Chemical Kinetics - Practice Problems

1. Which of the following graphs best describes the relationship between the rate of a reaction and the temperature of the reaction?
2. Describe the difference between the terms *rate of reaction, rate law, and rate constant.* Give an example of each.

3. What are the units of the rate constant for the following reaction?

\[ 2\text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \quad \text{Rate} = k \, [\text{NO}_2]^2 \]

4. Calculate the rate at which \( \text{N}_2\text{O}_4 \) is formed in the following reaction at the moment in time when \( \text{NO}_2 \) is being consumed at a rate of 0.0592 M/s:

\[ 2\text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \]

5. NO reacts with \( \text{H}_2 \) according to the following equation:

\[ 2\text{NO}(g) + 2\text{H}_2(g) \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(g) \]

The mechanism for this reaction involves two steps:

- \( 2\text{NO} + \text{H}_2 \rightarrow \text{N}_2 + \text{H}_2\text{O}_2 \) (slow step)
- \( \text{H}_2\text{O}_2 + \text{H}_2 \rightarrow 2\text{H}_2\text{O} \) (fast step)

What is the experimental rate law for this reaction?

6. The disproportionation of NO to \( \text{N}_2\text{O} \) and \( \text{NO}_2 \) is third order in NO:

\[ 3\text{NO}(g) \rightarrow \text{N}_2\text{O}(g) + \text{NO}_2(g) \quad \text{Rate} = k \, [\text{NO}]^3 \]

This rate law is consistent with which of the following mechanisms?

- a) \( 3\text{NO} \rightarrow \text{N}_2\text{O} + \text{NO}_2 \) (one-step)
- b) \( 2\text{NO} \rightarrow \text{N}_2\text{O}_2 \)
  \( \text{N}_2\text{O}_2 + \text{NO} \rightarrow \text{N}_2\text{O} + \text{NO}_2 \) (fast step)
- c) \( 2\text{NO} \quad \text{N}_2\text{O}_2 \)
  \( \text{N}_2\text{O}_2 + \text{NO} \rightarrow \text{N}_2\text{O} + \text{NO}_2 \) (slow step)
7. Which graph best describes the rate of the following reaction if this reaction is first order in $\text{N}_2\text{O}_4$?

$$\text{N}_2\text{O}_4(g) \rightarrow 2\text{NO}_2(g)$$
8. The reaction in which NO\textsubscript{2} forms a dimer is second order in NO\textsubscript{2}:

\[ 2\text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \quad \text{Rate} = k \left[\text{NO}_2\right]^2 \]

Calculate the rate constant for this reaction if it takes 0.005 s for the initial concentration of NO\textsubscript{2} to decrease from 0.50M to 0.25M.

9. The decomposition of hydrogen peroxide is first order in H\textsubscript{2}O\textsubscript{2}:

\[ 2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) \quad \text{Rate} = k \left[\text{H}_2\text{O}_2\right] \]

How long will it take for half of the H\textsubscript{2}O\textsubscript{2} in a 10-gal sample to be consumed if the rate constant for this reaction is 5.6 \times 10^{-2} \text{ s}^{-1}?

10. Calculate the rate constant for the following acid-base reaction if the half-life for this reaction is 0.0282 s at 25\textdegree C and the reaction is first order in the NH\textsubscript{4}\textsuperscript{+} ion:

\[ \text{NH}_4\textsuperscript{+}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq) \]

11. Use the following data to determine the rate law for the decomposition of N\textsubscript{2}O.

\[ 2\text{N}_2\text{O}(g) \rightarrow 2\text{N}_2(g) + \text{O}_2(g) \]

<table>
<thead>
<tr>
<th><a href="M">N\textsubscript{2}O</a>:</th>
<th>0.100</th>
<th>0.086</th>
<th>0.079</th>
<th>0.075</th>
<th>0.066</th>
<th>0.059</th>
<th>0.049</th>
</tr>
</thead>
<tbody>
<tr>
<td>Time (s):</td>
<td>0</td>
<td>80</td>
<td>120</td>
<td>160</td>
<td>240</td>
<td>320</td>
<td>480</td>
</tr>
</tbody>
</table>

12. Use the results of the preceding problem to calculate the rate constant for this reaction. Predict the concentration of N\textsubscript{2}O after 900 s.

13. Dimethyl ether, CH\textsubscript{3}OCH\textsubscript{3}, decomposes at high temperatures as shown in the following equation:

\[ \text{CH}_3\text{OCH}_3(g) \rightarrow \text{CH}_4(g) + \text{H}_2(g) + \text{CO}(g) \]

The following data were obtained when the partial pressure of CH\textsubscript{3}OCH\textsubscript{3} was studied as this compound decomposed at 500\textdegree C. Use these data to determine the order of this reaction.

<table>
<thead>
<tr>
<th>P_CH\textsubscript{3}OCH\textsubscript{3} (mmHg):</th>
<th>312</th>
<th>278</th>
<th>251</th>
<th>227</th>
<th>157</th>
</tr>
</thead>
<tbody>
<tr>
<td>Time (s):</td>
<td>0</td>
<td>390</td>
<td>777</td>
<td>1195</td>
<td>3155</td>
</tr>
</tbody>
</table>
14. The rate constant for the decomposition of $\text{N}_2\text{O}_5$ increases from $1.52 \times 10^{-5} \text{ s}^{-1}$ at $25^\circ \text{C}$ to $3.83 \times 10^{-3} \text{ s}^{-1}$ at $45^\circ \text{C}$. Calculate the activation energy for this reaction.

15. Calculate the activation energy for the following reaction if the rate constant for this reaction increases from $87.1 \text{ M}^{-1}\text{s}^{-1}$ at $500 \text{ K}$ to $1.53 \times 10^{3} \text{ M}^{-1}\text{s}^{-1}$ at $650 \text{ K}$:

$$2\text{NO}_2(g) \rightarrow 2\text{NO}(g) + \text{O}_2(g)$$

16. Calculate the activation energy for the decomposition of $\text{NO}_2$ from the temperature dependence of the rate constant for this reaction:

$$2\text{NO}_2(g) \rightarrow \text{N}_2(g) + 2\text{O}_2(g)$$

<table>
<thead>
<tr>
<th>Temperature (K)</th>
<th>319</th>
<th>329</th>
<th>352</th>
<th>381</th>
<th>389</th>
</tr>
</thead>
<tbody>
<tr>
<td>$k \ (\text{M}^{-1}\text{s}^{-1})$</td>
<td>0.522</td>
<td>0.755</td>
<td>1.70</td>
<td>4.02</td>
<td>5.03</td>
</tr>
</tbody>
</table>

17. Calculate the rate constant at $780 \text{ K}$ for the following reaction if the rate constant for the reaction is $3.5 \times 10^{-7} \text{ M}^{-1}\text{s}^{-1}$ at $550 \text{ K}$ and the activation energy is $188 \text{ kJ/mol}$:

$$2\text{HI}(g) \rightarrow \text{H}_2(g) + \text{I}_2(g)$$

18. Calculate the rate constant at $75^\circ \text{C}$ for the following reaction if the rate constant for this reaction is $6.5 \times 10^{-5} \text{ M}^{-1}\text{s}^{-1}$ at $25^\circ \text{C}$ and the activation energy is $92.9 \text{ kJ/mol}$:

$$\text{CH}_3\text{I}(aq) + \text{OH}^-(aq) \rightarrow \text{CH}_3\text{OH}(aq) + \text{I}^-(aq)$$