

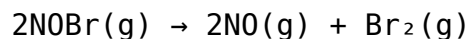
The Rate Law For The Reaction: $2\text{NOBr(g)} \rightarrow 2\text{NO(g)} + \text{Br}_2(\text{g})$ Is Rate $=k[\text{NOBr}]^2$ where The Rate Constant Is 0.55

The Rate Law For The Reaction: $2\text{NOBr(g)} \rightarrow 2\text{NO(g)} + \text{Br}_2(\text{g})$ Is Rate $=k[\text{NOBr}]^2$ where The Rate Constant Is 0.55 is a fundamental expression that describes how the speed of this chemical reaction depends on the concentration of the reactant NOBr. Understanding the rate law is essential for chemists to predict how quickly a reaction will proceed under various conditions, optimize industrial processes, and understand the underlying mechanisms at the molecular level. This article delves into the details of this specific reaction, exploring its kinetics, the meaning of the rate law, how the rate constant influences the reaction, and the broader implications in chemical science.

Understanding the Reaction and Its Rate Law

Reaction Overview

The chemical reaction under consideration is:



This is a decomposition reaction where two nitrogen oxybromide (NOBr) molecules break down to form nitric oxide (NO) and bromine (Br₂). Such reactions are significant in various industrial and environmental contexts, including nitric oxide production and bromine-based processes.

What Is a Rate Law?

The rate law expresses how the rate of a chemical reaction depends on the concentration(s) of reactant(s). It is generally written as:

$$\text{Rate} = k [\text{A}]^m [\text{B}]^n$$

where:

- k is the rate constant,
- $[\text{A}]$, $[\text{B}]$ are concentrations of reactants,
- m , n are the reaction orders with respect to each reactant.

In the case of our reaction, the rate law simplifies to:

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

indicating that the reaction is second order with respect to NOBr.

Details of the Rate Law for the NOBr Reaction

Reaction Order and Its Significance

The overall reaction order is determined by summing the exponents in the rate law. For this reaction:

- The order with respect to NOBr is 2,
- The total order is also 2.

This suggests that the reaction rate is proportional to the square of NOBr's concentration. As a consequence, doubling NOBr concentration increases the reaction rate by a factor of four, illustrating the quadratic dependence.

The Rate Constant (k) and Its Value

The rate constant, k , is a proportionality factor that links the concentration to the rate. It is temperature-dependent and specific to each reaction at a given temperature. In this case, the rate constant is given as:

$k = 0.55$ (units depend on the overall order; typically $\text{M}^{-1} \text{s}^{-1}$ for a second-order reaction).

This value indicates the intrinsic speed of the reaction under specific conditions (usually at a certain temperature). Higher values of k correspond to faster reactions.

Factors Influencing the Rate Constant

Temperature

The rate constant varies with temperature according to the Arrhenius equation:

$$k = A e^{(-E_a/RT)}$$

where:

- A is the pre-exponential factor,
- E_a is the activation energy,
- R is the gas constant,
- T is temperature in Kelvin.

An increase in temperature generally leads to an increase in k , accelerating the reaction.

Activation Energy

The activation energy (E_a) is the energy barrier that must be overcome for reactants to convert into products. Reactions with lower E_a tend to have higher rate constants at the same temperature.

Catalysts

Catalysts can lower the activation energy, thereby increasing k and speeding up the reaction without being consumed.

Mechanistic Insights Into the NOBr Decomposition

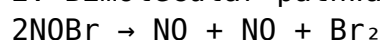
Possible Reaction Mechanisms

Understanding the mechanism provides context for the rate law. For the NOBr decomposition, possible pathways include:

1. Unimolecular decomposition:



2. Bimolecular pathways:



3. Radical chain mechanisms:

involving radical intermediates like NO and Br.

The observed second-order dependence on NOBr indicates that the rate-determining step likely involves two NOBr molecules interacting.

Implications of the Rate Law on Reaction Mechanism

A second-order rate law suggests a bimolecular step. It supports mechanisms where two NOBr molecules collide or interact to form products, possibly via a transition state involving both reactants.

Calculating Reaction Rates and Their Practical Significance

Example Calculation

Suppose the concentration of NOBr at a given moment is 0.1 M. The reaction rate can be calculated as:

$$\text{Rate} = 0.55 \times (0.1)^2 = 0.55 \times 0.01 = 0.0055 \text{ M/s}$$

This indicates that at this concentration, the NOBr decomposes at a rate of 0.0055 mol per liter per second.

Impact of Concentration Changes

Since the rate depends on $[\text{NOBr}]^2$, small changes in concentration significantly affect the reaction rate. For instance, increasing [NOBr] to 0.2 M results in:

$$\text{Rate} = 0.55 \times (0.2)^2 = 0.55 \times 0.04 = 0.022 \text{ M/s}$$

which is four times faster than at 0.1 M.

Industrial and Environmental Relevance

Industrial Applications

Understanding the kinetics of NOBr decomposition is essential in processes like:

- Production of nitric oxide,
- Bromine-related chemical manufacturing,
- Designing reactors with optimal conditions to control reaction rates.

Environmental Considerations

NOBr and its decomposition products are relevant in atmospheric chemistry, affecting air quality and ozone chemistry. Modeling how quickly NOBr decomposes under various conditions helps predict environmental impacts.

Summary and Key Takeaways

- The rate law for the reaction $2\text{NOBr(g)} \rightarrow 2\text{NO(g)} + \text{Br}_2\text{(g)}$ is $\text{Rate} = k[\text{NOBr}]^2$.
- The reaction is second order overall, indicating that two NOBr molecules are involved in the rate-determining step.
- The rate constant, 0.55, reflects how fast the reaction proceeds at a specific temperature.
- Variations in concentration dramatically influence the reaction rate due to

the quadratic dependence.

- Understanding the kinetics informs both industrial process optimization and environmental modeling.

Conclusion

Grasping the rate law of the NOBr decomposition reaction provides valuable insights into its mechanism, speed, and how various factors influence it. Experimental determination of the rate constant, along with temperature control, enables chemists to manipulate reaction conditions effectively. Whether for manufacturing or environmental assessment, the principles underlying the rate law serve as foundational knowledge in chemical kinetics and reaction engineering.

Frequently Asked Questions

What is the overall order of the reaction based on the rate law $\text{Rate} = k[\text{NOBr}]^2$?

The overall order of the reaction is 2, since the rate law shows the rate is proportional to the square of [NOBr].

How does the rate constant (k) influence the reaction rate in this reaction?

The rate constant k (0.55 in this case) directly affects the reaction rate; a higher k results in a faster reaction at the same concentration of NOBr.

What is the significance of the reaction being second order with respect to NOBr?

It indicates that the reaction rate depends on the square of NOBr concentration, meaning the rate will increase quadratically as [NOBr] increases.

How would the rate change if the concentration of NOBr is doubled?

If [NOBr] doubles, the rate would increase by a factor of $2^2 = 4$, since the reaction is second order with respect to NOBr.

Given the rate law and rate constant, how can you

calculate the rate of the reaction at a specific concentration of NOBr?

Use the formula $\text{Rate} = 0.55 \times [\text{NOBr}]^2$, substituting the given concentration of NOBr to find the reaction rate.

Additional Resources

Understanding the Rate Law for the Reaction: $2\text{NOBr(g)} \rightarrow 2\text{NO(g)} + \text{Br}_2\text{(g)}$

The study of chemical kinetics provides invaluable insights into how reactions proceed, their speed, and the factors influencing their rates. One fundamental aspect of this study is the determination and interpretation of the rate law—a mathematical expression that relates the reaction rate to the concentration of reactants. In this comprehensive review, we delve into the specific rate law for the reaction:



with the associated rate law:

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

where the rate constant (k) is given as 0.55 (units depend on the overall order, typically $\text{M}^{-2}\cdot\text{s}^{-1}$).

Introduction to Reaction Kinetics and Rate Laws

Understanding the kinetics of a chemical reaction involves quantifying how quickly reactants are converted into products. The rate law encapsulates this by expressing the reaction rate as a function of the concentrations of reactants, and sometimes products or catalysts.

Key Concepts:

- **Reaction Rate:** The change in concentration of a reactant or product per unit time, usually expressed as molarity per second (M/s).
- **Rate Law:** An equation that relates the reaction rate to the concentrations of reactants raised to specific powers (orders).
- **Rate Constant (k):** A proportionality constant that is temperature-dependent and characteristic of a given reaction at a specific temperature.
- **Reaction Order:** The sum of the exponents of the concentration terms in the rate law; it indicates how the rate responds to concentration changes.

In the case of the reaction under review, the rate law is $\text{Rate} = k[\text{NOBr}]^2$,

indicating a second-order overall reaction, first order in NOBr.

Detailed Analysis of the Given Reaction and Rate Law

Reaction Overview

The reaction:



involves the transformation of nitrogen oxybromide (NOBr) molecules into nitric oxide and bromine molecules. It is a typical example of a halogenation and nitrogen-oxygen compound reaction, often studied in inorganic and physical chemistry.

Key points:

- The stoichiometry indicates that two molecules of NOBr are consumed to produce two molecules of NO and one molecule of Br₂.
- The reaction occurs in the gaseous phase, which implies that kinetic parameters are influenced by gaseous collision processes, temperature, pressure, and concentration.

Interpreting the Rate Law: $\text{Rate} = k[\text{NOBr}]^2$

This rate law signifies:

- The reaction is second order overall, with the entire dependence on NOBr concentration.
- The reaction is first order with respect to NOBr—doubling NOBr concentration doubles the reaction rate.
- The rate constant (k) is 0.55, with units typically expressed as $\text{M}^{-2} \cdot \text{s}^{-1}$ for a second-order overall reaction.

Implications:

- Since the reaction rate depends solely on NOBr concentration, the rate-determining step likely involves NOBr directly or through a mechanism that produces an overall second-order dependence.
- The absence of other reactants in the rate law indicates that either other reactants are in large excess and their concentrations are effectively

constant or that the reaction mechanism involves only NOBr as the reactive species in the rate-determining step.

Mechanistic Insights and Reaction Pathway

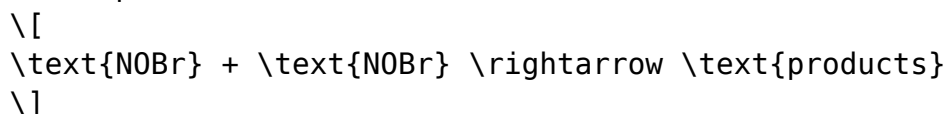
Understanding the mechanism behind this reaction helps explain why the rate law takes this particular form.

Possible Mechanisms Leading to the Rate Law

Given the second-order dependence on NOBr, potential mechanisms could include:

1. Bimolecular Elementary Step:

- Two NOBr molecules collide and react directly, forming products.
- Example:



- This aligns with the observed second-order kinetics, as the rate would depend on the concentrations of both reacting NOBr molecules.

2. Radical Chain Mechanism:

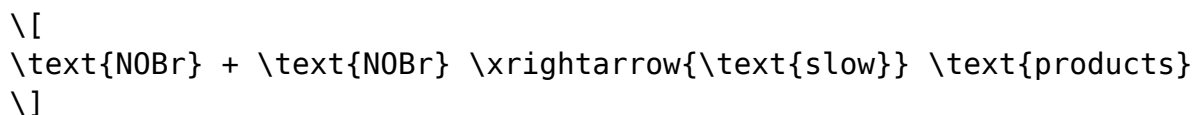
- NOBr dissociates into radicals, which then propagate chain reactions.
- The rate-determining step involves two radicals or molecules derived from radicals combining.

3. Stepwise Mechanism with Intermediate Formation:

- Formation of an intermediate from NOBr, which then reacts further.
- However, the overall rate law still depends on the concentration of NOBr, indicating that NOBr's involvement is direct and in the rate-determining step.

Likely Scenario:

Given the rate law, the reaction probably proceeds via a bimolecular elementary step involving two NOBr molecules, such as:



or a similar bimolecular process.

Role of the Rate Constant (k) and Temperature Dependence

The rate constant $k = 0.55$ (units depending on the overall order) is crucial in defining the speed of the reaction at a given temperature.

Factors affecting k :

- Temperature: According to the Arrhenius equation:

$$k = A e^{-E_a/RT}$$

where:

- A is the pre-exponential factor,
- E_a is the activation energy,
- R is the gas constant,
- T is the temperature in Kelvin.

- Activation Energy (E_a): A higher E_a indicates a slower reaction at a given temperature.

Implication of $k = 0.55$:

- At the temperature at which this rate constant is measured, the reaction proceeds at a rate proportional to the square of NOBr concentration.
- Small changes in temperature can significantly alter the rate due to the exponential dependence in Arrhenius' equation.

Units of k :

- For a second-order reaction, k typically has units of $\text{M}^{-2} \cdot \text{s}^{-1}$, ensuring the rate (M/s) is consistent with the concentration dependence.

Experimental Determination of the Rate Law

The rate law is not always apparent from stoichiometry alone; it must often be deduced experimentally.

Common methods include:

- Initial Rate Method: Measuring the initial reaction rate at various concentrations of NOBr while keeping other conditions constant.
- Method of Varying Concentrations: Systematically changing [NOBr] and observing the impact on reaction rate to determine the reaction order.
- Use of Spectroscopic Techniques: Monitoring concentration changes over time via UV-Vis or IR spectroscopy.

Key observations to determine the order:

- Doubling [NOBr] doubles the rate → First order.
- Quadrupling [NOBr] increases rate fourfold → Second order.

In this reaction, experimental data must have confirmed the quadratic dependence, leading to the rate law $\text{Rate} = k[\text{NOBr}]^2$.

Implications for Reaction Control and Industrial Applications

Understanding the kinetics and rate law allows chemists and engineers to optimize reaction conditions for desired outcomes.

Practical considerations:

- Adjusting Concentrations: To accelerate the reaction, increase NOBr concentration, bearing in mind the quadratic relationship.
- Temperature Control: Raising temperature increases k exponentially, thereby increasing rate; however, safety and stability considerations are paramount.
- Catalysis: Although none is specified here, catalysts could alter the mechanism, change the rate law, or lower activation energy.

Industrial relevance:

- The reaction could be involved in the synthesis of nitric oxide or bromine compounds.
- Efficient process design requires precise kinetic data to predict yields, optimize reaction time, and minimize by-products.

Summary and Final Remarks

Key takeaways:

- The reaction $2\text{NOBr(g)} \rightarrow 2\text{NO(g)} + \text{Br}_2\text{(g)}$ follows a second-order rate law: $\text{Rate} = k[\text{NOBr}]^2$.
- The rate constant (k) at the specified conditions is 0.55 (units consistent with second-order kinetics).
- The reaction mechanism likely involves bimolecular elementary steps involving NOBr molecules.
- The kinetic dependence suggests that doubling NOBr concentration results in a quadrupling of the reaction rate.
- Temperature plays a significant role due to the Arrhenius dependence of the rate constant.
- Experimental methods for determining this rate law involve monitoring initial reaction rates under varying concentrations.

Broader impact:

Understanding this rate law not only deepens comprehension of the specific reaction but also exemplifies fundamental principles in chemical kinetics, including how molecular interactions govern reaction speeds and how kinetic data inform practical applications. Such insight is vital for designing efficient chemical processes, controlling reaction pathways, and developing new chemical technologies.

In conclusion, the detailed analysis of the rate law for the reaction involving NOBr emphasizes the importance of kinetic studies in understanding reaction mechanisms and optimizing conditions for industrial and laboratory applications. The quadratic dependence on NOBr concentration highlights the significance of molecular interactions in dictating reaction speed, and the precise measurement of the rate constant enables accurate modeling and control of this chemical process.

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