At a certain temperature the following data were collected for the decomposition of HI.

\[ 2\text{HI} \rightarrow \text{H}_2 + \text{I}_2 \]

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Initial [HI] (mol L(^{-1}))</th>
<th>Initial rate of reaction (mol L(^{-1}) s(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>(1.0 \times 10^{-2})</td>
<td>(4.0 \times 10^{-6})</td>
</tr>
<tr>
<td>2</td>
<td>(2.0 \times 10^{-2})</td>
<td>(1.6 \times 10^{-5})</td>
</tr>
<tr>
<td>3</td>
<td>(3.0 \times 10^{-2})</td>
<td>(3.6 \times 10^{-5})</td>
</tr>
</tbody>
</table>

Determine the rate law for the reaction.

**Between experiment (1) and (2), the concentration of HI is doubled. This leads to the rate increasing by a factor of 4.**

**Between experiment (1) and (3), the concentration of HI is trebled. This leads the to rate increasing by a factor 9.**

The rate is proportional to \([\text{HI}]^2\):

\[ \text{rate} = k[\text{HI}]^2 \]

What is the value of the rate constant for the decomposition of HI?

**Using experiment (1), [HI] = 1.0 \times 10^{-2} \text{ mol L}^{-1} \text{ and rate} = 4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ s}^{-1}:**

\[
\text{rate} = k[\text{HI}]^2 \\
4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ s}^{-1} = (k) \times (1.0 \times 10^{-2} \text{ mol L}^{-1})^2 \\
k = \frac{(4.0 \times 10^{-6} \text{ mol L}^{-1} \text{ s}^{-1})}{(1.0 \times 10^{-2} \text{ mol L}^{-1})^2} = 4.0 \times 10^{-2} \text{ mol}^1 \text{ L}^{-1} \text{ s}^{-1}
\]

**Answer:** \(4.0 \times 10^{-2} \text{ mol}^1 \text{ L} \text{ s}^{-1}\)
Four experiments were conducted to discover how the initial rate of consumption of $\text{BrO}_3^-$ ions in the reaction below varied as the concentrations of the reactants were changed.

$$\text{BrO}_3^- + 5\text{Br}^- + 6\text{H}^+ \rightarrow 3\text{Br}_2 + 3\text{H}_2\text{O}$$

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Initial concentration (mol L$^{-1}$)</th>
<th>Initial rate (mol L$^{-1}$ s$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$\text{BrO}_3^-$</td>
<td>$\text{Br}^-$</td>
</tr>
<tr>
<td>1</td>
<td>0.10</td>
<td>0.10</td>
</tr>
<tr>
<td>2</td>
<td>0.20</td>
<td>0.10</td>
</tr>
<tr>
<td>3</td>
<td>0.10</td>
<td>0.30</td>
</tr>
<tr>
<td>4</td>
<td>0.20</td>
<td>0.10</td>
</tr>
</tbody>
</table>

Use the experimental data in the table above to determine the order of the reaction with respect to each reactant.

Between experiments (1) and (3), $[\text{Br}^-]$ and $[\text{H}^+]$ are kept constant but $[\text{BrO}_3^-]$ is doubled. This doubles the rate: the rate is proportional to $[\text{BrO}_3^-]_1$ and so is first order with respect to $\text{BrO}_3^-$.  

Between experiments (2) and (4), $[\text{BrO}_3^-]$ and $[\text{Br}^-]$ are kept constant but $[\text{H}^+]$ is increased by a factor of $(0.15/0.10) = 1.5$. This increases the rate by a factor of $(5.4 \times 10^{-3} / 2.4 \times 10^{-3}) = 2.25$: the rate is proportional to $[\text{H}^+]_2$ as $(1.5)^2 = 2.25$ and so is second order with respect to $\text{H}^+$.  

Between experiments (1) and (2), $[\text{BrO}_3^-]$ and $[\text{H}^+]$ are kept constant but $[\text{Br}^-]$ is increased by a factor of 3. This increases the rate by a factor of $(3.5 \times 10^{-3} / 1.2 \times 10^{-3}) = 2.9$: the rate is proportional to $[\text{Br}^-]_1$ and so is first order with respect to $\text{Br}^-$.  

Overall,  

$$\text{rate} = k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]_2$$

What is the rate of formation of $\text{Br}_2$ when $[\text{BrO}_3^-] = [\text{Br}^-] = [\text{H}^+] = 0.10$ M?

From the table, when $[\text{BrO}_3^-] = [\text{Br}^-] = [\text{H}^+] = 0.10$ M, the rate of consumption of $\text{BrO}_3^-$ is $1.2 \times 10^{-3}$ M s$^{-1}$. From the chemical equation, $\text{Br}_2$ is produced at three times this rate.

The rate of production of $\text{Br}_2$ is $3.6 \times 10^{-3}$ M s$^{-1}$.

Write the rate law for the reaction and determine the value of the rate constant, $k$.

From above, $\text{rate} = k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]_2$. Using experiment (1):
rate \ = \ k[\text{BrO}_3^-][\text{Br}^-][H^+]^2 \\
\quad \quad = k(0.10 \text{ M})(0.10 \text{ M})(0.10 \text{ M})^2 = k(0.00010 \text{ M}^4) = 1.2 \times 10^{-3} \text{ M s}^{-1}

So,

\[ k = \frac{(1.2 \times 10^{-3} \text{ M s}^{-1})}{(0.00010 \text{ M}^3)} = 12 \text{ M}^{-3} \text{ s}^{-1} \]
Nitrogen monoxide, a noxious pollutant, reacts with oxygen to produce nitrogen dioxide, another toxic gas:

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

The following rate data were collected at 225 °C.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>[NO]₀ (M)</th>
<th>[O₂]₀ (M)</th>
<th>Initial rate, –d[O₂]/dt, (M s⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1.3 × 10⁻²</td>
<td>1.1 × 10⁻²</td>
<td>1.6 × 10⁻³</td>
</tr>
<tr>
<td>2</td>
<td>1.3 × 10⁻²</td>
<td>2.2 × 10⁻²</td>
<td>3.2 × 10⁻³</td>
</tr>
<tr>
<td>3</td>
<td>2.6 × 10⁻²</td>
<td>1.1 × 10⁻²</td>
<td>6.4 × 10⁻³</td>
</tr>
</tbody>
</table>

Determine the rate law for the reaction.

Between experiments 1 and 2, [NO] is held constant and [O₂] doubles. This leads to a doubling of the rate: the reaction is 1st order with respect to O₂.

Between experiments 1 and 3, [O₂] is held constant and [NO] doubles. This leads to the rate increasing by a factor of 4: the rate is 2nd order with respect to NO.

The rate law is therefore:

\[ –\frac{d[\text{O}_2]}{dt} = k[\text{NO}]^2[\text{O}_2] \]

Calculate the value of the rate constant at 225 °C.

In experiment 1, [NO] = 1.3 × 10⁻² M, [O₂] = 1.1 × 10⁻² M and rate = 1.6 × 10⁻³ M s⁻¹.

Substituting these values into the rate law gives:

\[(1.6 \times 10^{-3} \text{ M s}^{-1}) = k \times (1.3 \times 10^{-2} \text{ M})^2 \times (1.1 \times 10^{-2} \text{ M})\]

Hence:

\[ k = 860 \text{ M}^{-2} \text{ s}^{-1} \]

Answer: 860 M⁻² s⁻¹

Calculate the rate of appearance of NO₂ when [NO] = [O₂] = 6.5 × 10⁻³ M.

Substituting the values into the rate law gives:

\[ –\frac{d[\text{O}_2]}{dt} = k[\text{NO}]^2[\text{O}_2] \]

\[ = (860 \text{ M}^{-2} \text{ s}^{-1}) \times (6.5 \times 10^{-3} \text{ M})^2 \times (6.5 \times 10^{-3} \text{ M}) = 2.35 \times 10^{-4} \text{ M s}^{-1} \]

From the chemical equation, the rate of appearance of NO₂ is twice the rate of loss of O₂:

\[ -\frac{d[\text{NO}_2]}{dt} = 2 \times -\frac{d[\text{O}_2]}{dt} = (2 \times 2.35 \times 10^{-4} \text{ M s}^{-1}) = 4.7 \times 10^{-4} \text{ M s}^{-1} \]

Answer: 4.7 × 10⁻⁴ M s⁻¹

Answer continues on the next page
Suggest a possible mechanism for the reaction based on the form of the rate law. Explain your answer.

A possible mechanism is:

Step 1: \( \text{NO} + \text{NO} \rightleftharpoons \text{N}_2\text{O}_2 \) fast equilibrium

Step 2: \( \text{N}_2\text{O}_2 + \text{O}_2 \rightarrow 2\text{NO}_2 \) slow (i.e. rate determining)

If the first step is at equilibrium with equilibrium constant \( K_1 \):

\[
K_1 = \frac{[\text{N}_2\text{O}_2]}{[\text{NO}]^2} \Rightarrow [\text{N}_2\text{O}_2] = K_1[\text{NO}]^2
\]

The rate of step 2 is therefore

\[
\text{rate} = k_2[\text{N}_2\text{O}_2][\text{O}_2]
\]

\[
= k_2K_1[\text{NO}]^2[\text{O}_2]
\]

This is consistent with the experiment rate law with \( k = k_1K \).
2-Bromo-2-methylpropane reacts with hydroxide ions to give 2-methyl-2-propanol.

\[(\text{CH}_3)_3\text{CBr} + \text{OH}^- \rightarrow (\text{CH}_3)_3\text{COH}\]

The following rate data were collected at 55 °C.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>([\text{(CH}_3)_3\text{CBr}]_0) (M)</th>
<th>([\text{OH}^-]_0) (M)</th>
<th>Initial rate ((\text{d}[\text{(CH}_3)_3\text{COH}]/\text{dt}, \text{M s}^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.050</td>
<td>0.10</td>
<td>(5.0 \times 10^{-4})</td>
</tr>
<tr>
<td>2</td>
<td>0.20</td>
<td>0.10</td>
<td>(2.0 \times 10^{-3})</td>
</tr>
<tr>
<td>3</td>
<td>0.20</td>
<td>0.30</td>
<td>(2.0 \times 10^{-3})</td>
</tr>
</tbody>
</table>

Determine the rate law for the reaction.

Between experiments (1) and (2), \([\text{OH}^-]_0\) is constant. \([\text{(CH}_3)_3\text{CBr}]_0\) is increased by a factor of 4 and this leads to the rate increasing by a factor of 4: rate \(\alpha\) \([\text{(CH}_3)_3\text{CBr}]\).

Between experiments (2) and (3), \([\text{CH}_3)_3\text{CBr}]_0\) is constant. \([\text{OH}^-]_0\) is increased by a factor of 3 and this leads to no change in the rate. The rate is independent of \([\text{OH}^-]_0\).

Overall, rate = \(k[\text{(CH}_3)_3\text{CBr}]\).

Calculate the value of the rate constant at 55 °C.

From the rate law, \(k = \text{rate} / [\text{(CH}_3)_3\text{CBr}]\).

For experiment 1, rate = \(5 \times 10^{-4}\) M s\(^{-1}\) and \([\text{(CH}_3)_3\text{CBr}] = 0.050\) M and so

\[
k = (5 \times 10^{-4} \text{ M s}^{-1}) / (0.050 \text{ M}) = 1.0 \times 10^{-2} \text{ s}^{-1}
\]

Answer: \(k = 1.0 \times 10^{-2} \text{ s}^{-1}\)

Suggest a possible mechanism for the reaction based on the form of the rate law. Explain your answer.

The rate is independent of \([\text{OH}^-]\) suggesting that it is involved in a step after the rate determining step.

The rate is directly proportional to \([\text{(CH}_3)_3\text{CBr}]\) suggesting that a rate determining step which is unimolecular in \((\text{CH}_3)_3\text{CBr}\).

A possible mechanism is:

1. \((\text{CH}_3)_3\text{Br} \rightarrow (\text{CH}_3)_3\text{}^+ + \text{Br}^-\) slow
2. \((\text{CH}_3)_3\text{}^+ + \text{OH}^- \rightarrow (\text{CH}_3)_3\text{OH}\) fast
The reaction is exothermic. Draw the potential energy vs reaction coordinate diagram for this mechanism, labelling all species that can be isolated.

From 2008-N-6, the mechanism involves two steps: a slow first step and a fast second step. As the first step is rate determining, it has a higher activation energy. As the reaction is exothermic overall, the products have lower enthalpy than the reactants.
Nitric oxide, a noxious pollutant, and hydrogen react to give nitrous oxide and water according to the following equation.

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

The following rate data were collected at 225 °C.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>[NO]₀ (M)</th>
<th>[H₂]₀ (M)</th>
<th>Initial rate (d[NO]/dt, M s⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>6.4 × 10⁻³</td>
<td>2.2 × 10⁻³</td>
<td>2.6 × 10⁻⁵</td>
</tr>
<tr>
<td>2</td>
<td>1.3 × 10⁻²</td>
<td>2.2 × 10⁻³</td>
<td>1.0 × 10⁻⁴</td>
</tr>
<tr>
<td>3</td>
<td>6.4 × 10⁻³</td>
<td>4.4 × 10⁻³</td>
<td>5.1 × 10⁻⁵</td>
</tr>
</tbody>
</table>

Determine the rate law for the reaction.

Between experiments (1) and (2), [H₂]₀ is constant and [NO]₀ doubles. As the rate increases by a factor of \( \frac{1.0 \times 10^{-4}}{2.6 \times 10^{-5}} = 3.8 \sim 4 \), the rate is second order with respect to [NO]₀.

Between experiments (1) and (3), [H₂]₀ doubles and [NO]₀ is constant. As the rate increases by a factor of \( \frac{5.1 \times 10^{-5}}{2.6 \times 10^{-5}} = 2.0 \), the rate is first order with respect to [NO]₀.

Overall,

rate = \( k[\text{NO}]^2[\text{H}_2] \)

Calculate the value of the rate constant at 225 °C.

Using experiment (1),

\[ 2.6 \times 10^{-5} \text{ M s}^{-1} = k \times (6.4 \times 10^{-3} \text{ M})^2 \times (2.2 \times 10^{-3} \text{ M}) \]

\[ k = 2.9 \times 10^2 \text{ M}^2 \text{ s}^{-1} \]

Answer: \( k = 2.9 \times 10^2 \text{ M}^2 \text{ s}^{-1} \)

Calculate the rate of appearance of N₂O when [NO] = [H₂] = 6.6 × 10⁻³ M.

rate of disappearance of NO = \( k[\text{NO}]^2[\text{H}_2] \)

\[ = (2.9 \times 10^2 \text{ M}^2 \text{ s}^{-1}) \times (6.6 \times 10^{-3} \text{ M})^2 \times (6.6 \times 10^{-3} \text{ M}) \]

\[ = 8.3 \times 10^{-5} \text{ M s}^{-1} \]

The rate of appearance of N₂O is half this value as, from the chemical equation, NO is disappearing at twice the rate than N₂O is appearing.

Answer: \( 4.1 \times 10^{-5} \text{ M s}^{-1} \)

Answer continues on the next page.
Suggest a possible mechanism for the reaction based on the form of the rate law. Explain your answer.

Step 1: \[2\text{NO}(g) \rightleftharpoons \text{N}_2\text{O}_2(g)\]

This is a fast equilibrium and so \(K = \frac{[\text{N}_2\text{O}_2(g)]}{[\text{NO}(g)]^2}\) or \([\text{N}_2\text{O}_2(g)] = K[\text{NO}(g)]^2\)

Step 2: \(\text{N}_2\text{O}_2(g) + \text{H}_2(g) \rightarrow \text{N}_2\text{O}(g) + \text{H}_2\text{O}(g)\) slow (i.e. rate determining)

As this is rate determining,

\[\text{rate} = k_2[\text{N}_2\text{O}_2][\text{H}_2]\]

As \([\text{N}_2\text{O}_2(g)] = K[\text{NO}(g)]^2\), this can be rewritten as,

\[\text{rate} = kK[\text{NO}]^2[\text{H}_2]\]

This is consistent with the experimentally determined rate law with \(k_{\text{exp}} = k_2K\).
The major pollutants NO(g), CO(g), NO₂(g) and CO₂(g) are emitted by cars and can react according to the following equation.

\[ \text{NO}_2(g) + \text{CO}(g) \rightarrow \text{NO}(g) + \text{CO}_2(g) \]

The following rate data were collected at 225 °C.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>([\text{NO}_2]_0) (M)</th>
<th>([\text{CO}]_0) (M)</th>
<th>Initial rate ((d[\text{NO}_2]/dt), M s(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.263</td>
<td>0.826</td>
<td>(1.44 \times 10^{-5})</td>
</tr>
<tr>
<td>2</td>
<td>0.263</td>
<td>0.413</td>
<td>(1.44 \times 10^{-5})</td>
</tr>
<tr>
<td>3</td>
<td>0.526</td>
<td>0.413</td>
<td>(5.76 \times 10^{-5})</td>
</tr>
</tbody>
</table>

Determine the rate law for the reaction.

Between experiments (1) and (2), \([\text{NO}_2]_0\) is constant and \([\text{CO}]_0\) is halved. The rate does not change. The rate is independent of \([\text{CO}]\): zero order with respect to \([\text{CO}]\).

Between experiments (2) and (3), \([\text{CO}]_0\) is kept constant and \([\text{NO}_2]_0\) is doubled. The rate increases by a factor of four: the rate is second order with respect to \([\text{NO}_2]\).

Overall,

\[ \text{rate} \propto [\text{NO}_2]^2 = k[\text{NO}_2]^2 \]

Calculate the value of the rate constant at 225 °C.

In experiment (1), rate = \(1.44 \times 10^{-5}\) M s\(^{-1}\) when \([\text{NO}_2] = 0.263\) M. Using the rate law:

\[ 1.44 \times 10^{-5} = k \times (0.263)^2 \quad \text{so} \quad k = 2.08 \times 10^{-4} \]

The units of \(k\) can be deduced from the rate law:

\[ \text{rate} = k[\text{NO}_2]^2 \]

\[ \text{M s}^{-1} = (\text{units of } k) \times (\text{M})^2 \quad \text{so} \quad k \text{ must have units of } \text{“M}^{-1} \text{s}^{-1}” \]

Answer: \(2.08 \times 10^{-4}\) M\(^{-1}\) s\(^{-1}\)
Calculate the rate of appearance of CO₂ when [NO₂] = [CO] = 0.500 M.

When [NO₂] = 0.500 M, rate = \( \frac{d[NO₂]}{dt} = (2.08 \times 10^{-4}) \times (0.500)^2 = 5.20 \times 10^{-5} \text{ M s}^{-1} \)

From the chemical equation, one mole of CO₂ is produced for every mole of NO₂ that is removed. Thus, rate of appearance of CO₂ = rate of loss of NO₂.

Answer: \( 5.20 \times 10^{-5} \text{ M s}^{-1} \)

Suggest a possible mechanism for the reaction based on the form of the rate law. Explain your answer.

A possible mechanism is:

\[
\begin{align*}
\text{NO}_2(g) + \text{NO}_2(g) & \rightarrow \text{NO}(g) + \text{NO}_3(g) \quad \text{(slow)} \\
\text{NO}_3(g) + \text{CO}(g) & \rightarrow \text{NO}_2(g) + \text{CO}_2(g) \quad \text{(fast)}
\end{align*}
\]

The first step is slow and is rate determining. For this step, rate \( \propto [\text{NO}_2]^2 \), as observed. The second step is fast and does not contribute to the overall rate of the reaction and so the rate is independent of [CO(g)].
2-Propanol can be oxidised to acetone using Cr\(_{2}\)O\(_7^{2-}\) in acidic solution as indicated in the reaction below. The rate of decrease of the Cr\(_{2}\)O\(_7^{2-}\) ion under a certain set of conditions is 3.0 mol L\(^{-1}\) s\(^{-1}\).

\[
3\text{CH}_3\text{CH(OH)CH}_3 + \text{Cr}_2\text{O}_7^{2-} + 8\text{H}^+ \rightarrow 3\text{CH}_3\text{COCH}_3 + 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

What is the rate of increase in the concentration of Cr\(^{3+}\)?

6.0 mol L\(^{-1}\) s\(^{-1}\)

What is the rate of decrease in the concentration of 2-propanol?

9.0 mol L\(^{-1}\) s\(^{-1}\)

The rate law for this reaction is:  
\[
\text{Rate} = k [(\text{Cr}_2\text{O}_7^{2-})[\text{CH}_3\text{CH(OH)CH}_3][\text{H}^+]^2
\]

Complete the following table by writing increase, decrease or no change in the box to indicate how the rate of the reaction is affected by each of the following changes.

<table>
<thead>
<tr>
<th>Change</th>
<th>Rate Effect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Increase in [CH(_3)CH(OH)CH(_3)]</td>
<td>increase</td>
</tr>
<tr>
<td>Increase in [CH(_3)COCH(_3)]</td>
<td>no change</td>
</tr>
<tr>
<td>Increase in pH</td>
<td>decrease</td>
</tr>
<tr>
<td>Increase in temperature</td>
<td>increase</td>
</tr>
</tbody>
</table>

Complete the following table.

<table>
<thead>
<tr>
<th>Formula</th>
<th>Systematic name</th>
<th>Oxidation state of transition metal</th>
<th>Number of d-electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>K(_2)[Pt(CN)(_4)]</td>
<td>potassium tetracyanidoplatinate(II)</td>
<td>+2</td>
<td>8</td>
</tr>
<tr>
<td>[Co(H(_2)O)(_6)]Cl(_2)</td>
<td>hexaquacobalt(II) chloride</td>
<td>+2</td>
<td>7</td>
</tr>
</tbody>
</table>
• Explain in terms of their electronic configurations and ionisation energies why the alkali metals (Group 1) are powerful reducing agents.

Ionisation energies increase across a period in the periodic table because the increasing nuclear charge holds the electrons more tightly. Hence, in any period, the Group I element is the one that most easily loses its electron (from the s subshell). This electron is then available to reduce another species.

• The half-life for the first order decomposition of N\textsubscript{2}O\textsubscript{5}(g) is 6.00 × 10\textsuperscript{4} s at 20 °C. Calculate the rate constant, k, at this temperature.

For a first-order reaction, the half-life, \( t_{1/2} \), is related to the rate constant, k, by:

\[
\frac{\ln 2}{k} = \frac{\ln 2}{(6.00 \times 10^4 \text{ s}^{-1})} = 1.16 \times 10^{-5} \text{ s}^{-1}
\]

\[
k = 1.16 \times 10^{-5} \text{ s}^{-1}
\]

What percentage of the N\textsubscript{2}O\textsubscript{5} molecules will have reacted after one hour?

For a first-order reaction:

\[
\ln[N\textsubscript{2}O\textsubscript{5}] = \ln[N\textsubscript{2}O\textsubscript{5}]_0 - kt \quad \text{or} \quad \frac{\ln[N\textsubscript{2}O\textsubscript{5}]}{\ln[N\textsubscript{2}O\textsubscript{5}]_0} = -kt
\]

Hence for \( t = 1 \text{ hour} = (60 \times 60) \text{ s} = 3600 \text{ s} \),

\[
\frac{\ln[N\textsubscript{2}O\textsubscript{5}]}{\ln[N\textsubscript{2}O\textsubscript{5}]_0} = -(1.16 \times 10^{-5} \text{ s}^{-1}) \times (3600 \text{ s}) = 0.0418
\]

\[
\frac{[N\textsubscript{2}O\textsubscript{5}]}{[N\textsubscript{2}O\textsubscript{5}]_0} = 0.959 \text{ or } 95.9\% \text{ remains.}
\]

The amount that has reacted is \( (100.0 - 95.9) = 4.1\% \)

ANSWER: 4.1%