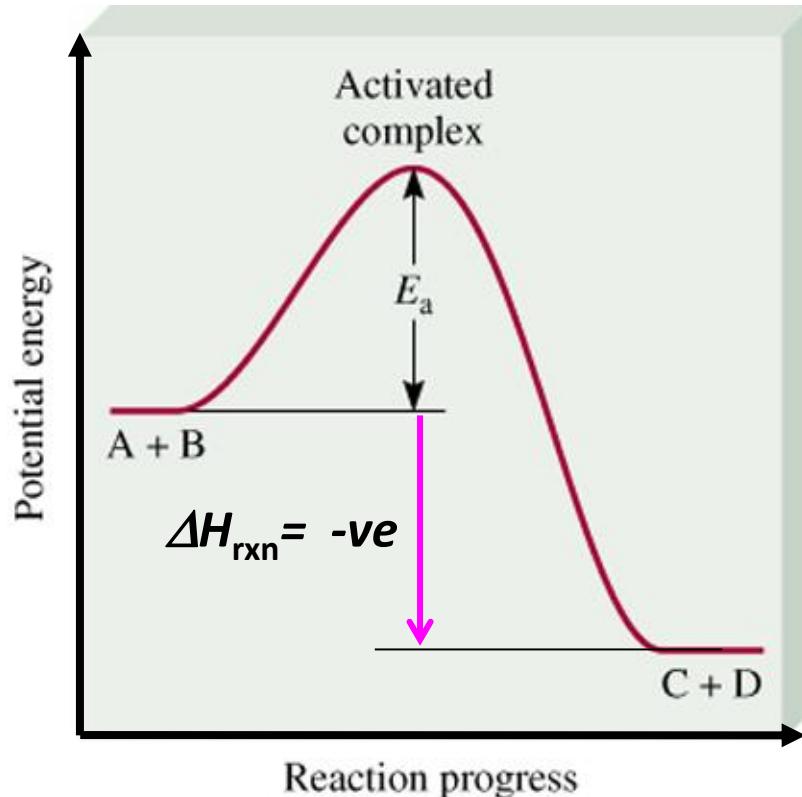
# Activation Energy, $E_a$

- The activation energy,  $E_a$  is the minimum energy required to initiate a chemical reaction.
- The activation energy,  $E_a$  may be gained through the collisions of molecules.
- The energy is used to break the bonds between atoms of the reactants
- Then these atoms will form new bonds in the product.

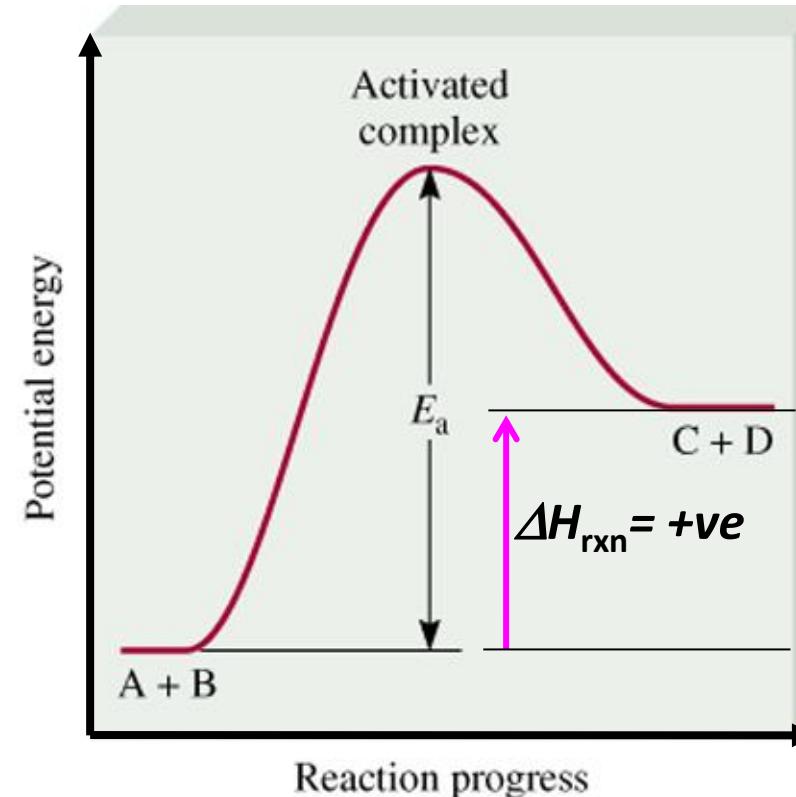
# Energy Profile Diagram



Exothermic Reaction



Endothermic Reaction



**EXAMPLE****12**

A key reaction in the upper atmosphere is

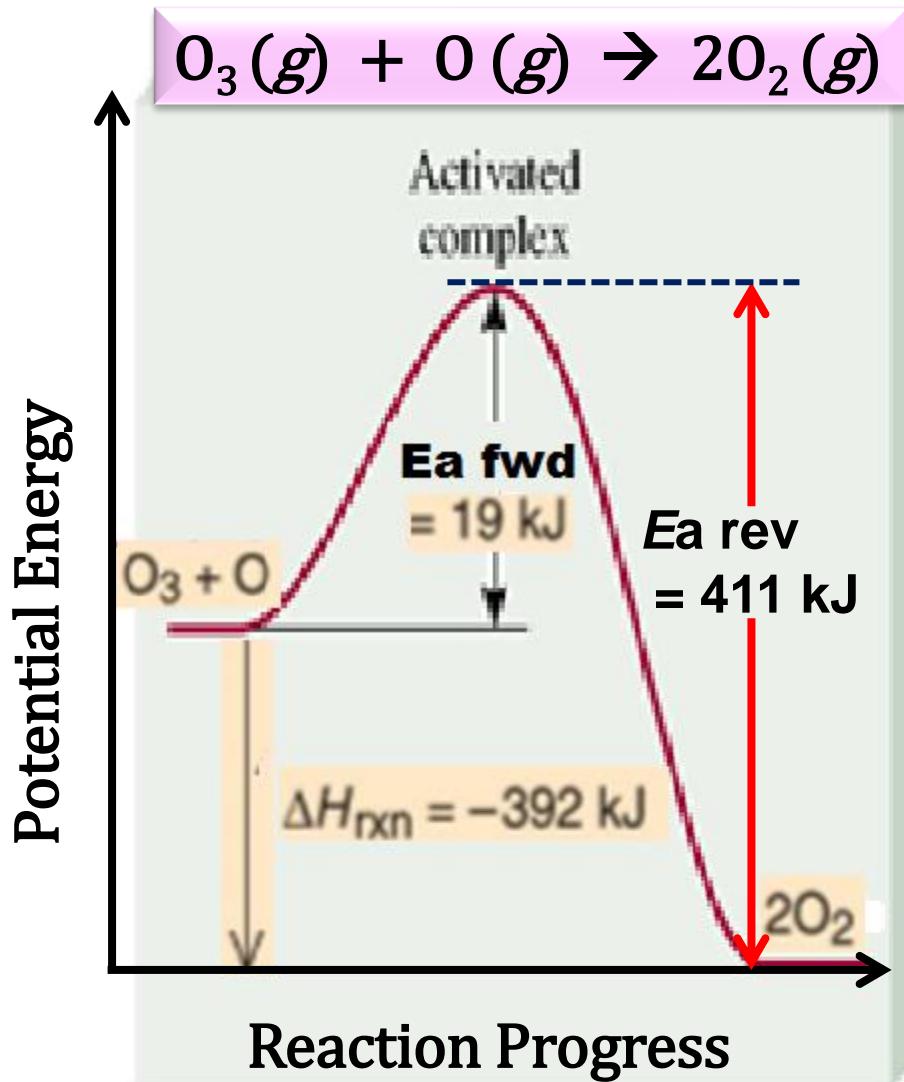


The  $E_a$ (forward) is 19 kJ, and  $\Delta H_{\text{rxn}}$  for the reaction is -392 kJ. Draw a reaction energy diagram and calculate  $E_a$ (reverse).

## EXAMPLE

12

ANSWERS



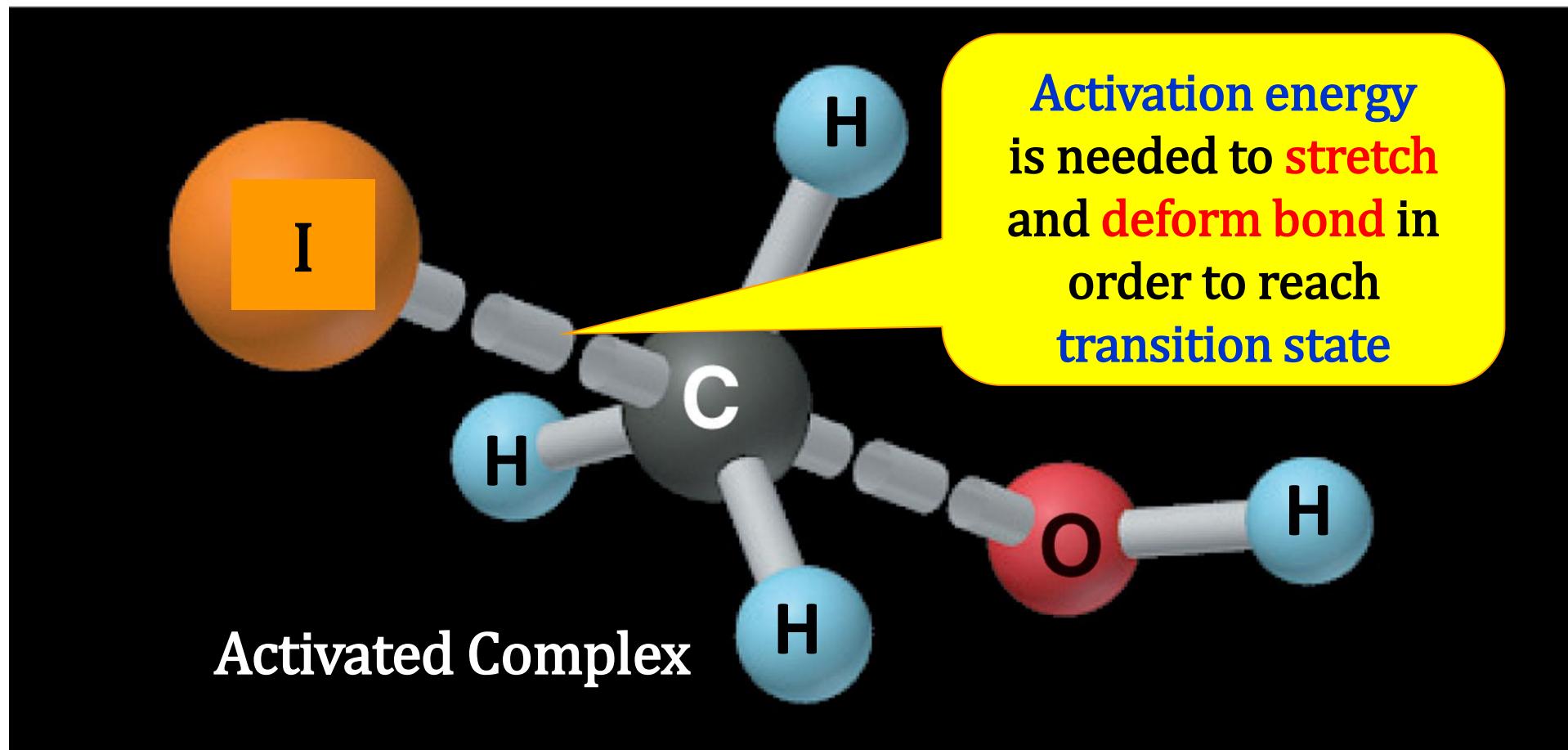
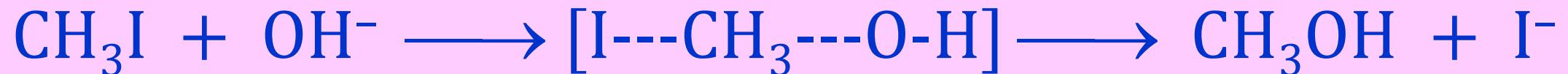
$$\begin{aligned}\Delta H_{rxn} &= Ea_{fwd} - Ea_{rev} \\ -392 \text{ kJ} &= 19 \text{ kJ} - Ea_{rev} \\ Ea_{rev} &= 19 \text{ kJ} + 392 \text{ kJ} \\ &= 411 \text{ kJ}\end{aligned}$$

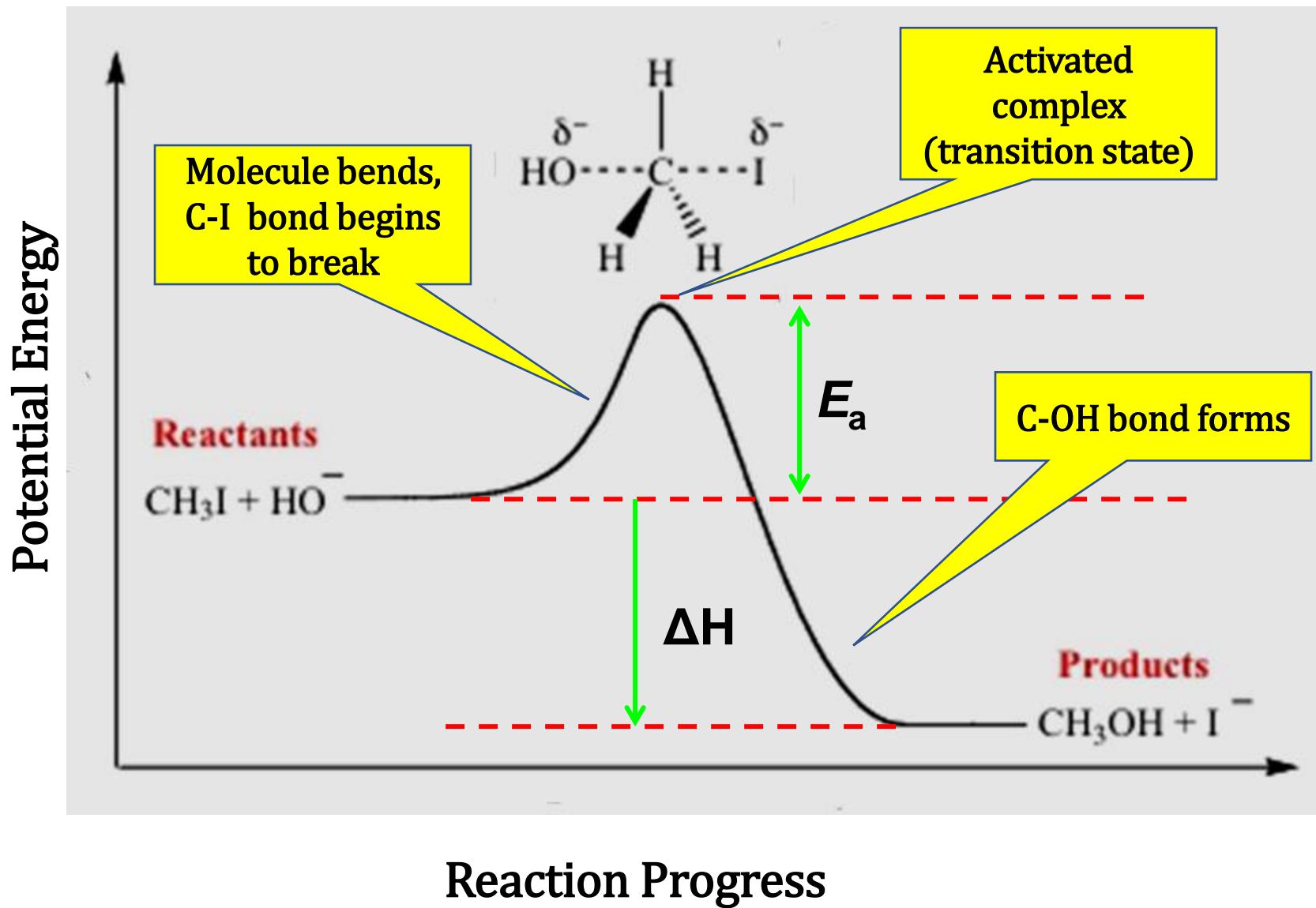
# Transition State Theory

- Explains the activation energy,  $E_a$  as the energy required by reactant molecules to form the **activated complex**.
- Activated complex is a temporary species formed by the reactant molecules as a result of the collision **before they form the products**.
- The transition theory is best explained by using the energy profile diagram.

EXAMPLE

13







## **1.3 FACTORS AFFECTING REACTION RATE**

# Factors Affecting Reaction Rate

## CONCENTRATIONS OF REACTANTS

- Reaction rates generally **increase** as the concentrations of the reactants are increased.



## TEMPERATURE

- Reaction rates generally **increase** rapidly as the **temperature** is increased.



## CATALYSTS

- Catalysts **speed up** reactions.



## PARTICLE SIZE

- The **reaction rate increases** as the **smaller the size** of reacting particles.



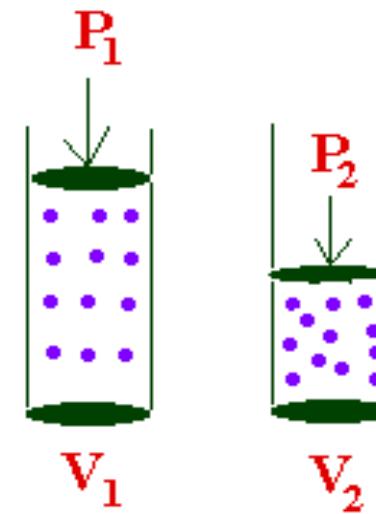
# 1. Concentration / Pressure

## □ Reactant concentration / Pressure of gases

$$PV = nR T$$

$$P = (n/V) R T \Rightarrow P = c R T$$

$\Rightarrow$  pressure,  $p \propto$  concentration,  $c$



Explanation:

*Pressure or Concentration increases,*

- particles are **closer** to each other
- More likely to collide (**higher collision frequency**)
- Probability of **effective collision increases**
- More molecules with kinetic energy **equal to or greater than  $E_a$**
- Reaction rate **increases**

# KEEP IN MIND

- Concentration change has  
No effect on **zero-order** reaction

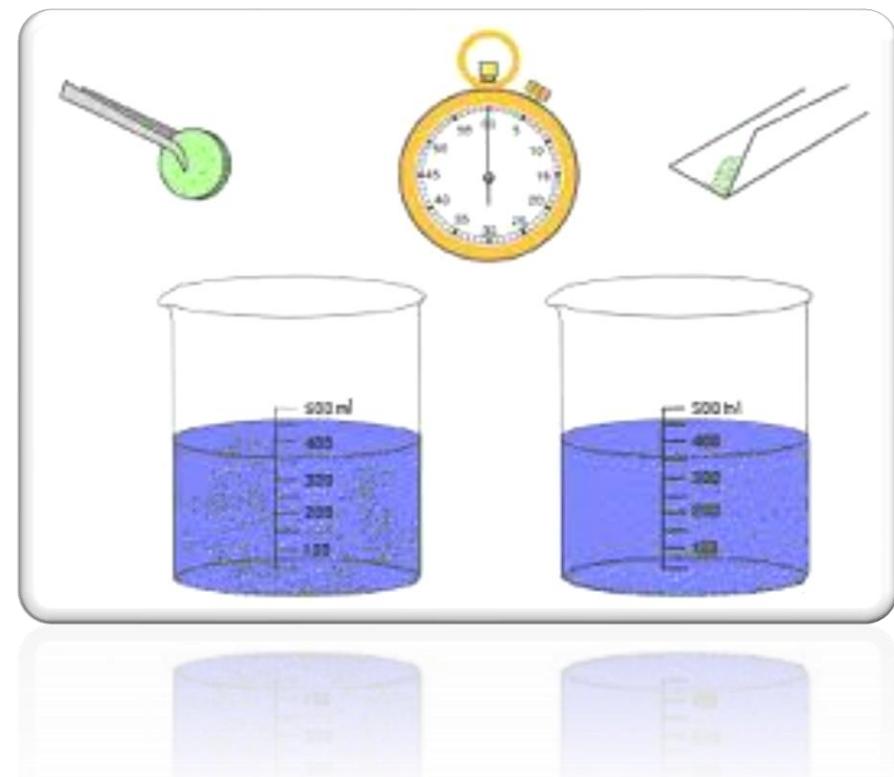
**Example:**  $R \rightarrow$  Product

Rate law:  $r = k [R]^{\circ}$

**Rate is independent of  $[R]$**

## 2. Particle Size

- The smaller the size of reacting particles, the larger the **contact surface area (solid reactant)**.
- Thus, frequency of collision increases.
- Probability of **effective collision** also increases.
- Thus, the **reaction rate increases**.



### 3. Temperature

- As **temperature** increases, **kinetic energy** of molecules increases.
- More **collisions** occur in a given time.
- More molecules will have energy greater than **activation energy,  $E_a$** .
- **Effective collisions** will increase.
- Thus, the **rate of reaction** increases.

## 4. Catalyst

- A **catalyst** is a substance that increases the rate of a chemical reaction without itself being consumed.
- A catalyst **provides an alternative pathway which has a lower activation energy** compared to the one without catalyst.
- Addition of a catalyst increases the reaction rate by **increasing the frequency of effective collision**. That is by
  - Decreasing the  $E_a$ , and
  - Correct orientation.

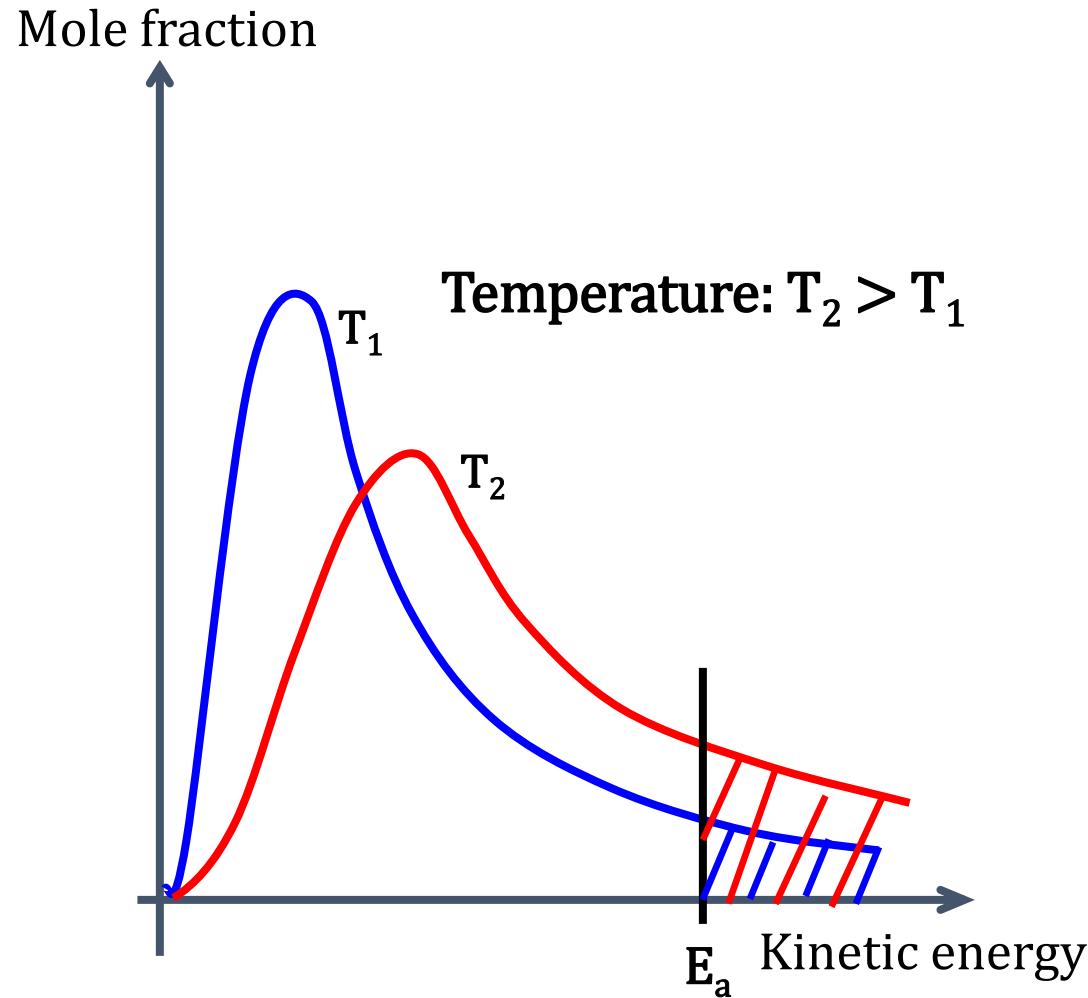
# Effect of Temperature on Reaction Rate

- The effect of temperature on the reaction rate can be explained in terms of kinetic theory.

Average Kinetic Energy  $\propto$  Temperature

- The higher the **temperature** the higher the **rate of reaction**.
- **Maxwell-Boltzmann distribution** shows the kinetic energy distributions for a reaction mixture at two different temperature.

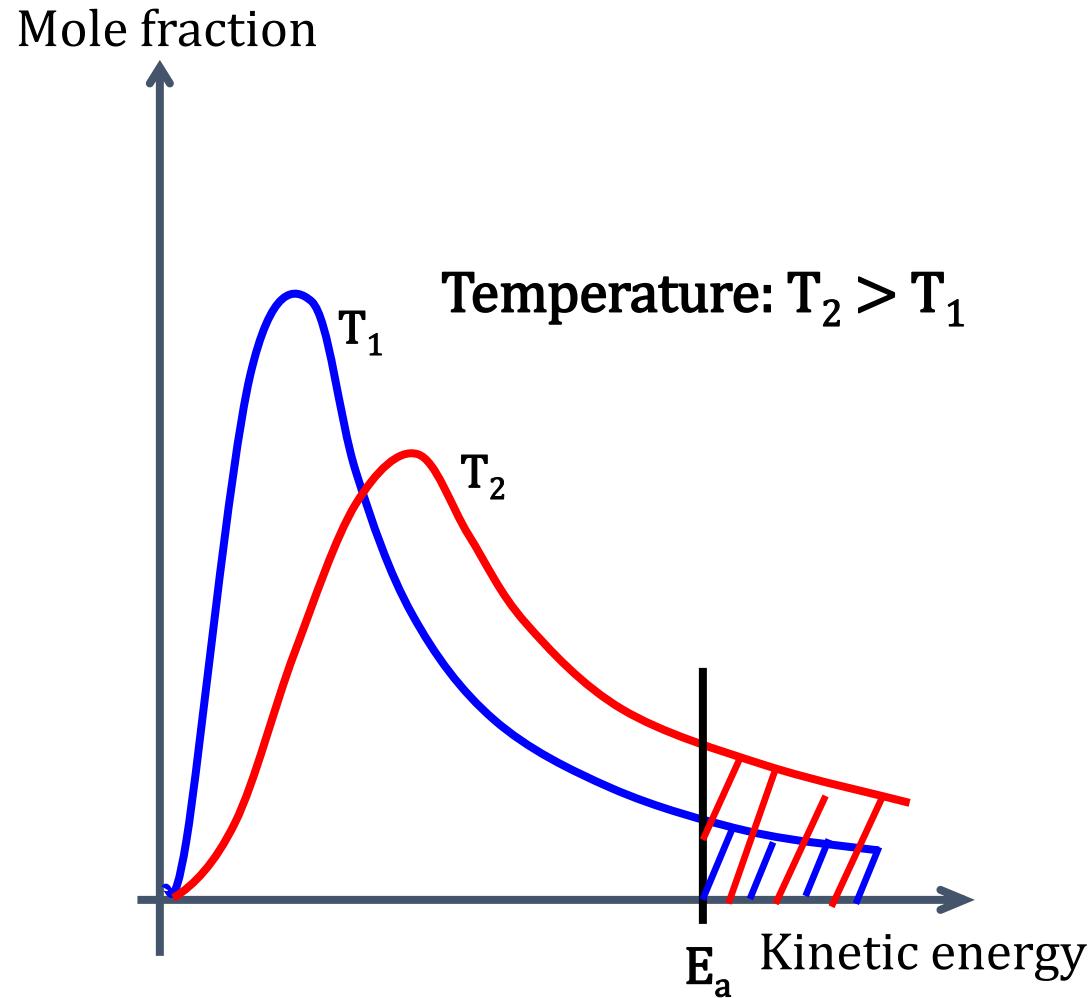
# Maxwell-Boltzmann Distribution Curve



## Features:

- $T_2 > T_1$  : At  $T_2$ , molecules that have kinetic energy equal or greater than activation energy is more compared to  $T_1$ .
- The area under the curve equal to the **total number of molecules**
- Number of molecules below **both** curve are the **same**
- The shaded area represent number of molecules possessing kinetic energy,  **$KE \geq$  activation energy,  $E_a$**

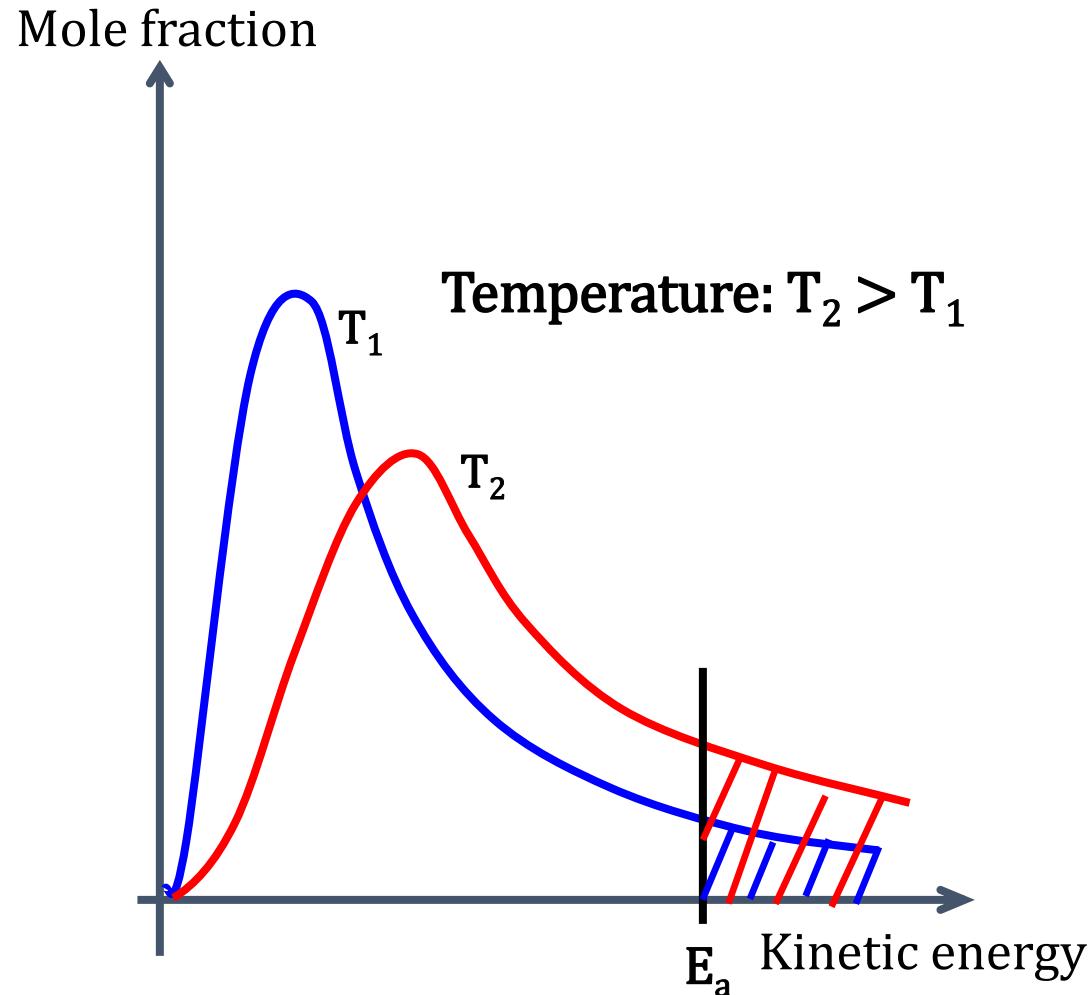
# Maxwell-Boltzmann Distribution Curve



## Features:

- At higher temperature ( $T_2$ ),
  - the peak of the curve **moves** to the right
  - the curves **flattens**
  - more molecules with **higher kinetic energy, KE** (larger pink shaded area)
- Very **few molecules** have low and high KE

# Maxwell-Boltzmann Distribution Curve



- ❑ At **higher temperature ( $T_2$ )**
  - ❑ Particles absorb energy and become **more energetic**
  - ❑ More molecules possess kinetic energy **equal to or higher** than  $E_a$
  - ❑ i.e. higher frequency of effective collision
- ❑ Reaction rate at  $T_2$  **increases**

$T \uparrow$ , average kinetic energy of the molecules  $\uparrow$ , number of molecules with energy equal or greater than  $E_a \uparrow$ , effective collision  $\uparrow$ , reaction rate  $\uparrow$

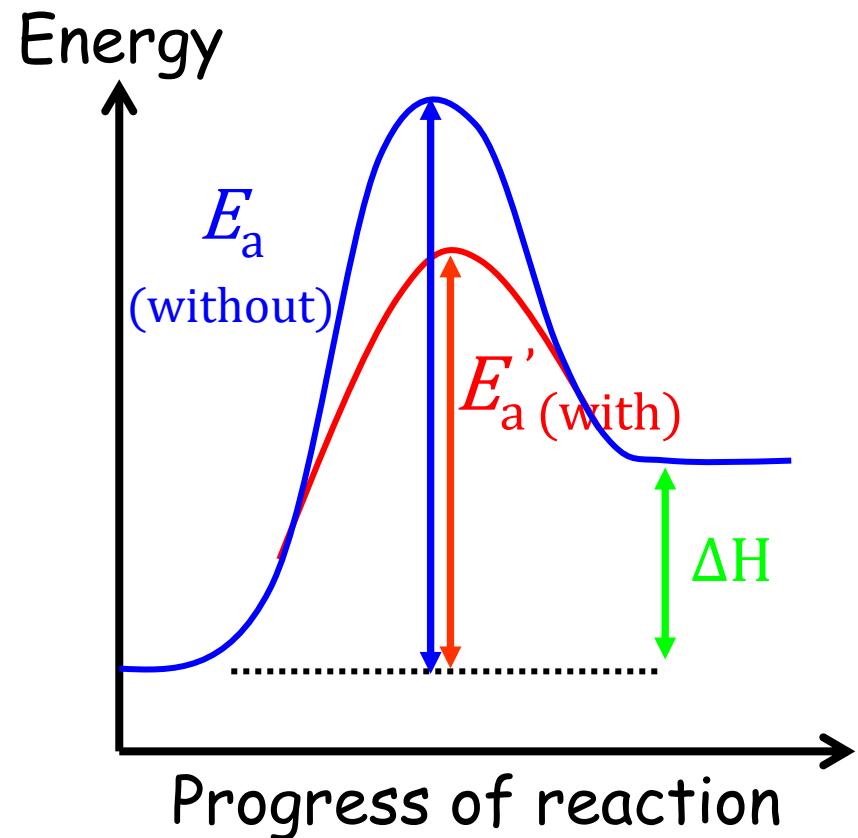
# Effect of Catalyst

- Energy profile diagram showing the difference reaction in the presence of catalyst and without catalyst.

In the presence of catalyst,  $E_a$  is lower;

- More molecules have kinetic energy equal to or higher than  $E_a$
- the probability of effective collisions also increases
- thus, reaction rate increases.

Endothermic reaction



# Characteristics of Catalyst

- They catalyse a **specific reaction**.
- Catalysts **lower by the same amount** the activation energies of the forward and backward reactions of a reversible reaction.
- A catalyst neither **alters the position** of equilibrium nor **increases the yield of products**.
- **Do not** change the value of  $\Delta H$  (enthalpy change) and  $K$  (equilibrium constant) but **change** the  $K$  (rate constant) – rate law @ Arrhenius equation.
- The catalyst may be **changed physically** but the **mass** of catalyst is **unchanged** at the end of the reaction.
- It won't be **denaturalised** at high temperature.

# Arrhenius Equation

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



$$\ln k = \ln A - \frac{E_a}{RT}$$

$k$  : rate constant

$R$  : universal gas constant

( $8.314 \text{ Jmol}^{-1} \text{ K}^{-1}$ )

$e$  : base of natural logarithm

$T$  : absolute temperature (in K)

$E_a$  : activation energy

$A$  : frequency factor

# Temperature and Rate

$$f = e^{-E_a/RT}$$

$f$ : Fraction of molecular collisions with **energy greater or equal to  $E_a$** .

- Temperature** affects the rate by affecting the **rate constant ( $k$ )**
- $T \uparrow$ , average **speed** of particles  $\uparrow$ , **frequency of collision**  $\uparrow$
- Only those collisions with enough **energy** to **exceed  $E_a$**  can lead to reaction.
- Temperature** rise enlarge fraction ( $f$ ) of collision with enough energy to exceed  $E_a$ .

## Arrhenius equation

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

$$f = e^{-E_a/RT}$$

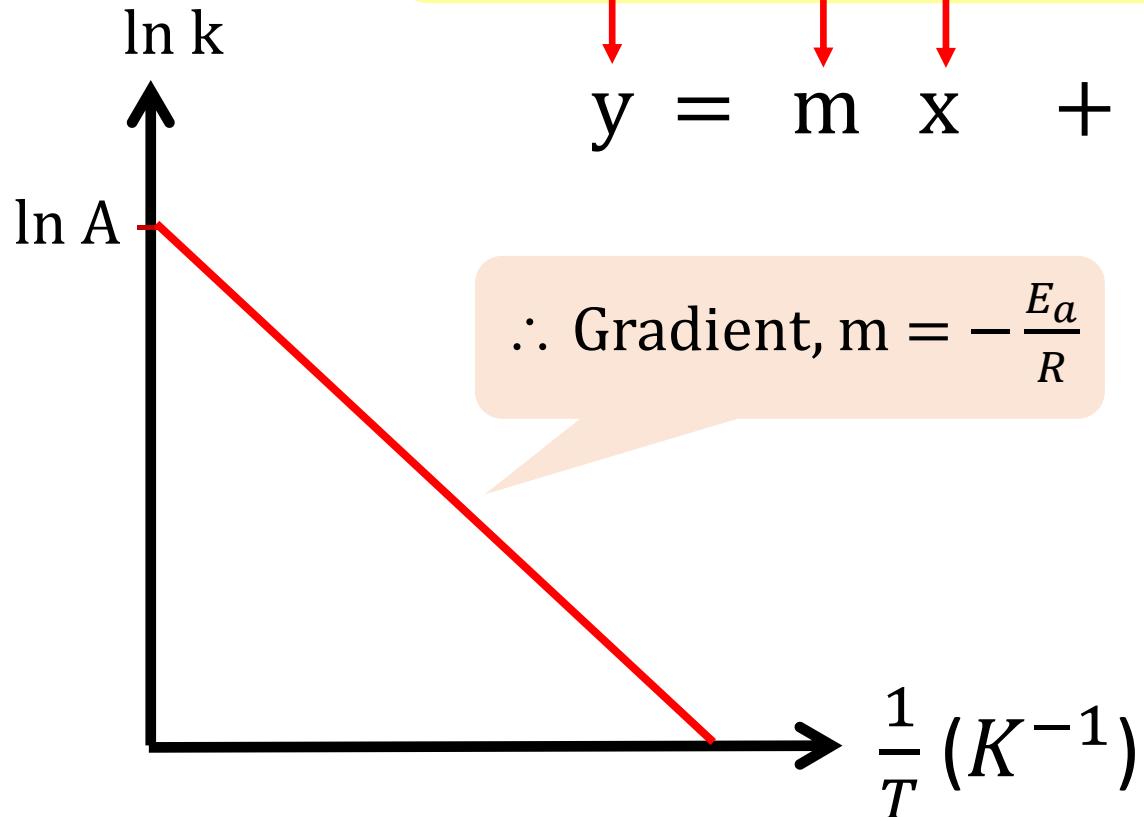
$f$ : Fraction of molecular collisions with energy greater or equal to  $E_a$ .

- Magnitudes of both  $E_a$  and  $T$  affect the fraction of sufficiently energetic collisions
  - Higher  $T$  ➔ larger  $k$  ➔ increased rate
  - Larger  $E_a$  ➔ smaller  $k$  ➔ decreased rate

# Graphically Method

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

$y = m x + c$



$R$  must be  $8.314 \text{ J mol}^{-1} \text{ K}^{-1}$   
not  $0.08206 \text{ atm L mol}^{-1} \text{ K}^{-1}$   
because the unit of  $E_a$   
is in  $\text{J}$  or  $\text{kJ}$  !

**EXAMPLE****14**

The rate constant for the decomposition of nitrous oxide ( $\text{N}_2\text{O}$ ) into  $\text{N}_2$  molecule and  $\text{O}$  atom has been measured at different temperatures:

$k(\text{M}^{-1} \text{s}^{-1})$	$T(\text{°C})$
0.00187	600
0.01130	650
0.05690	700
0.24400	750

Determine graphically the activation energy for the reaction.

**EXAMPLE**

14

ANSWERS

$k(M^{-1} s^{-1})$	$\ln k$	T (°C)	T (K)	$1/T (K^{-1})$
0.00187	- 6.281	600	873	$1.15 \times 10^{-3}$
0.01130	- 4.483	650	923	$1.08 \times 10^{-3}$
0.05690	- 2.866	700	973	$1.03 \times 10^{-3}$
0.24400	- 1.411	750	1023	$0.98 \times 10^{-3}$

 Plot graph:  $\ln k$  versus  $1/T$

**EXAMPLE**

14

 $\ln k$  $\ln A$ 

$$\begin{aligned}\text{Slope} &= -E_a/R \\ &= -2.89 \times 10^4 \text{ K}\end{aligned}$$

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

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

$$\begin{aligned}E_a &= (8.314 \text{ J/K mol})(2.89 \times 10^4 \text{ K}) \\ &= 2.40 \times 10^5 \text{ J/mol} \\ &= 240.2 \text{ kJ/mol}\end{aligned}$$

# Calculation Method

## KEEP IN MIND

Arrhenius equation

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

$$\ln K = \frac{E_a}{R} \frac{1}{T} + \ln A$$

$E_a$  must be in  $\text{J mol}^{-1}$

$R$  must be  $8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

$T$  must be in Kelvin

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

□ If we know the rate constant at TWO different temperature:

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

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

COMBINE

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

OR

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

**EXAMPLE****15**

The decomposition of  $\text{N}_2\text{O}_5$  has an activation energy of 103 kJ and a frequency factor of  $4.3 \times 10^{13} \text{ s}^{-1}$ . What is the rate constant for this decomposition at:

- (a) 20 °C
- (b) 100 °C

**EXAMPLE**

15

ANSWERS**Arrhenius equation**

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

(a)  $A = 4.3 \times 10^{13} \text{ s}^{-1}$

$$T = 20^\circ\text{C}$$

$$E_a = 103 \times 10^3 \text{ J mol}^{-1}$$

$$= (20 + 273.15)\text{K}$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$= 293.15 \text{ K}$$

$$k = 4.3 \times 10^{13} \text{ s}^{-1} \times e^{-\frac{103 \times 10^3 \text{ J mol}^{-1}}{8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 293.15 \text{ K}}}$$
$$= 1.9 \times 10^{-5} \text{ s}^{-1}$$

**Example 15**

- The decomposition of  $\text{N}_2\text{O}_5$  has an activation energy of 103 kJ and a frequency factor of  $4.3 \times 10^{13} \text{ s}^{-1}$ . What is the rate constant for this decomposition at
- $20^\circ\text{C}$
  - $100^\circ\text{C}$

**EXAMPLE**

15

**Arrhenius equation**

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

(b)  $A = 4.3 \times 10^{13} \text{ s}^{-1}$

$$T = 100 \text{ }^\circ\text{C}$$

$$E_a = 103 \times 10^3 \text{ J mol}^{-1}$$

$$= (100 + 273.15) \text{ K}$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$= 373 \text{ K}$$

$$k = 4.3 \times 10^{13} \text{ s}^{-1} \times e^{-\frac{103 \times 10^3 \text{ J mol}^{-1}}{8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 373.15 \text{ K}}}$$

$$= 0.16 \text{ s}^{-1}$$

The decomposition of  $\text{N}_2\text{O}_5$  has an activation energy of 103 kJ and a frequency factor of  $4.3 \times 10^{13} \text{ s}^{-1}$ . What is the rate constant for this decomposition at

- (a)  $20 \text{ }^\circ\text{C}$
- (b)  $100 \text{ }^\circ\text{C}$



**EXAMPLE**



**16**

**The decomposition of HI has rate constants**

$$k = 0.079 \text{ L mol}^{-1} \text{ s}^{-1} \text{ at } 508^\circ\text{C}$$

$$k = 0.24 \text{ L mol}^{-1} \text{ s}^{-1} \text{ at } 540^\circ\text{C}.$$

**What is the activation energy of this reaction in  $\text{kJ mol}^{-1}$ ?**

**EXAMPLE****16**ANSWERS

➤ Given,

$k_1 = 0.079 \text{ L mol}^{-1} \text{ s}^{-1}$	$T_1 = 508^\circ\text{C} = 781.15 \text{ K}$
$k_2 = 0.24 \text{ L mol}^{-1} \text{ s}^{-1}$	$T_2 = 540^\circ\text{C} = 813.15 \text{ K}$

➤  $E_a = ?$

➤ Use

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

**Example 16**

The decomposition of HI has rate constants

$k = 0.079 \text{ L mol}^{-1} \text{ s}^{-1}$  at  $508^\circ\text{C}$

$k = 0.24 \text{ L mol}^{-1} \text{ s}^{-1}$  at  $540^\circ\text{C}$ .

What is the activation energy of this reaction in  $\text{kJ mol}^{-1}$ ?

**EXAMPLE**16

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

$$\ln \frac{0.24 \text{ L mol}^{-1}\text{s}^{-1}}{0.079 \text{ L mol}^{-1}\text{s}^{-1}} = - \frac{E_a}{8.314 \text{ J mol}^{-1}\text{K}^{-1}} \left[ \frac{1}{813.15 \text{ K}} - \frac{1}{781.15 \text{ K}} \right]$$

$$E_a = 183311 \text{ J mol}^{-1}$$

$$= 183.3 \text{ kJ mol}^{-1}$$

The decomposition of HI has rate constants

$k = 0.079 \text{ L mol}^{-1} \text{s}^{-1}$  at  $508^\circ\text{C}$

$k = 0.24 \text{ L mol}^{-1} \text{s}^{-1}$  at  $540^\circ\text{C}$ .

What is the activation energy of this reaction in  $\text{kJ mol}^{-1}$ ?

# EXERCISE

# 8



The rate constant of a first-order reaction is  $4.6 \times 10^{-4} \text{ s}^{-1}$  at  $350^\circ\text{C}$ . If the activation energy is  $104 \text{ kJ mol}^{-1}$ , calculate the temperature at which its rate constant is  $8.80 \times 10^{-4} \text{ s}^{-1}$ .

370.6°C

## GLOSSARY

BIL	TERM	SYMBOL	DEFINE
1.	REACTION RATE		Changes in concentrations of a reactant (or product) with time
2.	DIFFERENTIAL RATE EQUATION		Shows the relationship between the rate of disappearance of reactants and formation of products.
3.	RATE LAW/ RATE EQUATION		An equation that relates the rate of a reaction to the rate constant and the concentrations of the reactants raised to some powers.
4.	ORDER OF REACTION		The power to which the concentration of a reactant is raised in a rate law
5.	HALF – LIFE	$t_{1/2}$	Time required for the concentration of a reactant to decrease to half of its initial concentration/value

BIL	TERM	SYMBOL	DEFINE
6.	INTEGRATED RATE LAW		The relationship between the concentration of the reactants and time.
7.	RATE CONSTANT	K	Constant of proportionality between the reaction rate and the concentration s of reactants
8.	ACTIVATION ENERGY	$E_A$	The minimum energy required to initiate a chemical reaction.
9.	EXOTHERMIC PROCESS		A process in which a system release heat to its surroundings
10	ENDOTHERMIC PROCESS		A process in which a system absorbs heat to its surroundings
11	ACTIVATED COMPLEX		The species temporarily formed by the reactant molecules as a result of the collision before they form the product