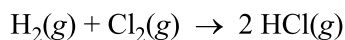


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Question 6
(8 points)



The table below gives data for a reaction rate study of the reaction represented above.

Experiment	Initial [H ₂] (mol L ⁻¹)	Initial [Cl ₂] (mol L ⁻¹)	Initial Rate of Formation of HCl (mol L ⁻¹ s ⁻¹)
1	0.00100	0.000500	1.82 × 10 ⁻¹²
2	0.00200	0.000500	3.64 × 10 ⁻¹²
3	0.00200	0.000250	1.82 × 10 ⁻¹²

(a) Determine the order of the reaction with respect to H₂ and justify your answer.

<p>The order of the reaction with respect to H₂ is 1. Comparing experiments 1 and 2, doubling the initial concentration of H₂ while keeping the initial concentration of Cl₂ constant results in a doubling of the reaction rate.</p>	<p>One point is earned for the correct order with justification.</p>
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(b) Determine the order of the reaction with respect to Cl₂ and justify your answer.

<p>The order of the reaction with respect to Cl₂ is 1. Comparing experiments 2 and 3, halving the initial concentration of Cl₂ while keeping the initial concentration of H₂ constant results in a halving of the reaction rate.</p>	<p>One point is earned for the correct order with justification.</p>
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(c) Write the overall rate law for the reaction.

$\text{rate} = k [\text{H}_2][\text{Cl}_2]$	<p>One point is earned for a rate law consistent with part (a) and part (b).</p>
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(d) Write the units of the rate constant.

$k = \frac{\text{rate}}{[\text{H}_2][\text{Cl}_2]} = \frac{\text{mol L}^{-1} \text{ s}^{-1}}{\text{mol L}^{-1} \text{ mol L}^{-1}}$ $= \frac{\text{s}^{-1}}{\text{mol L}^{-1}} = \text{L mol}^{-1} \text{ s}^{-1}$	<p>One point is earned for units consistent with part (c).</p>
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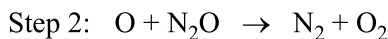
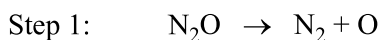
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Question 6 (continued)

- (e) Predict the initial rate of the reaction if the initial concentration of H_2 is $0.00300 \text{ mol L}^{-1}$ and the initial concentration of Cl_2 is $0.000500 \text{ mol L}^{-1}$.

For this reaction, the initial concentration of Cl_2 is the same as in Experiment 1 but the initial concentration of H_2 is three times as large. And because the reaction is first order with respect to each reactant, the initial rate of the reaction would be $5.46 \times 10^{-12} \text{ mol L}^{-1} \text{ s}^{-1}$, which is three times the rate of the initial rate of the reaction in Experiment 1.	One point is earned for the correct numerical answer or correct multiplier consistent with the rate law from part (c).
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The gas-phase decomposition of nitrous oxide has the following two-step mechanism.



- (f) Write the balanced equation for the overall reaction.

$2 \text{N}_2\text{O} \rightarrow 2 \text{N}_2 + \text{O}_2$	One point is earned for the correct balanced equation.
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- (g) Is the oxygen atom, O, a catalyst for the reaction or is it an intermediate? Explain.

The O atom is an intermediate because it is formed and then consumed during the course of the reaction. (Had it been a catalyst, it would have been present both at the beginning and the end of the reaction.)	One point is earned for the correct choice with explanation.
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- (h) Identify the slower step in the mechanism if the rate law for the reaction was determined to be $\text{rate} = k [\text{N}_2\text{O}]$. Justify your answer.

Step 1 is slower because N_2O appears in Step 1 as the single reactant, which is consistent with the given rate law.	One point is earned for the correct choice with justification.
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