

# Chemical Kinetics Practice Problems And Answers

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Chemical Kinetics Practice Problems and Answers Mastering the Rate of Reactions Chemical kinetics is the study of reaction rates and reaction mechanisms. It's a fascinating field that delves into the intricate dance of molecules as they transform from reactants to products. Understanding chemical kinetics is crucial for a wide range of applications, from designing efficient catalysts to predicting the stability of materials. This article aims to provide a comprehensive set of practice problems and solutions designed to help you solidify your understanding of chemical kinetics. We'll cover essential concepts like rate laws, integrated rate laws, activation energy, and reaction mechanisms, along with examples that illustrate their applications.

## 1 Rate Laws and Rate Constants

The rate law for a reaction describes the relationship between the rate of reaction and the concentration of reactants. For example, the general rate law for a reaction  $A + B \rightarrow C$  can be written as  $\text{Rate} = k[A]^m[B]^n$ . Where:

- Rate: The rate of the reaction
- $k$ : The rate constant, a proportionality constant specific to the reaction at a given temperature
- $A$  and  $B$ : The concentrations of reactants
- $m$  and  $n$ : The reaction orders with respect to  $A$  and  $B$  respectively, which are determined experimentally

**Practice Problem 1** The following data were collected for the reaction  $2A + B \rightarrow C$ :

Experiment	$[A]$ (M)	$[B]$ (M)	Initial Rate (Ms <sup>-1</sup> )
1	0.10	0.10	0.050
2	0.20	0.10	0.200
3	0.10	0.20	0.100

**Determine the rate law for the reaction and calculate the rate constant.**

**Solution 1**

- Determine the order with respect to  $A$ . Compare experiments 1 and 2 where  $B$  is constant but  $A$  doubles. The rate quadruples, indicating a second-order dependence on  $A$ .  $2^2 = 4$ .
- Determine the order with respect to  $B$ . Compare experiments 1 and 3 where  $A$  is constant but  $B$  doubles. The rate doubles, indicating a first-order dependence on  $B$ .
- Write the rate law:  $\text{Rate} = k[A]^2[B]$ .
- Calculate the rate constant. Using data from any experiment, we can solve for  $k$ .  $0.050 \text{ Ms}^{-1} = k(0.10 \text{ M})^2(0.10 \text{ M})$ .  $k = 5.0 \text{ M}^{-2}\text{s}^{-1}$ .

## 2 Integrated Rate Laws

Integrated rate laws relate the concentration of reactants to time. The form of the integrated rate law depends on the order of the reaction.

**For a first-order reaction:** Integrated rate law:  $\ln[A]_t = \ln[A]_0 - kt$ . Half-life:  $t_{1/2} = \frac{\ln 2}{k}$ .

**For a second-order reaction:** Integrated rate law:  $\frac{1}{[A]_t} = \frac{1}{[A]_0} + kt$ . Half-life:  $t_{1/2} = \frac{1}{k[A]_0}$ .

**Practice Problem 2** The decomposition of  $\text{N}_2\text{O}_5$  is a first-order reaction with a rate constant of  $6.8 \times 10^{-3} \text{ s}^{-1}$  at a certain temperature.

- Calculate the half-life of the reaction.
- If the initial concentration of  $\text{N}_2\text{O}_5$  is  $0.50 \text{ M}$ , what will be the concentration after 2 minutes?

**Solution 3**

- Half-life  $t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{6.8 \times 10^{-3} \text{ s}^{-1}} = 101.6 \text{ s}$ .
- Concentration after 2 minutes: First, convert 2 minutes to seconds:  $2 \text{ minutes} = 120 \text{ seconds}$ .  $\ln[\text{N}_2\text{O}_5]_t = \ln[\text{N}_2\text{O}_5]_0 - kt$ .  $\ln[\text{N}_2\text{O}_5]_t = \ln(0.50 \text{ M}) - (6.8 \times 10^{-3} \text{ s}^{-1})(120 \text{ s})$ .  $\ln[\text{N}_2\text{O}_5]_t = -0.816$ .  $[\text{N}_2\text{O}_5]_t = e^{-0.816} = 0.296 \text{ M}$ .

## 3 Activation Energy and the

**Arrhenius Equation** The activation energy  $E_a$  is the minimum amount of energy required for reactants to overcome the energy barrier and form products. The Arrhenius equation relates the rate constant  $k$  to the activation energy and temperature  $T$ :  $k = Ae^{-E_a/RT}$ . Where  $A$  is the preexponential factor which is related to the frequency of collisions between molecules,  $R$  is the ideal gas constant  $8.314 \text{ J/mol}\cdot\text{K}$ .

**Practice Problem 3** The rate constant for a reaction is  $12 \times 10^2 \text{ s}^{-1}$  at  $25^\circ\text{C}$ . The activation energy is  $50 \text{ kJ/mol}$ . Calculate the rate constant at  $45^\circ\text{C}$ .

**Solution** First convert temperatures to Kelvin:  $T_1 = 25^\circ\text{C} = 273.15 + 25 = 298.15 \text{ K}$ ,  $T_2 = 45^\circ\text{C} = 273.15 + 45 = 318.15 \text{ K}$ . Use the Arrhenius equation in the form  $\ln k_2 = \ln A - E_a/RT_2$  and  $\ln k_1 = \ln A - E_a/RT_1$ . Subtracting these equations gives  $\ln k_2 - \ln k_1 = E_a/R \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$ .  
 $\ln \left( \frac{k_2}{k_1} \right) = \frac{50,000 \text{ J/mol}}{8.314 \text{ J/mol}\cdot\text{K}} \left( \frac{1}{298.15 \text{ K}} - \frac{1}{318.15 \text{ K}} \right)$   
 $\ln \left( \frac{k_2}{12 \times 10^2 \text{ s}^{-1}} \right) = 6.0693$   
 $\frac{k_2}{12 \times 10^2 \text{ s}^{-1}} = e^{6.0693} = 424$   
 $k_2 = 424 \times 12 \times 10^2 \text{ s}^{-1} = 5.09 \times 10^5 \text{ s}^{-1}$

**Reaction Mechanisms** A reaction mechanism is a step-by-step description of how a reaction proceeds. Each step involves a single molecular event called an elementary reaction.

**Practice Problem 4** The following mechanism has been proposed for the reaction  $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$ :  
 Step 1:  $2\text{NO} + \text{O}_2 \rightarrow \text{N}_2\text{O}_2$  (fast)  
 Step 2:  $\text{N}_2\text{O}_2 + \text{O}_2 \rightarrow 2\text{NO}_2$  (slow)

**a** Identify the rate-determining step. **b** Derive the rate law for the overall reaction.

**Solution** **a** Rate-determining step: The slow step in a reaction mechanism is the rate-determining step as it determines the overall rate of the reaction. In this case, Step 2 is the rate-determining step. **b** Rate law: The rate law for the overall reaction is based on the rate-determining step:  $\text{Rate} = k_2[\text{N}_2\text{O}_2][\text{O}_2]$ . However,  $\text{N}_2\text{O}_2$  is an intermediate and its concentration cannot be measured directly. We need to express  $\text{N}_2\text{O}_2$  in terms of reactants using the equilibrium constant for the fast step:  $K_1 = \frac{[\text{N}_2\text{O}_2]}{[\text{NO}]^2[\text{O}_2]}$ . Substitute this back into the rate law:  $\text{Rate} = k_2 K_1 [\text{NO}]^2 [\text{O}_2]^2 = k [\text{NO}]^2 [\text{O}_2]^2$ , where  $k = k_2 K_1$ .

**Key Takeaways** Chemical kinetics deals with the rate and mechanism of chemical reactions. Rate laws describe how the rate of a reaction depends on reactant concentrations. Integrated rate laws relate reactant concentration to time. Activation energy is the minimum energy required for a reaction to occur.

**5 Reaction mechanisms** provide a detailed description of how a reaction proceeds step-by-step. Further Practice To further enhance your understanding of chemical kinetics, explore additional practice problems from textbooks and online resources. Focus on understanding the concepts behind each problem and applying the appropriate equations. Remember, chemical kinetics is a vast field with numerous applications. By mastering the fundamental principles, you'll be well-equipped to tackle real-world problems and contribute to scientific advancements.

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