

CHEMICAL KINETICS

Chemical Kinetics

- The area of chemistry that concerns **reaction rates**.

Reaction Rates

- The questions posed in this chapter will be:
 - How is the rate of a reaction measured?
 - What conditions will affect the rate of a reaction?
 - How do you express the relationship of rate to the variables affecting the rate?
 - What happens on a molecular level during a chemical reaction?

Reaction Rate

- **Change in concentration** (conc) of a reactant or product **per unit time**.

$$\text{Rate} = \frac{\text{conc of A at time } t_2 - \text{conc of A at time } t_1}{t_2 - t_1}$$
$$= \frac{\Delta[\text{A}]}{\Delta t}$$

Definition of Reaction Rate

- The **reaction rate** is the **increase in molar concentration of a product of a reaction per unit time**.
 - It can also be expressed as the **decrease** in molar concentration of a **reactant** per unit time.

1. **Rate of formation of product =**
 $\Delta[\text{product}]/\Delta t$
This is the average rate over time
2. **Rate of reaction of reactant =**
 $-\Delta[\text{Reactant}]/\Delta t$
This is the average rate over time

Definition of Reaction Rates

- Consider the gas-phase decomposition of dinitrogen pentoxide.



- If we denote molar concentrations using brackets, then the change in the molarity of O_2 would be represented as

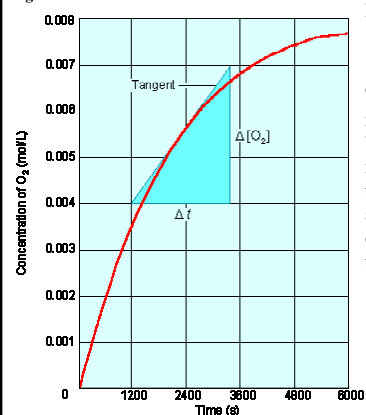
$$\Delta[\text{O}_2]$$

where the symbol, Δ (capital Greek delta), means “the change in.”

Definition of Reaction Rates

- Figure 14.5 shows the increase in concentration of O_2 during the decomposition of N_2O_5 .
- Note that the **rate decreases** as the reaction proceeds.

Figure 14.4 The instantaneous rate of reaction



In the reaction



The concentration of O_2 increases over time. You obtain the instantaneous rate from the slope of the tangent at the point of the curve corresponding to that time.

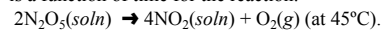
TABLE 12.3

Concentration/Time Data for the Reaction $2\text{N}_2\text{O}_5(\text{soln}) \rightarrow 4\text{NO}_2(\text{soln}) + \text{O}_2(\text{g})$ (at 45°C)

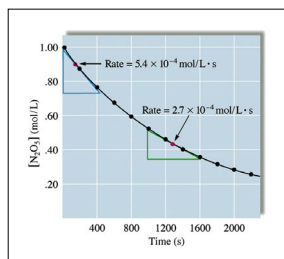
$[\text{N}_2\text{O}_5]$ (mol/L)	Time (s)
1.00	0
0.88	200
0.78	400
0.69	600
0.61	800
0.54	1000
0.48	1200
0.43	1400
0.38	1600
0.34	1800
0.30	2000

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Figure 12.3: A plot of the concentration of N_2O_5 is a function of time for the reaction:



Note that the reaction rate at $[\text{N}_2\text{O}_5] = 0.90 \text{ M}$ is twice that at $[\text{N}_2\text{O}_5] = 0.45 \text{ M}$.



Definition of Reaction Rates

- Because the amounts of products and reactants are related by stoichiometry, any substance in the reaction can be used to express the rate.

$$\text{Rate of decomposition of } \text{N}_2\text{O}_5 = -\frac{\Delta[\text{N}_2\text{O}_5]}{\Delta t}$$

- Note the negative sign. This results in a positive rate as reactant concentrations decrease.

Definition of Reaction Rates

- The rate of decomposition of N_2O_5 and the formation of O_2 are easily related.

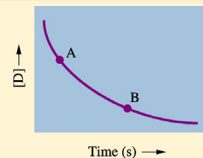
$$\frac{\Delta[\text{O}_2]}{\Delta t} = \frac{1}{2} \left(-\frac{\Delta[\text{N}_2\text{O}_5]}{\Delta t} \right)$$

- Since two moles of N_2O_5 decompose for each mole of O_2 formed, the rate of the decomposition of N_2O_5 is twice the rate of the formation of O_2 .

CONCEPT CHECK 14.1

Shown here is a plot of the concentration of a reactant D versus time.

- How do the instantaneous rates at points A and B compare?
- Is the rate for this reaction constant at all points in time?



Factors Affecting Reaction Rates

- Concentration of reactants.**
 - More often than not, the rate of a reaction **increases** when the concentration of a reactant is increased.
 - Increasing the population of reactants increases the likelihood of a successful collision.
 - In some reactions, however, the rate is unaffected by the concentration of a particular reactant, as long as it is present at some concentration.

Figure 14.7: The precipitate forms more slowly in a solution of lower concentration.

Photo courtesy of American Color.

Time = 0 seconds

Lower Concentration



Figure 14.7: The solution gains the bright yellow precipitate.

Photo courtesy of American Color.

Time = 20 seconds

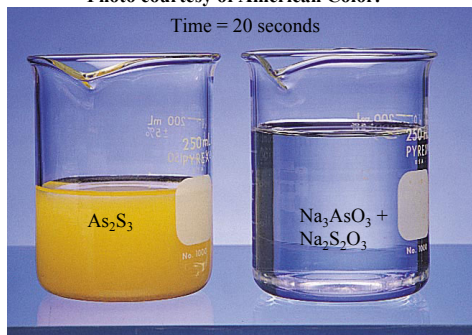
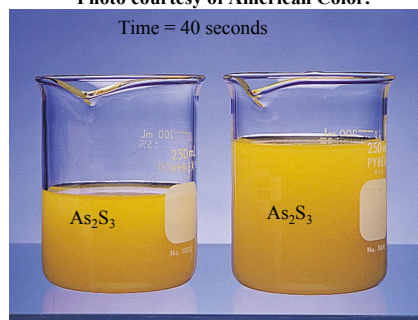


Figure 14.7: The beaker on the right contains more water.

Photo courtesy of American Color.

Time = 40 seconds

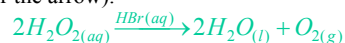


Factors Affecting Reaction Rates

- **Temperature at which a reaction occurs.**
 - Usually reactions speed up when the temperature increases.
 - A good “rule of thumb” is that reactions approximately double in rate with a 10 °C rise in temperature.

Factors Affecting Reaction Rates

- **Concentration of a catalyst.**
 - A **catalyst** is a **substance that increases the rate of a reaction without being consumed in the overall reaction.**
 - The catalyst generally does not appear in the overall balanced chemical equation (although its presence may be indicated by writing its formula over the arrow).

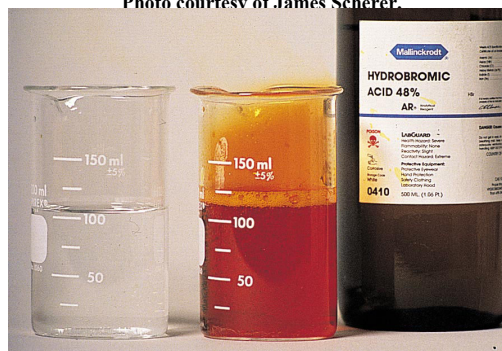


Factors Affecting Reaction Rates

- **Concentration of a catalyst.**
 - Figure 14.2 shows the HBr catalyzed decomposition of H_2O_2 to H_2O and O_2 .
- $$2\text{H}_2\text{O}_{2(aq)} \xrightarrow{\text{HBr}(aq)} 2\text{H}_2\text{O}_{(l)} + \text{O}_{2(g)}$$
- A catalyst speeds up reactions by reducing the “activation energy” needed for a successful reaction.
 - A catalyst may also provide an alternative mechanism, or pathway, that results in a faster rate.

Figure 14.2: Catalytic decomposition of hydrogen peroxide.

Photo courtesy of James Scherer.



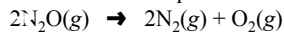
Factors Affecting Reaction Rates

Effect of Surface Area of Solid or Catalyst (in case of solids) or Pressure (gases)

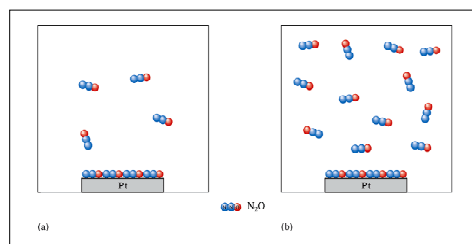
Increasing the surface area of a solid reactant or catalyst or the pressure of a gaseous reactant (without increasing the actual amount of the gas) usually increases the reaction rate.

In the case of the solid reactant or catalyst - there are more available reaction sites. In the case of the gas, the likelihood collisions resulting in reaction increases

Figure 12.8: The decomposition reaction:



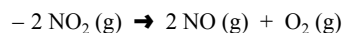
takes place on a platinum surface.



Types of Rate Laws

- **Differential Rate Law:** expresses how rate depends on concentration.
- **Integrated Rate Law:** expresses how concentration depends on time.

Dependence of Rate on Concentration



- When the concentration of nitrogen dioxide is doubled, the reaction rate doubles.
- **We need a mathematical expression** to relate the rate of the reaction to the concentrations of the reactants.

$$\text{Rate} = - \Delta[\text{NO}_2] / \Delta t = ?$$

Figure 12.2: Representation of the reaction:
 $2\text{NO}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{O}_2(\text{g})$.

- (a) The reaction at the very beginning ($t = 0$). (b) and (c) As time passes, NO_2 is converted to NO and O_2 .

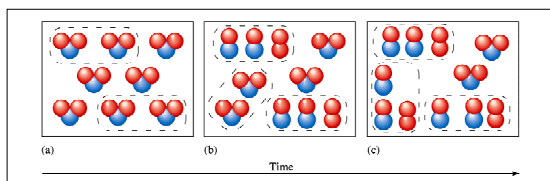


TABLE 12.1 Concentrations of Reactant and Products as a Function of Time for the Reaction $2\text{NO}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{O}_2(\text{g})$ (at 300°C)

Time (± 1 s)	Concentration (mol/L)		
	NO_2	NO	O_2
0	0.0100	0	0
50	0.0079	0.0021	0.0011
100	0.0065	0.0035	0.0018
150	0.0055	0.0045	0.0023
200	0.0048	0.0052	0.0026
250	0.0043	0.0057	0.0029
300	0.0038	0.0062	0.0031
350	0.0034	0.0066	0.0033
400	0.0031	0.0069	0.0035

Figure 12.1: Starting with a flask of nitrogen dioxide at 300°C , the concentrations of nitrogen dioxide, nitric oxide, and oxygen are plotted versus time.

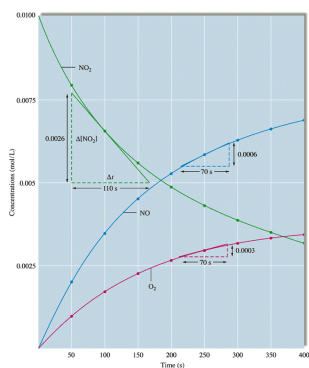


TABLE 12.2 Average Rate (in $\text{mol/L} \cdot \text{s}$) of Decomposition of Nitrogen Dioxide as a Function of Time*

$\frac{\Delta[\text{NO}_2]}{\Delta t}$	Time Period (s)
4.2×10^{-5}	0 \rightarrow 50
2.8×10^{-5}	50 \rightarrow 100
2.0×10^{-5}	100 \rightarrow 150
1.4×10^{-5}	150 \rightarrow 200
1.0×10^{-5}	200 \rightarrow 250

Rate Laws

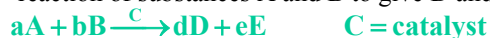
$$\text{Rate} = k[\text{NO}_2]^m$$

- k = rate constant
- m = rate order

If rate doubles when $[\text{NO}_2]$ is doubled, what is m ?

Dependence of Rate on Concentration

- As a more general example, consider the reaction of substances A and B to give D and E.



- You could write the rate law in the form

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

- The exponents m , n , and p are frequently, but not always, integers. They must be determined experimentally and cannot be obtained by simply looking at the balanced equation.

Dependence of Rate on Concentration

- Reaction Order
 - The **reaction order** with respect to a given reactant species **equals the exponent of the concentration of that species in the rate law, as determined experimentally.**
 - The **overall order** of the reaction **equals the sum of the orders of the reacting species in the rate law.**

Overall Reaction Order

- Sum of the order of each component in the rate law.

$$\bullet \text{rate} = k[\text{H}_2\text{SeO}_3][\text{H}^+]^2[\text{I}^-]^3$$

- The overall reaction order is $1 + 2 + 3 = 6$.

Dependence of Rate on Concentration

- Reaction Order
 - Consider the reaction of nitric oxide with hydrogen according to the following equation.



- The experimentally determined rate law is

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

- Thus, the reaction is second order in NO, first order in H_2 , and third order overall.

Dependence of Rate on Concentration

- Reaction Order
 - Although reaction orders frequently have whole number values (particularly 1 and 2), they can be **fractional**.
 - **Zero** and **negative** orders are also possible.
 - The concentration of a reactant with a zero-order dependence has **no effect** on the rate of the reaction.

Dependence of Rate on Concentration

- Determining the Rate Law.
 - One method for determining the order of a reaction with respect to each reactant is the **method of initial rates**.
 - It involves running the experiment multiple times, each time varying the concentration of only one reactant and measuring its initial rate.
 - The resulting change in rate indicates the order with respect to that reactant.

Dependence of Rate on Concentration

- Determining the Rate Law.
 - If **doubling the concentration** of a reactant has a **doubling** effect on the rate, then one would deduce it was a **first-order dependence**.
 - If **doubling the concentration** had a **quadrupling** effect on the rate, one would deduce it was a **second-order dependence**.
 - A **doubling of concentration** that results in an **eight-fold** increase in the rate would be a **third-order dependence**.

Table 14.1
Effect on Rate of Doubling the Initial Concentration of Reactant

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

m	Rate Is Multiplied by:
-1	$\frac{1}{2}$
0	1
1	2
2	4

A Problem to Consider

- Iodide ion is oxidized in acidic solution to triiodide ion, I_3^- , by hydrogen peroxide.



- A series of four experiments was run at different concentrations, and the initial rates of I_3^- formation were determined.
- From the following data, obtain the reaction orders with respect to H_2O_2 , I^- , and H^+ .
- Calculate the numerical value of the rate constant.

A Problem to Consider

	Initial Concentrations (mol/L)			Initial Rate [mol/(L·s)]
	H_2O_2	I^-	H^+	
Exp. 1	0.010	0.010	0.00050	1.15×10^{-6}
Exp. 2	0.020	0.010	0.00050	2.30×10^{-6}
Exp. 3	0.010	0.020	0.00050	2.30×10^{-6}
Exp. 4	0.010	0.010	0.00100	1.15×10^{-6}

- Comparing Experiment 1 and Experiment 2, you see that when the H_2O_2 concentration doubles (with other concentrations constant), the rate doubles.
- This implies a **first-order dependence with respect to H_2O_2** .

A Problem to Consider

	Initial Concentrations (mol/L)			Initial Rate [mol/(L·s)]
	H_2O_2	I^-	H^+	
Exp. 1	0.010	0.010	0.00050	1.15×10^{-6}
Exp. 2	0.020	0.010	0.00050	2.30×10^{-6}
Exp. 3	0.010	0.020	0.00050	2.30×10^{-6}
Exp. 4	0.010	0.010	0.00100	1.15×10^{-6}

- Comparing Experiment 1 and Experiment 3, you see that when the I^- concentration doubles (with other concentrations constant), the rate doubles.
- This implies a **first-order dependence with respect to I^-** .

A Problem to Consider

	Initial Concentrations (mol/L)			Initial Rate [mol/(L·s)]
	H ₂ O ₂	I ⁻	H ⁺	
Exp. 1	0.010	0.010	0.00050	1.15 × 10 ⁻⁶
Exp. 2	0.020	0.010	0.00050	2.30 × 10 ⁻⁶
Exp. 3	0.010	0.020	0.00050	2.30 × 10 ⁻⁶
Exp. 4	0.010	0.010	0.00100	1.15 × 10 ⁻⁶

- Comparing Experiment 1 and Experiment 4, you see that when the H⁺ concentration doubles (with other concentrations constant), the rate is unchanged.
- This implies a **zero-order dependence with respect to H⁺**.

A Problem to Consider

	Initial Concentrations (mol/L)			Initial Rate [mol/(L·s)]
	H ₂ O ₂	I ⁻	H ⁺	
Exp. 1	0.010	0.010	0.00050	1.15 × 10 ⁻⁶
Exp. 2	0.020	0.010	0.00050	2.30 × 10 ⁻⁶
Exp. 3	0.010	0.020	0.00050	2.30 × 10 ⁻⁶
Exp. 4	0.010	0.010	0.00100	1.15 × 10 ⁻⁶

- Because [H⁺]⁰ = 1, the rate law is:

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

- The reaction orders with respect to H₂O₂, I⁻, and H⁺, are **1**, **1**, and **0**, respectively.

A Problem to Consider

	Initial Concentrations (mol/L)			Initial Rate [mol/(L·s)]
	H ₂ O ₂	I ⁻	H ⁺	
Exp. 1	0.010	0.010	0.00050	1.15 × 10 ⁻⁶
Exp. 2	0.020	0.010	0.00050	2.30 × 10 ⁻⁶
Exp. 3	0.010	0.020	0.00050	2.30 × 10 ⁻⁶
Exp. 4	0.010	0.010	0.00100	1.15 × 10 ⁻⁶

- You can now calculate the rate constant by substituting values from any of the experiments. Using Experiment 1 you obtain:

$$k = \frac{1.15 \times 10^{-6} \text{ s}^{-1}}{0.010 \times 0.010 \text{ mol/L}} = 1.2 \times 10^{-2} \text{ L}/(\text{mol} \cdot \text{s})$$

A Problem to Consider

	Initial Concentrations (mol/L)			Initial Rate [mol/(L·s)]
	H ₂ O ₂	I ⁻	H ⁺	
Exp. 1	0.010	0.010	0.00050	1.15 × 10 ⁻⁶
Exp. 2	0.020	0.010	0.00050	2.30 × 10 ⁻⁶
Exp. 3	0.010	0.020	0.00050	2.30 × 10 ⁻⁶
Exp. 4	0.010	0.010	0.00100	1.15 × 10 ⁻⁶

- You can now calculate the rate constant by substituting values from any of the experiments. Using Experiment 1 you obtain:

$$1.15 \times 10^{-6} \frac{\text{mol}}{\text{L} \cdot \text{s}} = k \times 0.010 \frac{\text{mol}}{\text{L}} \times 0.010 \frac{\text{mol}}{\text{L}}$$

TABLE 12.5 The Results from Four Experiments to Study the Reaction
 $\text{BrO}_3^-(\text{aq}) + 5\text{Br}^-(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow 3\text{Br}_2(\text{l}) + 3\text{H}_2\text{O}(\text{l})$

Experiment	Initial Concentration of BrO ₃ ⁻ (mol/L)	Initial Concentration of Br ⁻ (mol/L)	Initial Concentration of H ⁺ (mol/L)	Measured Initial Rate (mol/L · s)
1	0.10	0.10	0.10	8.0 × 10 ⁻⁴
2	0.20	0.10	0.10	1.6 × 10 ⁻³
3	0.20	0.20	0.10	3.2 × 10 ⁻³
4	0.10	0.10	0.20	3.2 × 10 ⁻³

TABLE 12.4 Initial Rates from Three Experiments for the Reaction
 $\text{NH}_4^+(\text{aq}) + \text{NO}_2^-(\text{aq}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Experiment	Initial Concentration of NH ₄ ⁺	Initial Concentration of NO ₂ ⁻	Initial Rate (mol/L · s)
1	0.100 M	0.0050 M	1.35 × 10 ⁻⁷
2	0.100 M	0.010 M	2.70 × 10 ⁻⁷
3	0.200 M	0.010 M	5.40 × 10 ⁻⁷

Change of Concentration with Time

- A rate law simply tells you how the rate of reaction changes as reactant concentrations change.
- A more useful mathematical relationship would show how a reactant concentration changes over a period of time.

Change of Concentration with Time

- A rate law simply tells you how the rate of reaction changes as reactant concentrations change.
- Using calculus we can transform a rate law into a mathematical relationship between concentration and time.
- This provides a **graphical method** for determining rate laws.

Concentration-Time Equations

- First-Order Rate Law
 - You could write the rate law in the form

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

Concentration-Time Equations

- First-Order Rate Law
 - Using calculus, you get the following equation.

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

- Here $[A]_t$ is the concentration of reactant A at time t, and $[A]_0$ is the initial concentration.
- The ratio $[A]_t/[A]_0$ is the fraction of A remaining at time t.

TABLE 12.3
Concentration/Time Data for the Reaction $2\text{N}_2\text{O}_5(\text{soln}) \rightarrow 4\text{NO}_2(\text{soln}) + \text{O}_2(\text{g})$ (at 45°C)

$[\text{N}_2\text{O}_5]$ (mol/L)	Time (s)
1.00	0
0.88	200
0.78	400
0.69	600
0.61	800
0.54	1000
0.48	1200
0.43	1400
0.38	1600
0.34	1800
0.30	2000

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Figure 12.3: A plot of the concentration of N_2O_5 is a function of time for the reaction:
 $2\text{N}_2\text{O}_5(\text{soln}) \rightarrow 4\text{NO}_2(\text{soln}) + \text{O}_2(\text{g})$ (at 45°C).
Note that the reaction rate at $[\text{N}_2\text{O}_5] = 0.90 \text{ M}$ is twice that at $[\text{N}_2\text{O}_5] = 0.45 \text{ M}$.

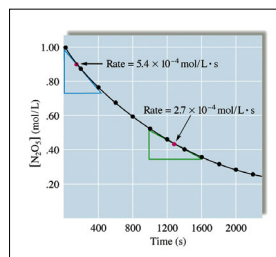
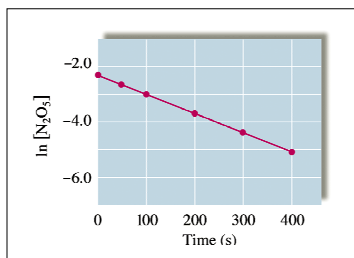


Figure 12.4: A plot of $\ln[\text{N}_2\text{O}_5]$ versus time.



A Problem to Consider

- The decomposition of N_2O_5 to NO_2 and O_2 is first order with a rate constant of $4.8 \times 10^{-4} \text{ s}^{-1}$. If the initial concentration of N_2O_5 is $1.65 \times 10^{-2} \text{ mol/L}$, **what is the concentration of N_2O_5 after 825 seconds?**
- The first-order time-concentration equation for this reaction would be:

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

A Problem to Consider

- The decomposition of N_2O_5 to NO_2 and O_2 is first order with a rate constant of $4.8 \times 10^{-4} \text{ s}^{-1}$. If the initial concentration of N_2O_5 is $1.65 \times 10^{-2} \text{ mol/L}$, what is the concentration of N_2O_5 after 825 seconds?
- Substituting the given information we obtain:

$$\ln \frac{[\text{N}_2\text{O}_5]_t}{1.65 \times 10^{-2} \text{ mol/L}} = -(4.80 \times 10^{-4} \text{ s}^{-1}) \times (825 \text{ s})$$

A Problem to Consider

- The decomposition of N_2O_5 to NO_2 and O_2 is first order with a rate constant of $4.8 \times 10^{-4} \text{ s}^{-1}$. If the initial concentration of N_2O_5 is $1.65 \times 10^{-2} \text{ mol/L}$, what is the concentration of N_2O_5 after 825 seconds?
- Substituting the given information we obtain:

$$\ln \frac{[\text{N}_2\text{O}_5]_t}{1.65 \times 10^{-2} \text{ mol/L}} = -0.396$$

A Problem to Consider

- The decomposition of N_2O_5 to NO_2 and O_2 is first order with a rate constant of $4.8 \times 10^{-4} \text{ s}^{-1}$. If the initial concentration of N_2O_5 is $1.65 \times 10^{-2} \text{ mol/L}$, what is the concentration of N_2O_5 after 825 seconds?
- Taking the inverse natural log of both sides we obtain:

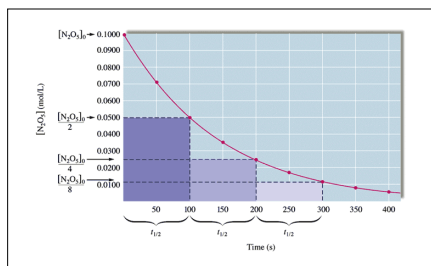
$$\frac{[\text{N}_2\text{O}_5]_t}{1.65 \times 10^{-2} \text{ mol/L}} = e^{-0.396} = 0.673$$

A Problem to Consider

- The decomposition of N_2O_5 to NO_2 and O_2 is first order with a rate constant of $4.8 \times 10^{-4} \text{ s}^{-1}$. If the initial concentration of N_2O_5 is $1.65 \times 10^{-2} \text{ mol/L}$, what is the concentration of N_2O_5 after 825 seconds?
- Solving for $[\text{N}_2\text{O}_5]$ at 825 s we obtain:

$$[\text{N}_2\text{O}_5] = (1.65 \times 10^{-2} \text{ mol/L}) \times (0.673) = 0.0111 \text{ mol/L}$$

Figure 12.5: A plot of $[\text{N}_2\text{O}_5]$ versus time for the decomposition reaction of N_2O_5 .



Half-Life of a First-Order Reaction

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

- $t_{1/2}$ = half-life of the reaction
- k = rate constant
- For a first-order reaction, the half-life does not depend on concentration.

Second-Order Rate Law

- For $aA \rightarrow$ products in a second-order reaction,

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

- Integrated rate law is

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

Concentration-Time Equations

- Second-Order Rate Law
 - You could write the rate law in the form

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

Concentration-Time Equations

- Second-Order Rate Law
 - Using calculus, you get the following equation.

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

- Here $[A]_t$ is the concentration of reactant A at time t , and $[A]_0$ is the initial concentration.

Figure 12.6: (a) A plot of $\ln[C_4H_6]$ versus t . (b) A plot of $1/[C_4H_6]$ versus t .

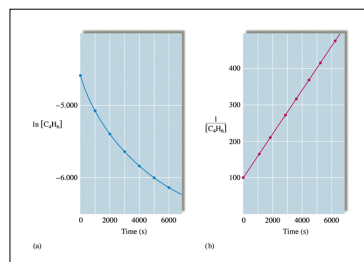
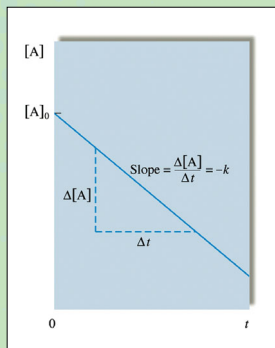


Figure 12.7:
A plot of $[A]$
versus t for a
zero-order
reaction.



Half-Life of a Second-Order Reaction

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

- $t_{1/2}$ = half-life of the reaction
- k = rate constant
- A_0 = initial concentration of A
- The half-life is dependent upon the initial concentration.

A Summary

1. **Simplification:** Conditions are set such that only forward reaction is important.
2. **Two types:** differential rate law
integrated rate law
3. **Which type?** Depends on the type of data collected - differential and integrated forms can be interconverted.

A Summary (continued)

4. **Most common:** method of initial rates.
5. **Concentration v. time:** used to determine integrated rate law, often graphically.
6. **For several reactants:** choose conditions under which only one reactant varies significantly (pseudo first-order conditions).

TABLE 12.6 Summary of the Kinetics for Reactions of the Type $aA \rightarrow$ Products That Are Zero, First, or Second Order in $[A]$

	Order		
	Zero	First	Second
Rate Law:	Rate = k	Rate = $k[A]$	Rate = $k[A]^2$
Integrated Rate Law:	$[A] = -kt + [A]_0$	$\ln[A] = -kt + \ln[A]_0$	$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$
Plot Needed to Give a Straight Line:	$[A]$ versus t	$\ln[A]$ versus t	$\frac{1}{[A]}$ versus t
Relationship of Rate Constant to the Slope of Straight Line:	Slope = $-k$	Slope = $-k$	Slope = k
Half-life:	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{0.693}{k}$	$t_{1/2} = \frac{1}{k[A]_0}$