The reaction \(2 \text{NOCl} \rightarrow 2 \text{NO} + \text{Cl}_2\) exhibits the differential rate law: rate = \(k[\text{NOCl}]^2\). In an experiment, 0.100 M NOCl is present initially. After 12 minutes, 0.085 M remains.

A) Determine \(k\) for this reaction.

\[
\frac{1}{0.085 \text{M}} = k \left( \frac{12 \text{ min}}{0.100 \text{ M}} \right) + \frac{1}{0.100 \text{ M}}
\]

\[
0.85 = k \times 0.12 + 1 = 0.147 \text{ M}^{-1} \text{min}^{-1}
\]

B) Determine \(t_{1/2}\) for this reaction.

\[
\frac{1}{\frac{1}{2} [\text{NOCl}]_0} = (0.147 \text{ M}^{-1} \text{min}^{-1}) t_{1/2} + \frac{1}{[\text{NOCl}]_0}
\]

\[
\frac{1}{0.100 \text{ M}} = (0.147 \text{ M}^{-1} \text{min}^{-1}) t_{1/2}
\]

\[
t_{1/2} = 68.0 \text{ min} = 4080 \text{ s}
\]

For a given reaction, \(A + B \rightarrow C\), the rate constant triples when the temperature is increased from 25°C to 45°C. What is the activation energy for this reaction?

\[
\frac{k_2}{k_1} = 3, \quad T_a = 318 \text{ K}, \quad T_f = 298 \text{ K}
\]

\[
\ln \left( \frac{3k_2}{k_1} \right) = - \frac{E_a}{8.3145 \times 10^{-5} \text{ K} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}} \left( \frac{1}{318 \text{ K}} - \frac{1}{298 \text{ K}} \right)
\]

\[
E_a = 43 \text{ kJ/mol}
\]

For the reaction: \(2 \text{B} \rightarrow \text{products}\), a plot of \(\ln \left( \frac{[\text{B}]_0}{[\text{B}]} \right)\) versus time yields a straight-line graph. In a separate experiment, the half-life is determined to be 3.2 hours. If you start with a sample that contains 3.87 g of B, how long will it take to reduce 0.25 g of B to 0.001 g?

\[
k = \frac{0.693}{t_{1/2}} = \frac{0.693}{3.2 \text{ h}} = 0.217 \text{ h}^{-1}
\]

\[
\ln \left( \frac{[\text{B}]}{[\text{B}]_0} \right) = -kt
\]

\[
\ln \left( \frac{0.25 \text{ g}}{3.87 \text{ g}} \right) = -(0.217 \text{ h}^{-1}) t
\]

\[
t = 12.6 \text{ h} = 13 \text{ h}
\]
Consider the proposed reaction mechanism for phosphorus monoxide and hydrogen gas:

\[
\begin{align*}
2 \text{PO} & \rightleftharpoons \text{P}_2\text{O}_3 \\
\text{P}_2\text{O}_3 + \text{H}_2 & \rightarrow \text{P}_4 \text{O}_6 + \text{H}_2\text{O} \\
\text{P}_2\text{O}_3 + \text{H}_2 & \rightarrow \text{P}_2 + \text{H}_2\text{O}
\end{align*}
\]

run 1: fast, equilibrium
run 2: slow
run 3: fast

A) Write the balanced equation for the overall reaction.

\[2 \text{PO} + 2 \text{H}_2 \rightarrow \text{P}_2 + 2 \text{H}_2\text{O}\]

B) Which species serve as intermediates in the reaction mechanism?

\[\text{P}_2\text{O}_3 \text{ and } \text{P}_2\text{O}\]

C) Write the differential rate law expression, excluding any intermediates.

\[
\text{rate} = k_2 [\text{P}_2\text{O}_3] [\text{H}_2]
\]

\[k_1 [\text{PO}]^2 = k_7 [\text{P}_2\text{O}_3]
\]

\[[\text{P}_2\text{O}_3] = \frac{k_1}{k_7} [\text{PO}]^2
\]

\[\text{rate} = k_2 \frac{k_1}{k_7} [\text{PO}]^2 [\text{H}_2]
\]

\[\text{rate} = k [\text{PO}]^2 [\text{H}_2]
\]

Sketch in detail the Reaction-Energy Diagram for the endothermic overall reaction \(A+B \rightarrow 2C+D\). Assume that the reaction mechanism for this process involves three elementary steps, and the second step is the slow step (rate-determining step).

- Label the axes.
- Label the relative energies of the reactants, products, transition state(s), and intermediate(s).
- Clearly indicate the Activation energy (E_a).
- Clearly indicate ΔE for the reaction.