# Chemical Kinetics: Understanding Reaction Rates

## I. Introduction to Reaction Rates

- **A. What is Chemical Kinetics?**
    - The study of the rates of chemical reactions.
    - Kinetics addresses *how fast* a reaction goes, whereas thermodynamics predicts the *spontaneous direction and extent* of the change.
- **B. Defining and Measuring Rate**
    - The rate is the change in the molar concentration of a reactant or product divided by the time interval over which the change takes place.

## II. Molecular Basis: Reaction Mechanisms

- **A. Mechanisms and Elementary Reactions**
  - A **reaction mechanism** is the proposed pathway (sequence of steps) for an overall reaction.
  - **Elementary reactions** are distinct single steps (often collisions of particles) that make up a mechanism.
  - The rate law for an elementary reaction *can* be written directly from its chemical equation, with coefficients serving as exponents.
  - The molecularity is the number of reactant particles involved (e.g., unimolecular, bimolecular, or the rare termolecular).
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- **B. The Rate-Determining Step (RDS)**
  - The speed of the overall reaction is governed by the slowest step in the sequence, known as the **rate-determining step**.
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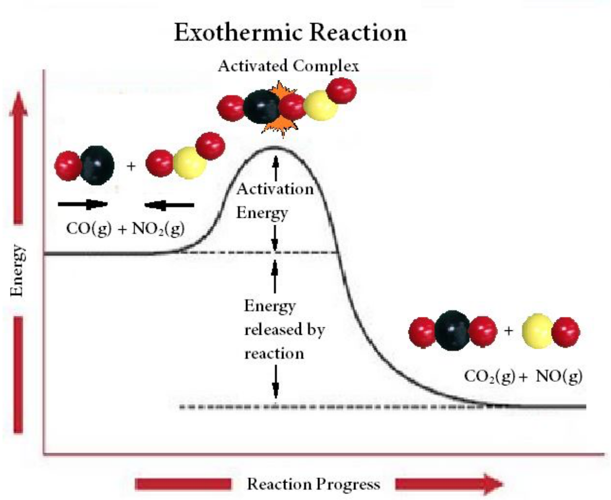
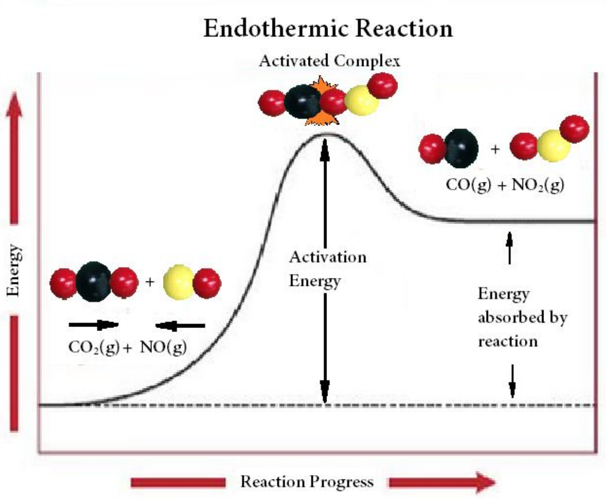
- **C. Rates and Equilibrium**
  - At equilibrium, the rates of the forward and reverse reactions are equal.
  - The equilibrium constant ($K$) for a reaction is equal to the ratio of the rate constants for the forward ($k$) and reverse ($k'$) elementary reactions: $K = k/k'$.

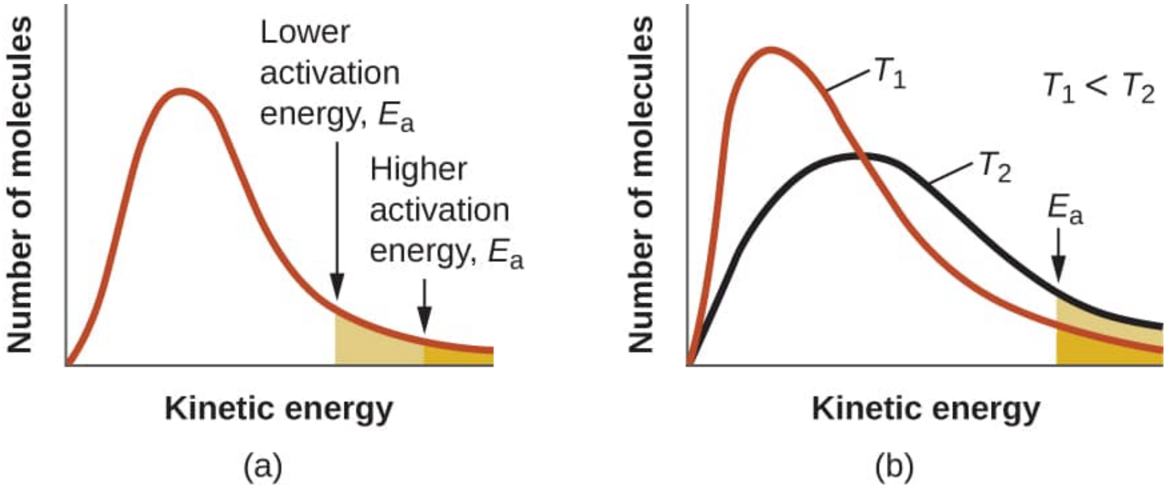
## III. Determining Reaction Speed: Energy and Temperature

**A. Collision Theory**: Requires that reactant molecules collide with kinetic energy $\geq E_a$ and in the correct relative orientation (steric requirement, $A$).





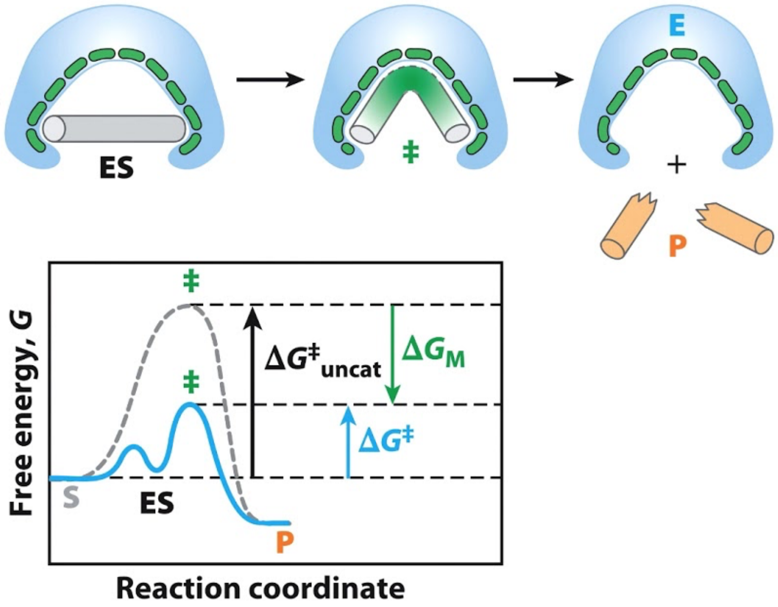
- **B. The Activation Energy ($E_a$)**
  - The activation energy is the **minimum energy needed for reaction**.
  - It represents the height of the **activation barrier**.
  - Reactions are slow when they have high activation energies, meaning hardly any collisions result in reaction.
  - Reactions with high $E_a$ have rates that depend strongly on temperature (a steep Arrhenius plot).
  - **Transition State Theory**: Explains that molecules must acquire sufficient energy to form an unstable **activated complex** (transition state) and successfully cross the energy barrier. The activation energy measures the energy of this activated complex relative to the reactants.



- **A. Effect of Temperature**
    - Reaction rates generally increase with temperature.
    - This dependence is described by the Arrhenius equation: $\ln k = \ln A - \frac{E_a}{RT}$

Consider kinetic molecular theory of gases, what is RT?

Enzymes stabilize the **transition state**.



## V. Accelerating Reactions: Catalysis

### A. Catalyst Function

- A **catalyst** is a substance that increases the rate of a reaction without being consumed.
- A catalyst works by providing an **alternative reaction pathway** that has a lower activation energy ($E_a$) than the original path.

### B. Effect on Equilibrium

- A catalyst accelerates both the forward and reverse reactions equally, thus having **no effect** on the final equilibrium composition.

### C. Types of Catalysts

- **Homogeneous catalyst**: In the same phase as the reactants (e.g., dissolved in a liquid solution).
- **Heterogeneous catalyst**: In a different phase from the reactants (e.g., a solid surface used in gas-phase reactions).
- **Enzymes**: Biological catalysts, often proteins, that utilize a slot-like **active site** to interact with a substrate molecule.