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Understanding the dynamics of chemical reactions is fundamental in chemistry, especially when dealing with reaction kinetics. In this comprehensive guide, we will explore the specifics of a first-order reaction, using the example where the rate constant (k) is 5.67 s^{-1}, the initial concentration of reactant A is 0.500 M, and the question centers around determining the time required for the concentration to decrease to 0.125 M. Whether you're a student preparing for exams, a researcher designing experiments, or a professional in the chemical industry, mastering these calculations enhances your grasp of reaction mechanisms and process optimization.

Introduction to First-Order Reactions

What Is a First-Order Reaction?

A first-order reaction is a type of chemical reaction where the rate depends linearly on the concentration of a single reactant. Mathematically, it can be expressed as:

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

where:
- $[A]$ is the concentration of reactant A at any given time,
- k is the rate constant, with units of inverse seconds (s^{-1}).

Characteristics of First-Order Reactions

First-order reactions possess several distinctive features:
- The half-life ($t_{1/2}$) is independent of initial concentration.
- They follow a simple exponential decay pattern.
- The integrated rate law relates concentration and time straightforwardly.

Understanding the Rate Law and Integrated Rate Law

The Rate Law for First-Order Reactions

The rate law for a first-order reaction is:

$$ \frac{d[A]}{dt} = -k[A] $$

This indicates that the rate at which A decreases is proportional to its current concentration.

Integrated Rate Law

By integrating the differential rate law, we arrive at the integrated form:

$$ [A] = [A]_o e^{-kt} $$

where:
- $[A]_o$ is the initial concentration,
- $[A]$ is the concentration at time t,
- e is Euler's number (~2.71828).

This equation allows us to determine the concentration of A at any given time or vice versa.

Calculating the Time to Reach a Specific Concentration

Given Data

- Rate constant, $k = 5.67 \ \text{s}^{-1}$
- Initial concentration, $[A]_o = 0.500 \ \text{M}$
- Target concentration, $[A] = 0.125 \ \text{M}$

Step-by-Step Calculation

To find the time t it takes for the concentration to decrease from $[A]_o$ to $[A]$, we use the integrated rate law:

$$ [A] = [A]_o e^{-kt} $$

Rearranged to solve for t:

$$ t = -\frac{1}{k} \ln \left(\frac{[A]}{[A]_o} \right) $$

Plugging in the values:

$$ t = -\frac{1}{5.67} \ln \left(\frac{0.125}{0.500} \right) $$

Calculate the ratio:

$$ \frac{0.125}{0.500} = 0.25 $$

Compute the natural logarithm:

$$ \ln(0.25) = -1.3863 $$

Finally, calculate t:

$$ t = -\frac{1}{5.67} \times (-1.3863) $$
$$ t = \frac{1.3863}{5.67} $$
$$ t \approx 0.2446 \ \text{seconds} $$

Result: It takes approximately 0.245 seconds for the concentration of A to decrease from 0.500 M to 0.125 M.

Implications and Applications of First-Order Reaction Kinetics

Industrial Reactions and Process Design

Understanding the kinetics of first-order reactions is crucial in designing chemical processes. For instance:
- Reactor sizing depends on the reaction rate and desired conversion.
- Time estimates help optimize reaction conditions to maximize efficiency.

Pharmacokinetics and Medicine

First-order kinetics are common in drug metabolism, where:
- The rate of drug elimination is proportional to its concentration.
- Dosage regimens are designed based on half-life calculations.

Environmental Chemistry

Pollutant degradation often follows first-order kinetics, helping predict:
- The time needed for contaminants to reduce to safe levels.
- The effectiveness of remediation strategies.

Key Points to Remember

- The integrated rate law for a first-order reaction is $[A] = [A]_o e^{-kt}$.
- The time to reach a certain concentration can be calculated using:

$$t = -\frac{1}{k} \ln \left(\frac{[A]}{[A]_o} \right)$$

- A higher rate constant k implies a faster reaction.
- The half-life of a first-order reaction is:

$$t_{1/2} = \frac{\ln 2}{k}$$

which, for $k = 5.67 \ \text{s}^{-1}$, is approximately 0.122 seconds, indicating rapid reaction kinetics.

Conclusion

Mastering the calculation of reaction times in first-order kinetics is essential for chemists and related professionals. The process involves understanding the relationship between concentration and time, applying the integrated rate law, and performing straightforward logarithmic calculations. In the example provided, with a rate constant of 5.67 s^{-1}, it takes just about 0.245 seconds for the concentration of reactant A to decrease from 0.500 M to 0.125 M, showcasing the rapidity of such reactions. By incorporating these principles into your studies or industrial applications, you can better predict reaction behavior, optimize conditions, and design efficient processes that leverage the fundamental laws of chemical kinetics.

Keywords for SEO Optimization:
- First-order reaction kinetics
- Reaction rate constant k
- Concentration vs. time calculations
- Integrated rate law
- How long for reaction to reach a specific concentration
- Reaction half-life
- Chemical reaction kinetics examples
- Rate constant units
- Reaction time calculation formula
- Industrial process optimization

Frequently Asked Questions

What is the rate constant for the first-order reaction AB?

The rate constant (k) for the reaction AB is 5.67 s^{-1}.

Given the initial concentration $[A]_0$ = 0.500 M, what is the target concentration [A] to reach?

The target concentration [A] is 0.125 M.

How do you calculate the time required for a first-order reaction to decrease from $[A]_0$ to a specific concentration?

Use the first-order integrated rate law: $t = (1/k) \ln([A]_0 / [A])$.

What is the formula to find the time to reach [A] = 0.125 M in this reaction?

$t = (1/5.67) \ln(0.500 / 0.125)$.

Calculate the time needed to reach [A] = 0.125 M from 0.500 M with k=5.67 s^{-1}.

$t = (1/5.67) \ln(4) \approx 0.176$ seconds.

Is the reaction faster or slower with a higher rate constant?

A higher rate constant indicates a faster reaction, meaning it reaches the target concentration sooner.

What is the significance of the natural logarithm in this calculation?

The natural logarithm relates the ratio of initial and final concentrations over time for a first-order reaction.

If the rate constant were lower, how would the time to reach [A] = 0.125 M change?

The time would increase, taking longer for the concentration to decrease to 0.125 M.

Additional Resources

6. The First-order Reaction AB, Has K= 5.67 S. If [A]o= 0.500 M, How Long Will It Take To Get To [A] = 0.125 M

Understanding reaction kinetics is fundamental for chemists and engineers alike, especially when it

comes to predicting how long a chemical process will take to reach a desired concentration. Consider a simple first-order reaction involving species A, where the rate constant (k) is given as 5.67 s^{-1}. If the initial concentration of A ($[A]_0$) is 0.500 M, the question arises: how long will it take for the concentration to decrease to 0.125 M? This scenario not only exemplifies the principles of first-order kinetics but also demonstrates how to perform practical calculations that can inform process design, safety protocols, and reaction optimization.

In this article, we will explore the mathematical foundations of first-order reactions, apply these principles to our specific problem, and delve into the implications of the calculated time. Through a detailed step-by-step approach, we aim to make the concepts accessible yet precise, equipping readers with the tools necessary to analyze similar kinetic problems.

Understanding First-Order Reactions

What Defines a First-Order Reaction?

A chemical reaction is classified as first-order when its rate depends linearly on the concentration of a single reactant. Mathematically, this is expressed as:

Rate = k [A]

where:

- Rate is the reaction rate,
- k is the rate constant (units of s^{-1} for first-order reactions),
- [A] is the concentration of reactant A at any given time.

This linear dependence means that as [A] decreases, the reaction rate slows proportionally, resulting in an exponential decay of the reactant concentration over time.

The Integrated Rate Law for First-Order Reactions

The key to solving problems involving time and concentration lies in the integrated rate law:

$\ln([A] / [A]_0) = -k t$

or equivalently,

$[A] = [A]_0 e^{(-k t)}$

where:

- $[A]_0$ is the initial concentration at time t = 0,

- [A] is the concentration at time t,
- t is the elapsed time.

This equation allows us to calculate the time required for the concentration to change from an initial value to a specified final value, given the rate constant.

Applying the Rate Law to Our Scenario

Given Data and Objective

In our specific example:

- Rate constant, $k = 5.67\ s^{-1}$
- Initial concentration, $[A]_0 = 0.500\ M$
- Final concentration, $[A] = 0.125\ M$

Our goal is to determine t, the time it takes for [A] to decrease from 0.500 M to 0.125 M.

Step-by-Step Calculation

Using the integrated rate law:

$$\ln([A] / [A]_0) = -k\ t$$

Rearranged to solve for t:

$$t = (1 / k) \ln([A]_0 / [A])$$

Plugging in the known values:

$$t = (1 / 5.67\ s^{-1}) \ln(0.500\ M / 0.125\ M)$$

Calculating the ratio:

$$0.500 / 0.125 = 4$$

Calculating the natural logarithm:

$$\ln(4) \approx 1.3863$$

Finally, computing t:

$$t = (1 / 5.67)\ 1.3863 \approx 0.2447 \text{ seconds}$$

Result: It will take approximately 0.245 seconds for the concentration of A to decrease from 0.500 M to 0.125 M.

Interpreting the Results and Practical Implications

Understanding the Speed of the Reaction

A rate constant of 5.67 s^{-1} indicates an extremely rapid reaction. The calculated time of roughly 0.245 seconds underscores how quickly reactants can be consumed in such a system. This has several implications:

- Industrial Processes: For reactions with such high rate constants, reactors need to be designed to ensure proper mixing and control, as the reaction proceeds almost instantaneously on a human timescale.
- Safety Considerations: Rapid reactions may lead to sudden temperature spikes or pressure build-ups, requiring safety measures to prevent accidents.
- Analytical Challenges: Measuring concentrations accurately in such short timescales demands high-speed analytical techniques.

Limitations and Assumptions

While the calculations provide a clear answer, it's essential to acknowledge assumptions:

- The reaction strictly follows first-order kinetics throughout.
- The rate constant remains unchanged over the course of the reaction.
- External factors like temperature and pressure are constant.
- No side reactions or reversibility affects the concentration change.

In real-world scenarios, deviations from these assumptions may occur, and more complex models might be necessary.

Extending the Analysis

The approach used here can be generalized to similar kinetic problems:

- Different initial concentrations: Adjust $[A]_0$ accordingly.
- Other target concentrations: Substitute different [A] values.
- Varying rate constants: Understand how changes in k influence reaction speed.

By mastering these calculations, scientists can predict reaction times, optimize processes, and design experiments more effectively.

Concluding Remarks

Understanding the kinetics of reactions is crucial for controlling and optimizing chemical processes. In the case of a first-order reaction with a high rate constant, the reaction proceeds rapidly, as exemplified by our calculation showing a reaction time of less than a quarter of a second to reduce the concentration of A from 0.500 M to 0.125 M. This example highlights how the integrated rate law serves as a powerful tool in translating kinetic constants into practical timeframes.

Whether in industrial synthesis, environmental monitoring, or academic research, the ability to predict reaction times based on concentration changes and rate constants empowers chemists to make informed decisions. As reactions become more complex, the principles illustrated here serve as foundational knowledge, paving the way for deeper exploration into advanced kinetic models.

In summary, the precise calculation of reaction times enables better control, safety, and efficiency in chemical applications. Recognizing the speed of reactions with large rate constants like 5.67 s^{-1} emphasizes the importance of meticulous experimental design and real-time monitoring—an essential aspect of modern chemistry and chemical engineering.

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