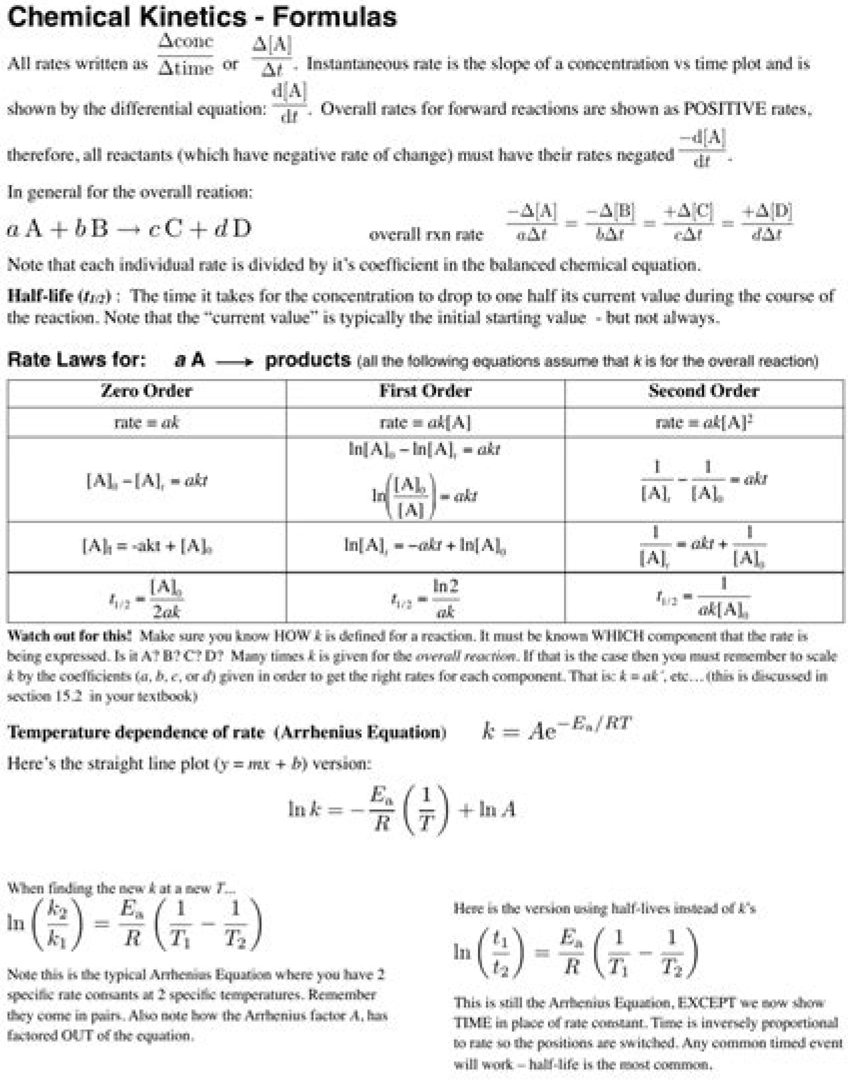
Chemistry Kinetics Cheat Sheet



Chemical Kinetics - Formulas

All rates written as $\frac{\Delta \text{conc}}{\Delta \text{time}}$ or $\frac{\Delta[\text{A}]}{\Delta t}$. Instantaneous rate is the slope of a concentration vs time plot and is shown by the differential equation: $\frac{\text{d}[\text{A}]}{\text{d}t}$. Overall rates for forward reactions are shown as POSITIVE rates, therefore, all reactants (which have negative rate of change) must have their rates negated $\frac{-\text{d}[\text{A}]}{\text{d}t}$.

In general for the overall reation:

$a\,\text{A} + b\,\text{B} \rightarrow c\,\text{C} + d\,\text{D}$ overall rxn rate $\frac{-\Delta[\text{A}]}{a\Delta t} = \frac{-\Delta[\text{B}]}{b\Delta t} = \frac{+\Delta[\text{C}]}{c\Delta t} = \frac{+\Delta[\text{D}]}{d\Delta t}$

Note that each individual rate is divided by it's coefficient in the balanced chemical equation.

Half-life ($t_{1/2}$) : The time it takes for the concentration to drop to one half its current value during the course of the reaction. Note that the "current value" is typically the initial starting value - but not always.

Rate Laws for: $a\,\text{A} \longrightarrow$ products (all the following equations assume that k is for the overall reaction)

Zero Order	First Order	Second Order
rate = ak	rate = $ak[\text{A}]$	rate = $ak[\text{A}]^2$
$[\text{A}]_0 - [\text{A}]_t = akt$	$\ln[\text{A}]_0 - \ln[\text{A}]_t = akt$ $\ln\left(\frac{[\text{A}]_0}{[\text{A}]}\right) = akt$	$\frac{1}{[\text{A}]_t} - \frac{1}{[\text{A}]_0} = akt$
$[\text{A}]_t = -akt + [\text{A}]_0$	$\ln[\text{A}]_t = -akt + \ln[\text{A}]_0$	$\frac{1}{[\text{A}]_t} = akt + \frac{1}{[\text{A}]_0}$
$t_{1/2} = \frac{[\text{A}]_0}{2ak}$	$t_{1/2} = \frac{\ln 2}{ak}$	$t_{1/2} = \frac{1}{ak[\text{A}]_0}$

Watch out for this! Make sure you know HOW k is defined for a reaction. It must be known WHICH component that the rate is being expressed. Is it A? B? C? D? Many times k is given for the *overall reaction*. If that is the case then you must remember to scale k by the coefficients (a, b, c, or d) given in order to get the right rates for each component. That is: $k = ak'$, etc… (this is discussed in section 15.2 in your textbook)

Temperature dependence of rate (Arrhenius Equation) $k = A\text{e}^{-E_a/RT}$

Here's the straight line plot ($y = mx + b$) version:

$$\ln k = -\frac{E_a}{R}\left(\frac{1}{T}\right) + \ln A$$

When finding the new k at a new T...

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Note this is the typical Arrhenius Equation where you have 2 specific rate consants at 2 specific temperatures. Remember they come in pairs. Also note how the Arrhenius factor A, has factored OUT of the equation.

Here is the version using half-lives instead of k's

$$\ln\left(\frac{t_1}{t_2}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

This is still the Arrhenius Equation, EXCEPT we now show TIME in place of rate constant. Time is inversely proportional to rate so the positions are switched. Any common timed event will work – half-life is the most common.

Chemistry kinetics cheat sheet is an essential tool for students and professionals alike in the field of chemistry. Understanding the principles of chemical kinetics is vital for predicting how fast reactions occur and the factors influencing these rates. This article aims to provide a comprehensive overview of the fundamentals of chemical kinetics, including key concepts, equations, and examples, making it an invaluable resource for anyone studying this intriguing branch of chemistry.

Introduction to Chemical Kinetics

Chemical kinetics studies the rates of chemical reactions and the factors that affect these rates. This discipline provides insights into how reactions proceed, allowing chemists to manipulate conditions to control the speed of a reaction.

Key Concepts in Chemical Kinetics

1. Reaction Rate: This is the change in concentration of a reactant or product per unit time. It can be expressed mathematically as:
$$
\text{Rate} = -\frac{d[A]}{dt} \quad \text{(for reactants)}
$$
$$
\text{Rate} = \frac{d[B]}{dt} \quad \text{(for products)}
$$
where $[A]$ and $[B]$ are concentrations of reactants and products, respectively.

2. Factors Affecting Reaction Rates:
- Concentration: Increasing the concentration of reactants typically increases the reaction rate.
- Temperature: Higher temperatures usually increase reaction rates due to increased kinetic energy.
- Catalysts: Substances that increase the reaction rate without being consumed in the process.
- Surface Area: For solid reactants, increasing the surface area (e.g., by grinding) can enhance the reaction rate.

3. Order of Reaction: The order indicates how the rate of reaction depends on the concentration of reactants. It can be zero, first, second, or higher.

Rate Laws and Reaction Orders

Rate laws express the relationship between the rate of a reaction and the concentration of its reactants. The general form of a rate law is:
$$
\text{Rate} = k[A]^m[B]^n
$$
where:
- k = rate constant
- $[A]$ and $[B]$ = concentrations of reactants
- m and n = orders of the reaction with respect to each reactant

Determining Reaction Order

To determine the order of a reaction experimentally, one can use methods such as:

1. Method of Initial Rates: Measure the initial reaction rates at varying concentrations of reactants.
2. Integrated Rate Laws: Analyze concentration vs. time data.
- Zero-Order Reactions: Rate is constant, independent of concentration.
- First-Order Reactions: Rate is directly proportional to the concentration of one reactant.
- Second-Order Reactions: Rate is proportional to the square of the concentration of one reactant or to the product of the concentrations of two reactants.

Half-Life of Reactions

Half-life ($t_{1/2}$) is the time required for the concentration of a reactant to decrease to half its initial concentration. The half-lives for different order reactions are:

- Zero-Order:
$$
t_{1/2} = \frac{[A]_0}{2k}
$$

- First-Order:
$$
t_{1/2} = \frac{0.693}{k}
$$

- Second-Order:
$$
t_{1/2} = \frac{1}{k[A]_0}
$$

Activation Energy and the Arrhenius Equation

Activation energy (E_a) is the minimum energy required for a reaction to occur. The Arrhenius equation relates the rate constant k to temperature T and activation energy:
$$
k = A e^{-\frac{E_a}{RT}}
$$
where:
- A = pre-exponential factor
- R = universal gas constant (8.314 J/(mol·K))
- T = temperature in Kelvin

Understanding the Arrhenius Equation

The Arrhenius equation can be linearized to facilitate analysis:
\[
\ln k = \ln A - \frac{E_a}{R}\left(\frac{1}{T}\right)
\]
This linear form allows for the determination of activation energy by plotting \(\ln k\) against \(\frac{1}{T}\).

Reaction Mechanisms

A reaction mechanism is a step-by-step description of how a reaction occurs at the molecular level. Each step in the mechanism is called an elementary reaction.

Elementary Steps and Rate Determining Step

1. Elementary Steps: These are individual processes that lead to the overall reaction. Each step has its own rate law.
2. Rate-Determining Step (RDS): The slowest step in a mechanism that determines the overall reaction rate. The rate law can often be derived from this step.

Examples of Reaction Mechanisms

- Unimolecular Reactions: Involves a single reactant molecule and may follow first-order kinetics.
- Bimolecular Reactions: Involves two reactant molecules and may follow second-order kinetics.

Catalysis in Chemical Kinetics

Catalysts play a crucial role in chemical kinetics by lowering the activation energy, thus increasing the reaction rate without being consumed.

Types of Catalysts

1. Homogeneous Catalysts: In the same phase as the reactants (e.g., gases or liquids).
2. Heterogeneous Catalysts: In a different phase, often solid catalysts in

contact with gaseous or liquid reactants.

Examples of Catalytic Reactions

- Haber Process: Ammonia synthesis using iron as a heterogeneous catalyst.
- Enzymatic Reactions: Biological catalysts (enzymes) that speed up biochemical reactions.

Practical Applications of Chemical Kinetics

Understanding kinetics is crucial in various fields including:

- Pharmaceuticals: Designing drugs with optimal reaction rates.
- Environmental Chemistry: Studying pollutant degradation.
- Industrial Processes: Enhancing the efficiency of chemical manufacturing.

Conclusion

The chemistry kinetics cheat sheet serves as a quick reference to key concepts, equations, and examples that are fundamental in understanding reaction rates and mechanisms. By mastering these principles, students and professionals can better predict and manipulate chemical reactions in both laboratory and industrial settings. This foundational knowledge of kinetics not only aids in academic pursuits but also has practical implications across various scientific and engineering disciplines.

Frequently Asked Questions

What is chemistry kinetics?

Chemistry kinetics is the branch of physical chemistry that studies the rates of chemical reactions and the factors affecting these rates.

What are the main factors affecting reaction rates in chemistry kinetics?

The main factors include concentration of reactants, temperature, presence of catalysts, surface area of reactants, and the nature of the reactants.

What is the difference between reaction rate and

rate constant?

The reaction rate is the speed at which reactants are converted to products, while the rate constant is a specific value that relates the reaction rate to the concentrations of reactants in a rate law expression.

What is a rate law in chemistry kinetics?

A rate law is an equation that relates the rate of a reaction to the concentration of reactants, typically expressed as rate = k[A]^m[B]^n, where k is the rate constant and m and n are the reaction orders.

What is the Arrhenius equation?

The Arrhenius equation describes the temperature dependence of reaction rates, given by k = Ae^(-Ea/RT), where A is the pre-exponential factor, Ea is the activation energy, R is the gas constant, and T is the temperature in Kelvin.

What role do catalysts play in reaction kinetics?

Catalysts increase the rate of a reaction by lowering the activation energy required for the reaction to proceed, without being consumed in the process.

How can reaction mechanisms be determined?

Reaction mechanisms can be determined through experimental data, such as reaction rates at various concentrations, and by proposing a series of elementary steps that add up to the overall reaction.

What are zero-order reactions in kinetics?

Zero-order reactions are those where the rate of reaction is constant and does not depend on the concentration of reactants; the rate is independent of the concentration, often due to saturation of a catalyst or surface.

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