If A Reaction Is First Order With A Rate Constant Of 0.0450 S', How Much Time Is Required For 65% Of

If A Reaction Is First Order With A Rate Constant Of 0.0450 S', How Much Time Is Required For 65% Of the reactant to be consumed or transformed into the product? This question is fundamental in chemical kinetics, as understanding the relationship between reaction rates and time allows chemists and engineers to control and optimize industrial processes, laboratory reactions, and environmental systems. In this article, we will explore the principles behind first-order reactions, derive the relevant equations, perform calculations for the specific case of 65% conversion, and discuss practical implications and applications of these concepts.

Understanding 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, this can be expressed as:

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

where:
- Rate is the reaction rate,
- k is the rate constant (with units of s^{-1}),
- $[A]$ is the concentration of reactant A.

This linear dependence means that the rate of reaction decreases exponentially as the reactant is consumed.

Characteristics of First-Order Reactions

- The half-life ($t_{1/2}$) is independent of initial concentration and is given by:

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

- The concentration of reactant decreases exponentially over time:

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

where:
- $[A]_0$ is the initial concentration,
- $[A](t)$ is the concentration at time t.

Deriving the Time for a Specific Conversion in a First-Order Reaction

Knowing the rate constant k and the desired extent of reaction (e.g., 65% conversion), we can calculate the required time using the integrated first-order rate law.

Integrated Rate Law for First-Order Reactions

The integrated rate law relates the concentration at any time t to the initial concentration:

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

Alternatively, expressing the fractional conversion:

$$
\frac{[A](t)}{[A]_0} = e^{-kt}
$$

If X is the fraction of reactant converted, then:

$$
X = \frac{[A]_0 - [A](t)}{[A]_0}
$$

or

$$
[A](t) = [A]_0 (1 - X)
$$

Thus,

$$
1 - X = e^{-kt}
$$

which can be rearranged to solve for t:

$$
t = -\frac{1}{k} \ln(1 - X)
```
```

\]

Calculating Time for 65% Conversion

Given:
- \(k = 0.0450\, \text{s}^{-1} \),
- \(X = 0.65 \) (since 65% of reactant is converted).

Applying the formula:

\[
t = -\frac{1}{0.0450} \ln(1 - 0.65)
\]

\[
t = -\frac{1}{0.0450} \ln(0.35)
\]

Calculate \(\ln(0.35) \):

\[
\ln(0.35) \approx -1.0498
\]

Now, compute \(t \):

\[
t = -\frac{1}{0.0450} \times (-1.0498) = \frac{1.0498}{0.0450}
\]

\[
t \approx 23.33\, \text{seconds}
\]

Therefore, approximately 23.33 seconds are required for 65% of the reactant to be consumed in a first-order reaction with a rate constant of 0.0450 s\(^{-1}\).

Practical Implications and Applications

Industrial Reactor Design

Understanding the time required for a certain conversion helps in designing reactors such as continuous stirred-tank reactors (CSTR), plug flow reactors (PFR), or batch reactors. For example, knowing that 65% conversion occurs in about 23 seconds allows engineers to determine the necessary residence time in a reactor to achieve desired yields.

Reaction Monitoring and Control

Monitoring reaction progress in real-time is crucial for process control. Using rate constants and the integrated rate law, operators can predict how long it will take to reach target conversions, enabling better scheduling, safety, and efficiency.

Environmental and Pharmaceutical Processes

In environmental chemistry, understanding reaction times for pollutant degradation is essential for assessing cleanup durations. Similarly, in pharmaceuticals, reaction kinetics influence the formulation and stability of drug compounds.

Factors Influencing Reaction Rates

While the calculation assumes ideal conditions, real-world reactions are influenced by various factors:

- **Temperature:** Higher temperatures generally increase k due to Arrhenius' law.
- **Concentration:** For first-order reactions, k remains constant regardless of initial concentration, but in other reaction orders, it can vary.
- **Catalysts:** Catalysts can significantly increase k, reducing reaction time.
- **Pressure and Physical State:** Changes can affect reaction kinetics, especially for gases or heterogeneous reactions.

Extensions and Related Calculations

Beyond calculating the time for a certain percentage of reaction completion, other related calculations include:

1. **Half-life ($t_{1/2}$):** Time for 50% conversion:

   $$
   t_{1/2} = \frac{\ln 2}{k} \approx \frac{0.693}{0.0450} \approx 15.4\, \text{seconds}
   $$

2. **Time for other conversions:** For example, 90% conversion:

   $$
   t = -\frac{1}{k} \ln(1 - 0.9) = -\frac{1}{0.0450} \ln(0.10) \approx 51.13\, \text{seconds}
   $$

\]

Summary and Key Takeaways

- For a first-order reaction with \(k = 0.0450\, \text{s}^{-1} \), approximately 23.33 seconds are needed to achieve 65% conversion.
- The exponential nature of first-order kinetics allows straightforward calculation of reaction times for desired conversions.
- Understanding these principles aids in designing efficient chemical processes, optimizing reaction conditions, and ensuring safety and consistency.

Conclusion

Mastering the relationship between reaction rate constants and reaction times is vital for chemists and chemical engineers. The ability to predict how long a reaction will take to reach a certain conversion enables better process control and optimization. By applying the integrated rate law for first-order reactions, we can easily determine that with a rate constant of 0.0450 S\(^{-1}\), a reaction will take approximately 23.33 seconds to achieve 65% conversion. This knowledge not only enhances our understanding of reaction kinetics but also directly impacts practical applications across various industries.

Note: Always consider real-world factors such as temperature fluctuations, catalyst presence, and measurement uncertainties when applying these calculations to actual systems.

Frequently Asked Questions

How do you calculate the time required for 65% of a first-order reaction to complete with a rate constant of 0.0450 s^{-1}?

Use the first-order integrated rate law: t = (1/k) ln(1/(1 - fraction reacted)). For 65%, fraction reacted = 0.65, so t = (1/0.0450) ln(1/0.35) ≈ 22.22 1.049 = approximately 23.33 seconds.

What is the significance of the rate constant in determining reaction time?

The rate constant (k) indicates how fast a reaction proceeds; a higher k results in a shorter time for a given percentage of reaction, while a lower k means the reaction takes longer.

Can you explain the first-order kinetics formula used in this calculation?

Yes, for a first-order reaction, the time t for a certain fraction to react is given by t = (1/k) ln(1/(1 - fraction reacted)). This relates the rate constant to the extent of reaction over time.

What does a rate constant of 0.0450 s^{-1} tell us about the reaction speed?

A rate constant of 0.0450 s^{-1} indicates a relatively moderate reaction speed, meaning it takes about 23 seconds for 65% of the reactant to be consumed.

How would the required time change if the reaction were 90% complete instead of 65%?

For 90% completion (fraction reacted = 0.9), t = (1/0.0450) ln(1/0.1) ≈ 22.22 2.303 ≈ 51.14 seconds, longer than for 65% completion.

What assumptions are made when using the first-order rate law in this calculation?

The assumptions include that the reaction is truly first order, the rate constant remains constant over time, and the reaction conditions are constant without side reactions or reversibility.

How does temperature affect the rate constant and reaction time in this context?

Increasing temperature generally increases the rate constant (k), thereby decreasing the time required for a given reaction extent; decreasing temperature has the opposite effect.

Is the calculation affected if the reaction is not purely first order?

Yes, if the reaction is not truly first order, the calculation may not be accurate; different kinetic models would need to be used to determine the time.

What practical applications can this calculation have in chemical processes?

It helps in designing reactors, optimizing reaction times, and controlling processes to ensure desired product yields within specific time frames.

How can this information be used to control reaction

conditions in industrial settings?

By adjusting temperature, catalysts, or reactant concentrations to modify the rate constant, operators can achieve desired reaction extents within target times efficiently.

Additional Resources

Understanding Reaction Kinetics: How Long Does It Take for 65% of a First-Order Reaction to Complete?

When exploring chemical reactions, especially in the realm of kinetics, understanding the relationship between reaction order, rate constants, and reaction completion times is fundamental. This article delves into the specifics of a first-order reaction with a given rate constant, and how to determine the time required for a certain percentage of reactant consumption—in this case, 65%.

Introduction to Reaction Kinetics and First-Order Reactions

Reaction kinetics is the branch of chemistry that studies the speed or rate at which chemical reactions occur and the factors influencing these rates. It provides insights into reaction mechanisms and helps in designing processes in industrial, biological, and environmental applications.

First-order reactions are among the simplest kinetic models. They are characterized by a rate that is directly proportional to the concentration of a single reactant:

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

where:
- $ [A] $ is the concentration of reactant A,
- $ k $ is the rate constant (units: $ \text{s}^{-1} $ for first-order reactions).

Mathematical Foundations of First-Order Reaction Kinetics

Understanding the mathematical basis is crucial for calculating the time required for a certain degree of reaction completion.

Integrated Rate Law for First-Order Reactions:

$$

$$
\ln \left(\frac{[A]_0}{[A]} \right) = kt
$$

where:
- $[A]_0$ is the initial concentration,
- $[A]$ is the concentration at time t,
- k is the rate constant,
- t is the elapsed time.

This equation relates concentration changes over time for first-order reactions, simplifying calculations of reaction times for specific extents of conversion.

Calculating Time for 65% Reaction Completion

The question: How much time is required for 65% of the reactant to be consumed?

Step 1: Define the extent of reaction

If 65% of the reactant has reacted, then 35% remains:

$$
\frac{[A]}{[A]_0} = 1 - 0.65 = 0.35
$$

Step 2: Apply the integrated rate law

Rearranged for time:

$$
t = \frac{1}{k} \ln \left(\frac{[A]_0}{[A]} \right)
$$

Substitute:

$$
t = \frac{1}{0.0450\, \text{s}^{-1}} \times \ln \left(\frac{1}{0.35} \right)
$$

Step 3: Calculate the natural logarithm

$$
\ln \left(\frac{1}{0.35} \right) = \ln(2.8571) \approx 1.0498
$$

Step 4: Final calculation

$$
\[

$$
t = \frac{1}{0.0450} \times 1.0498 \approx 23.33\, \text{s}
$$

Result: It takes approximately 23.33 seconds for 65% of the reactant to be consumed in this first-order reaction.

Deeper Insights and Practical Considerations

1. Significance of the Rate Constant k:

The rate constant $k = 0.0450\, \text{s}^{-1}$ indicates the speed of the reaction. Larger k values correspond to faster reactions, meaning less time is needed for a given extent of conversion.

- For reactions with the same k, the time for 65% completion can be directly compared.
- Adjusting k (e.g., through temperature changes or catalysts) influences reaction times significantly.

2. Effect of Temperature (Arrhenius Equation):

The rate constant is temperature-dependent, described by the Arrhenius equation:

$$
k = A e^{-\frac{E_a}{RT}}
$$

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

An increase in temperature generally increases k, reducing the time needed for a certain reaction extent.

3. Reaction Monitoring and Experimental Validation:

In laboratory or industrial settings, reaction progress is often monitored through spectroscopic methods, chromatography, or titrations. Knowing the theoretical time helps in planning experiments and processes.

4. Limitations and Assumptions:

- The calculation assumes ideal first-order behavior throughout the reaction.
- No side reactions or reversibility are considered.
- Conditions such as temperature, pressure, and solvent environment are constant.

Extending the Analysis: Time for Other Percentages

Understanding how to calculate the time for other extents of reaction is straightforward with the same method:

$$
t = \frac{1}{k} \ln \left(\frac{[A]_0}{[A]} \right)
$$

For example:
- For 90% reacted ($[A]/[A]_0 = 0.10$):

$$
t = \frac{1}{0.0450} \times \ln(10) \approx 51.12\, \text{s}
$$

- For 50% reacted ($[A]/[A]_0 = 0.50$):

$$
t = \frac{1}{0.0450} \times \ln(2) \approx 15.4\, \text{s}
$$

This illustrates the exponential nature of first-order reactions: as the reactant diminishes, the rate slows, but the time to reach high conversion levels increases logarithmically.

Practical Applications and Implications

1. Industrial Chemical Processes:

Knowing the time for a specific percentage of conversion allows engineers to optimize reactor design, batch times, and process efficiencies.

2. Pharmacokinetics:

In drug metabolism, first-order kinetics are common. Calculating how long it takes for a drug to reach a certain level of clearance informs dosing schedules.

3. Environmental Chemistry:

Pollutant degradation often follows first-order kinetics. Estimating degradation times helps in environmental risk assessments.

4. Biological Systems:

Enzymatic reactions frequently exhibit first-order kinetics under certain conditions, guiding biochemical pathway analysis.

Key Takeaways and Summary

- The rate constant $k = 0.0450\, \text{s}^{-1}$ indicates a relatively rapid first-order reaction.
- To determine the time for 65% of reactant consumption, use the integrated rate law:

$$
t = \frac{1}{k} \ln \left(\frac{1}{1 - \text{fraction reacted}} \right)
$$

which, for 65%, becomes:

$$
t \approx 23.33\, \text{s}
$$

- The method is extendable to any reaction extent, providing a versatile tool for kinetic analysis.
- Real-world applications depend on maintaining the assumptions of ideal first-order kinetics; deviations may require more complex models.

Final Remarks

Understanding reaction times based on kinetic parameters is essential across many scientific disciplines. The ability to accurately predict how long a reaction takes to reach a certain completion level enables better process control, safety, and efficiency. This example demonstrates the power of straightforward mathematical relationships in chemistry, emphasizing the importance of grasping fundamental concepts like the integrated rate law.

By mastering these calculations, chemists and engineers can design more effective reactions, optimize conditions, and interpret experimental data with confidence. Whether in laboratory research or industrial production, the principles outlined here form the backbone of kinetic analysis for first-order reactions.

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Disclaimer: The calculations assume ideal conditions and do not account for factors such as temperature fluctuations, catalyst presence, or reaction reversibility which may influence real-world

reaction times.

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