1. The compound RX₃ decomposes according to the equation
   \[ 3RX₃ \rightarrow R + R₂X₃ + 3X₂ \]
   In an experiment the following data were collected for the decomposition at 100°C. What is the average rate of reaction over the entire experiment?
   \[ \begin{array}{cc}
   t(s) & [RX₃](\text{mol L}^{-1}) \\
   0 & 0.85 \\
   2 & 0.67 \\
   6 & 0.41 \\
   8 & 0.33 \\
   12 & 0.20 \\
   14 & 0.16 \\
   \end{array} \]
   A) 0.011 mol L⁻¹s⁻¹
   B) 0.019 mol L⁻¹s⁻¹
   C) 0.044 mol L⁻¹s⁻¹
   D) 0.049 mol L⁻¹s⁻¹
   E) 0.069 mol L⁻¹s⁻¹

2. Consider the general reaction
   \[ 5Br⁻(aq) + BrO₃⁻(aq) + 6H⁺(aq) \rightarrow 3Br₂(aq) + 3H₂O(aq) \]
   For this reaction, the rate when expressed as \( \Delta[Br₂]/\Delta t \) is the same as
   A) \( -\Delta[H₂O]/\Delta t \)
   B) \( 3\Delta[BrO₃⁻]/\Delta t \)
   C) \( -5\Delta[Br⁻]/\Delta t \)
   D) \( -0.6\Delta[Br⁻]/\Delta t \)
   E) none of the above
3. For the reaction
   \[3A(g) + 2B(g) \rightarrow 2C(g) + 2D(g)\]
the following data were collected at constant temperature. Determine the correct rate law
for this reaction.

<table>
<thead>
<tr>
<th>Trial</th>
<th>Initial [A] (mol/L)</th>
<th>Initial [B] (mol/L)</th>
<th>Initial Rate (mol/(L·min))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.200</td>
<td>0.100</td>
<td>6.00 \times 10^{-2}</td>
</tr>
<tr>
<td>2</td>
<td>0.100</td>
<td>0.100</td>
<td>1.50 \times 10^{-2}</td>
</tr>
<tr>
<td>3</td>
<td>0.200</td>
<td>0.200</td>
<td>1.20 \times 10^{-1}</td>
</tr>
<tr>
<td>4</td>
<td>0.300</td>
<td>0.200</td>
<td>2.70 \times 10^{-1}</td>
</tr>
</tbody>
</table>

A) \(\text{Rate} = k[A][B]\)
B) \(\text{Rate} = k[A][B]^2\)
C) \(\text{Rate} = k[A]^2[B]^2\)
D) \(\text{Rate} = k[A]^3[B]\)
E) \(\text{Rate} = k[A]^2[B]\)

4. For the reaction
   \[2A + B + 2C \rightarrow D + E\]
the following initial rate data were collected at constant temperature. Determine the
correct rate law for this reaction. All units are arbitrary.

<table>
<thead>
<tr>
<th>Trial</th>
<th>[A]</th>
<th>[B]</th>
<th>[C]</th>
<th>Rate</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.225</td>
<td>0.150</td>
<td>0.350</td>
<td>0.0217</td>
</tr>
<tr>
<td>2</td>
<td>0.320</td>
<td>0.150</td>
<td>0.350</td>
<td>0.0439</td>
</tr>
<tr>
<td>3</td>
<td>0.225</td>
<td>0.250</td>
<td>0.350</td>
<td>0.0362</td>
</tr>
<tr>
<td>4</td>
<td>0.225</td>
<td>0.150</td>
<td>0.600</td>
<td>0.01270</td>
</tr>
</tbody>
</table>

A) \(\text{Rate} = k[A][B][C]\)
B) \(\text{Rate} = k[A]^2[B][C]\)
C) \(\text{Rate} = k[A]^2[B][C]^{-1}\)
D) \(\text{Rate} = k[A][B]^2[C]^{-1}\)
E) none of the above

5. The rate constant for a reaction is 4.65 L·mol\(^{-1}\)·s\(^{-1}\). What is the overall order of the
reaction?
A) zero
B) first
C) second
D) third
E) More information is needed to determine the overall order.
6. When the reaction $A \rightarrow B + C$ is studied, a plot of $\ln[A]$, vs. time gives a straight line with a negative slope. What is the order of the reaction?
   A) zero
   B) first
   C) second
   D) third
   E) More information is needed to determine the order.

7. When the reaction $A \rightarrow B + C$ is studied, a plot $1/[A]$, vs. time gives a straight line with a positive slope. What is the order of the reaction?
   A) zero
   B) first
   C) second
   D) third
   E) More information is needed to determine the order.

8. Which one of the following sets of units is appropriate for a second-order rate constant?
   A) $s^{-1}$
   B) $\text{mol L}^{-1} \text{s}^{-1}$
   C) $L \text{mol}^{-1} \text{s}^{-1}$
   D) $\text{mol}^2 \text{L}^{-2} \text{s}^{-1}$
   E) $L^2 \text{mol}^{-2} \text{s}^{-1}$

9. Ammonium cyanate ($\text{NH}_4\text{CNO}$) reacts to form urea ($\text{NH}_2\text{CONH}_2$). At 65°C the rate constant, $k$, is $3.60 \text{ L mol}^{-1} \text{s}^{-1}$. What is the rate law for this reaction?
   A) Rate = $3.60 \text{ L mol}^{-1} \text{s}^{-1}[\text{NH}_4\text{CNO}]$
   B) Rate = $3.60 \text{ L mol}^{-1} \text{s}^{-1}[\text{NH}_4\text{CNO}]^2$
   C) Rate = $0.28 \text{ mol L}^{-1} \text{s}^{-1}[\text{NH}_4\text{CNO}]$
   D) Rate = $0.28 \text{ mol L}^{-1} \text{s}^{-1}[\text{NH}_4\text{CNO}]^2$
   E) Rate = $3.60 \text{ L mol}^{-1} \text{s}^{-1}[\text{NH}_2\text{CONH}_2]^{-1}$
10. $2\text{NOBr}(g) \rightarrow 2\text{NO}(g) + \text{Br}_2(g)$

<table>
<thead>
<tr>
<th>$[\text{NOBr}](\text{mol L}^{-1})$</th>
<th>Rate (mol L$^{-1}$s$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0450</td>
<td>$1.62 \times 10^{-3}$</td>
</tr>
<tr>
<td>0.0310</td>
<td>$7.69 \times 10^{-4}$</td>
</tr>
<tr>
<td>0.0095</td>
<td>$7.22 \times 10^{-5}$</td>
</tr>
</tbody>
</table>

Based on the initial rate data above, what is the value of the rate constant?
A) $0.0360$ L mol$^{-1}$s$^{-1}$
B) $0.800$ L mol$^{-1}$s$^{-1}$
C) $1.25$ L mol$^{-1}$s$^{-1}$
D) $27.8$ L mol$^{-1}$s$^{-1}$
E) $0.0360$ s$^{-1}$

11. A reaction is first-order with respect to the reactant R. Which of the following plots will produce a straight line?
A) $[R]$ vs. 1/time
B) $1/[R]$ vs. time
C) $[R]^2$ vs. time
D) $1/[R]^2$ vs. time
E) ln$[R]$ vs. time

12. Cyclopropane is converted to propene in a first-order process. The rate constant is $5.4 \times 10^{-2}$ hr$^{-1}$. If the initial concentration of cyclopropane is $0.150$ M, what will its concentration be after 22.0 hours?
A) $0.0457$ M
B) $0.105$ M
C) $0.127$ M
D) $0.492$ M
E) none of the above

13. A gas-phase decomposition is first-order with respect to the reactant, R. If the initial concentration of R is $1.0 \times 10^{-4}$ mol L$^{-1}$ and the rate constant $k = 1.08 \times 10^{-6}$ s$^{-1}$, what concentration of R remains after 25 days?
A) $1.0 \times 10^{-3}$ mol L$^{-1}$
B) $1.0 \times 10^{-4}$ mol L$^{-1}$
C) $9.6 \times 10^{-5}$ mol L$^{-1}$
D) $4.3 \times 10^{-5}$ mol L$^{-1}$
E) $9.7 \times 10^{-4}$ mol L$^{-1}$
14. The rate law for the rearrangement of CH₃NC to CH₂CN at 800 K is Rate = (1300 s⁻¹)[CH₃NC]. What is the half-life for this reaction?
   A) 7.69 × 10⁻⁴ s
   B) 5.3 × 10⁻⁴ s
   C) 1.9 × 10⁻³ s
   D) 520 s
   E) 1920 s

15. Consider the following mechanism for the oxidation of bromide ions by hydrogen peroxide in aqueous acid solution.
   \[ \text{H}^+ + \text{H}_2\text{O}_2 \rightleftharpoons \text{H}_2\text{O}^+\text{--OH} \text{ (rapid equilibrium)} \]
   \[ \text{H}_2\text{O}^+\text{--OH} + \text{Br}^- \rightarrow \text{HOBr} + \text{H}_2\text{O} \text{ (slow)} \]
   \[ \text{HOBr} + \text{H}^+ + \text{Br}^- \rightarrow \text{Br}_2 + \text{H}_2\text{O} \text{ (fast)} \]
   What is the overall reaction equation for this process?
   A) \(2\text{H}_2\text{O}^+\text{--OH} + 2\text{Br}^- \rightarrow \text{H}_2\text{O}_2 + \text{Br}_2 + 2\text{H}_2\text{O}\)
   B) \(2\text{H}^+ + 2\text{Br}^- + \text{H}_2\text{O}_2 \rightarrow \text{Br}_2^+ + 2\text{H}_2\text{O}\)
   C) \(2\text{H}^+ + \text{H}_2\text{O}_2 + \text{Br}^- + \text{HOBr} \rightarrow \text{H}_2\text{O}^+\text{--OH} + \text{Br}_2 + \text{H}_2\text{O}\)
   D) \(\text{H}_2\text{O}^+\text{--OH} + \text{Br}^- + \text{H}^+ \rightarrow \text{Br}_2 + \text{H}_2\text{O}\)
   E) none of the above

16. Consider the following mechanism for the oxidation of bromide ions by hydrogen peroxide in aqueous acid solution.
   \[ \text{H}^+ + \text{H}_2\text{O}_2 \rightleftharpoons \text{H}_2\text{O}^+\text{--OH} \text{ (rapid equilibrium)} \]
   \[ \text{H}_2\text{O}^+\text{--OH} + \text{Br}^- \rightarrow \text{HOBr} + \text{H}_2\text{O} \text{ (slow)} \]
   \[ \text{HOBr} + \text{H}^+ + \text{Br}^- \rightarrow \text{Br}_2 + \text{H}_2\text{O} \text{ (fast)} \]
   Which of the following rate laws is consistent with the mechanism?
   A) \(\text{Rate} = k[\text{H}_2\text{O}_2][\text{H}^+][\text{Br}^-]\)
   B) \(\text{Rate} = k[\text{H}_2\text{O}^+\text{--OH}][\text{Br}^-]\)
   C) \(\text{Rate} = k[\text{H}_2\text{O}_2][\text{H}^+][\text{Br}^-]\)
   D) \(\text{Rate} = k[\text{HOBr}][\text{H}^+][\text{Br}^-][\text{H}_2\text{O}_2]\)
   E) \(\text{Rate} = k[\text{Br}^-]\)

17. Which of the following affects the activation energy of a reaction?
   A) temperature of the reactants
   B) concentrations of reactants
   C) presence of a catalyst
   D) surface area of reactants
   E) reaction progress
18. A catalyst accelerates a reaction because

A) it increases the number of molecules with energy equal to or greater than the activation energy.
B) it lowers the activation energy for the reaction.
C) it increases the number of collisions between molecules.
D) it increases the temperature of the molecules in the reaction.
E) it supplies energy to reactant molecules.

19. The gas-phase reaction $\text{CH}_3\text{NC} \rightarrow \text{CH}_3\text{CN}$ has been studied in a closed vessel, and the rate equation was found to be: $\text{Rate} = -\Delta[\text{CH}_3\text{NC}] / \Delta t = k[\text{CH}_3\text{NC}]$. Which one of the following actions is least likely to cause a change in the rate of the reaction?

A) lowering the temperature
B) adding a catalyst
C) using a larger initial amount of CH$_3$NC in the same vessel
D) using a bigger vessel, but the same initial amount of CH$_3$NC
E) continuously removing CH$_3$CN as it is formed

20. For the reaction

$$\text{A(g)} + 2\text{B(g)} \rightarrow 2\text{C(g)} + 2\text{D(g)}$$

the following data were collected at constant temperature. Determine the correct rate law for this reaction.

<table>
<thead>
<tr>
<th>Trial</th>
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<th>Initial [B] (mol/L)</th>
<th>Initial Rate (mol/(L·min))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.125</td>
<td>0.200</td>
<td>7.25</td>
</tr>
<tr>
<td>2</td>
<td>0.375</td>
<td>0.200</td>
<td>21.75</td>
</tr>
<tr>
<td>3</td>
<td>0.250</td>
<td>0.400</td>
<td>14.50</td>
</tr>
<tr>
<td>4</td>
<td>0.375</td>
<td>0.400</td>
<td>21.75</td>
</tr>
</tbody>
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A) Rate = $k[A][B]$
B) Rate = $k[A]^2[B]$
C) Rate = $k[A][B]^2$
D) Rate = $k[A]$
E) Rate = $k[A]^3$