

Kinetics

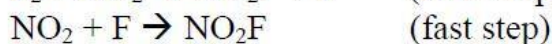
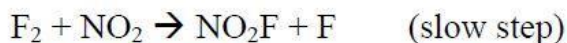
5.4 Elementary Reactions

5.7 Introduction to Reaction Mechanisms

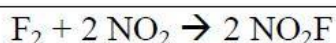
5.8 Reaction Mechanism and Rate Law

Worksheet Key

1) A two step mechanism is outlined below.



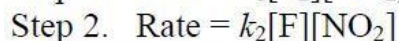
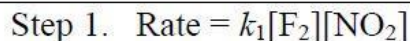
a. What is the overall reaction?



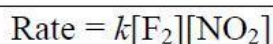
b. Identify the intermediate in this reaction mechanism. Justify your answer.

F is an intermediate, as it is produced and consumed in the reaction.

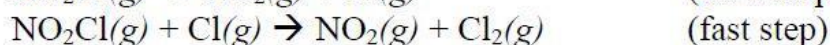
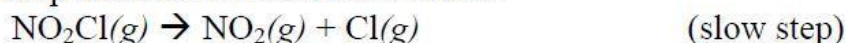
c. What is the rate law for each of the elementary steps?



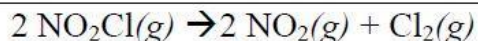
d. What is the rate law for the overall reaction?



2) A two step mechanism is outlined below.



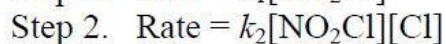
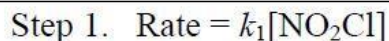
a. What is the overall reaction?



b. Identify the intermediate in this reaction mechanism. Justify your answer.

Cl(g) is an intermediate, as it is produced and consumed in the reaction.

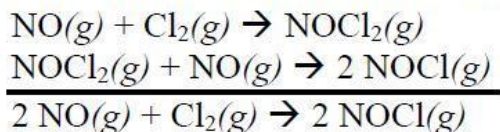
c. What is the rate law for each of the elementary steps?



d. What is the rate law for the overall reaction?

$$\text{Rate} = k_1[\text{NO}_2\text{Cl}]$$

3) The rate law for the overall reaction outlined below is: $\text{Rate} = k[\text{NO}][\text{Cl}_2]$.



a. Identify the intermediate in this reaction mechanism. Justify your answer.

NOCl_2 is an intermediate, as it is produced and consumed in the reaction.

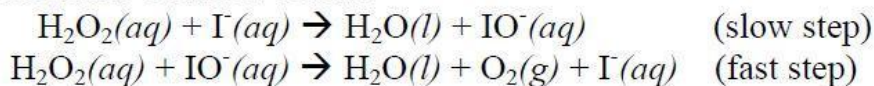
b. Which step occurs the fastest? Justify your answer.

Step 2. $\text{NOCl}_2(g) + \text{NO}(g) \rightarrow 2 \text{NOCl}(g)$ is the fast step, because it does not determine the rate of the overall reaction. The rate law for the overall reaction is the same as the rate law for the first step, which means that the first step is the slow, rate determining, step.

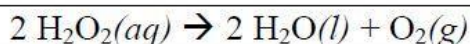
c. Which step determines the overall rate of the reaction? Explain.

Step 1. $\text{NO}(g) + \text{Cl}_2(g) \rightarrow \text{NOCl}_2(g)$ determines the rate of the overall reaction. We know this because the rate law of the overall reaction is: $\text{Rate} = k[\text{NO}][\text{Cl}_2]$. This is also the rate law for this step. Step 1 is the slow step, which determines the rate of the overall reaction.

4) A two step reaction is outlined below.



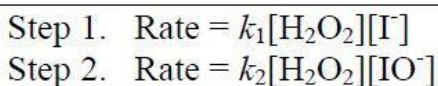
a. What is the overall reaction?



b. Identify the intermediate in this reaction mechanism. Justify your answer.

$\text{IO}^-(aq)$ is an intermediate, as it is produced and consumed in the reaction.

c. What is the rate law for each of the elementary steps?



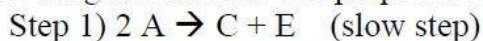
d. What is the rate law for the overall reaction?

$$\text{Rate} = k[\text{H}_2\text{O}_2][\text{I}^-]$$

5) The reaction below was found to be second order for A and zero order for B.



The following mechanism was proposed for this reaction.



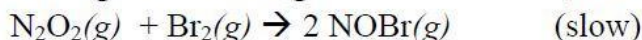
a. What is the rate law for this reaction?

$$\text{Rate} = k[\text{A}]^2$$

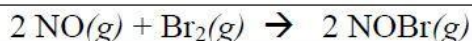
b. Give two reasons why the above mechanism is possible.

1. When the two reactions are added together, we get $\text{A} + \text{B} \rightarrow \text{C} + \text{D}$, which is the same as the overall reaction.
2. The slow step is consistent with the rate law, as B is zero order and thus is not included in the rate law; and A is second order and raised to the power of 2 in the rate law, which means that two A particles collide in the rate determining step.

6) A two step mechanism is outlined below.



a. What is the overall reaction?



b. Identify the intermediate in this reaction mechanism. Justify your answer.

$\text{N}_2\text{O}_2(\text{g})$ is an intermediate, as it is produced and consumed in the reaction.

c. What is the rate law for the overall reaction?

$$k_1[\text{NO}]^2 = k_{-1}[\text{N}_2\text{O}_2]$$

$$[\text{N}_2\text{O}_2] = \frac{k_1}{k_{-1}}[\text{NO}]^2$$

$$\text{Rate} = k_2[\text{N}_2\text{O}_2][\text{Br}_2]$$

$$\text{Rate} = k_2 \left(\frac{k_1}{k_{-1}}[\text{NO}]^2 \right) [\text{Br}_2]$$

$$\text{Rate} = k_3[\text{NO}]^2[\text{Br}_2]$$

- d. What is the order of the reaction with respect to NO according to the above mechanism? Justify your answer.

The reaction is second order with respect to NO, as the [NO] is raised to the power of 2 in the rate law.

- 7) Elementary steps that involve three or more reacting species are very rare. Explain why this is.

All *termolecular* elementary steps are very rare, as they require the simultaneous collision of three particles with correct orientations and enough energy. Elementary steps that require the simultaneous collision of more than three reactant particles are very rare for the same reason.