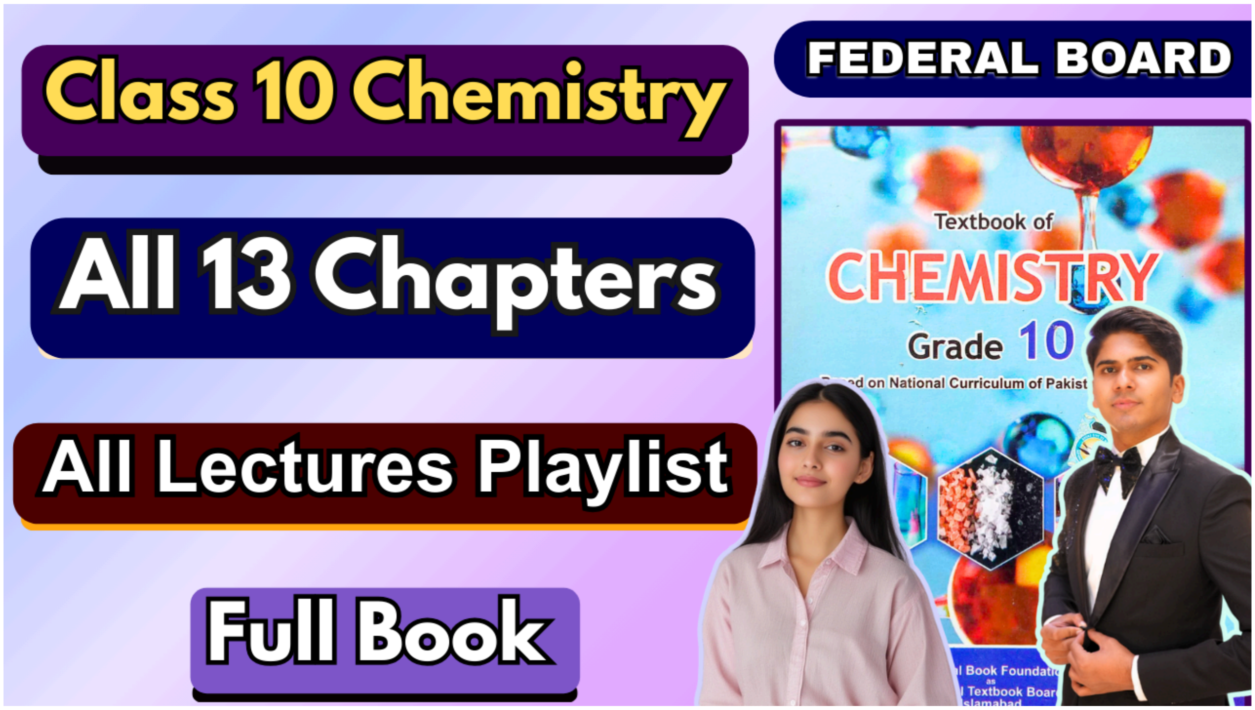
Chapter 5 - Chemical Kinetics

All Lectures Uploaded on YouTube:

https://tinyurl.com/fkm10-chemistry



Chemical Kinetics

Chemical Kinetics is the study of:

- The **rates** (speeds) of chemical reactions.
- The **mechanisms** (steps) by which reactions occur.
- The **factors** that influence these rates.

Reaction rates vary greatly:

- **Slow:** Fermentation (weeks), Digestion.
- **Fast:** Acid-base neutralization (microseconds).

- **Moderate:** Muscle contraction, nerve impulses, photography.

Importance: Understanding kinetics is crucial for industry to make chemical processes **cost-effective**.

5.1. Rates of Reactions

The rate of a reaction tells us how quickly **reactants are consumed** or **products are formed** over time.It is defined as the **change in concentration** of a reactant or product per unit time.

Mathematical Formula:

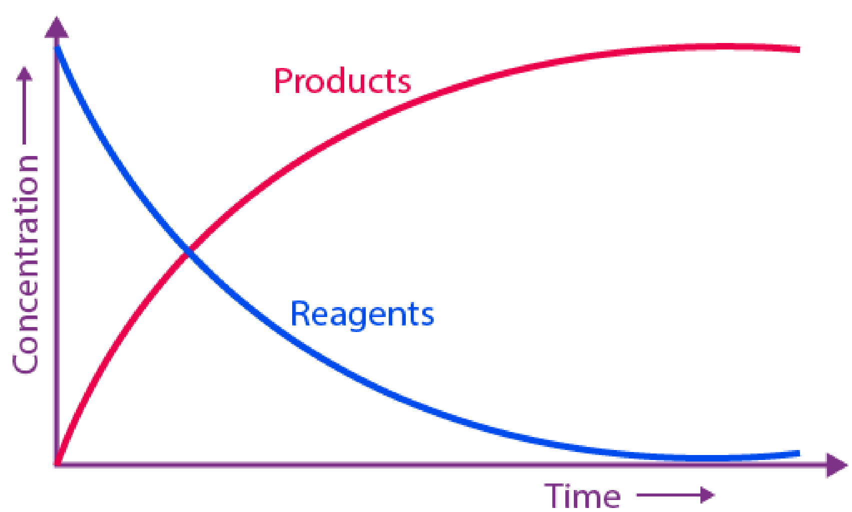
$$\text{Rate} = \frac{Change\ in\ concentration\ of\ a\ substance}{Time\ taken\ for\ change}$$

Common Units: moles per cubic decimeter per second ($mol\ dm^{-3}\ s^{-1}$).

Graphical Representation (Concentration vs. Time)

A typical graph shows two curves:

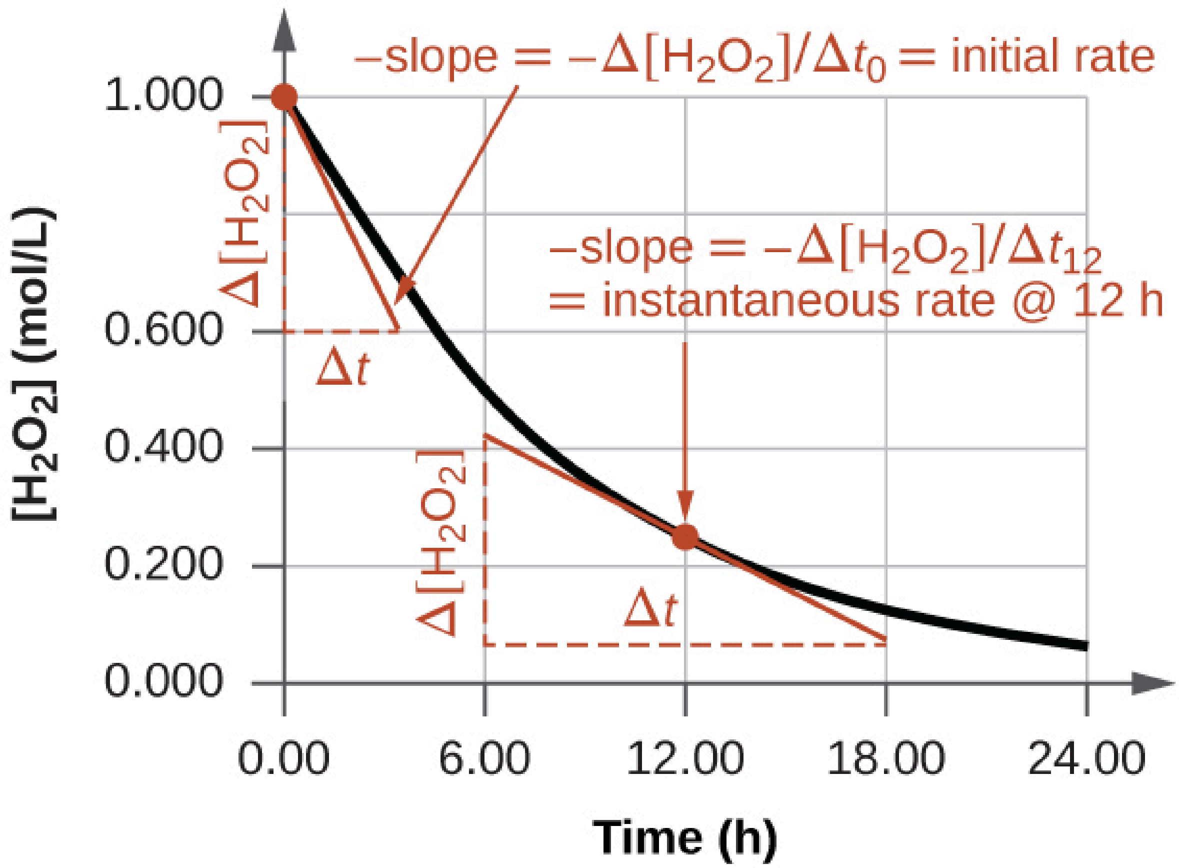
- **Reactant Curve:** Starts high and decreases over time.
- **Product Curve:** Starts at zero and increases over time.



Observations from the Graph:

1. **Initial Stage (Steep Slope):**

At the start, the curves are steep. This means the reaction is **fast** because there are many reactant particles available to collide.



2. **As Time Progresses (Slope Decreases):**

The curves become less steep. The reaction **slows down** because the concentration of reactants decreases, leading to fewer successful collisions.

3. **Final Stage (Flat Curve):**

The curves become horizontal (flat).The reaction has **stopped** because the reactants are fully used up, or the system has reached **equilibrium**.

Conclusion: The rate of a reaction is **not constant**; it is highest at the start and decreases over time.

Average Rate of Reaction

This gives the overall speed of the reaction over a specific time interval.

Formula:

$$\text{Average rate} = \frac{Total\ change\ in\ concentration}{Total\ time\ taken}$$

Expressing Reaction Rate Mathematically

For a simple reaction: **A ⟶ B**

- The rate can be expressed as the **disappearance of reactant A**:

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

  (The negative sign shows concentration is decreasing)

- Or as the **appearance of product B**:

  $$\text{Rate} = +\frac{d[B]}{dt}$$

  (The positive sign shows concentration is increasing)

Where d[A] and d[B] are the small changes in concentration, and dt is the small change in time.

Interpreting Reaction Rate Data

Example Reaction: A + B ⟶ C

The rate of this reaction can be followed by measuring the **concentration of the product (C)** at regular time intervals.

Data Table & Analysis:

Time(min)	Concentration of C (moldm^{-3})	Interpretation
0.0	0.00	Reaction has not yet started.
20	15	Product C is forming **rapidly**. The rate is high.
40	21	The increase in [C] is slower, indicating the **reaction is slowing down**.
60	23	The reaction continues to slow. Very little increase in [C].
80	25	No further change in concentration.
100	25	The reaction has **stopped**; likely because a reactant has been used up.

The data visually demonstrates that the **reaction rate decreases over time** as reactants are consumed.

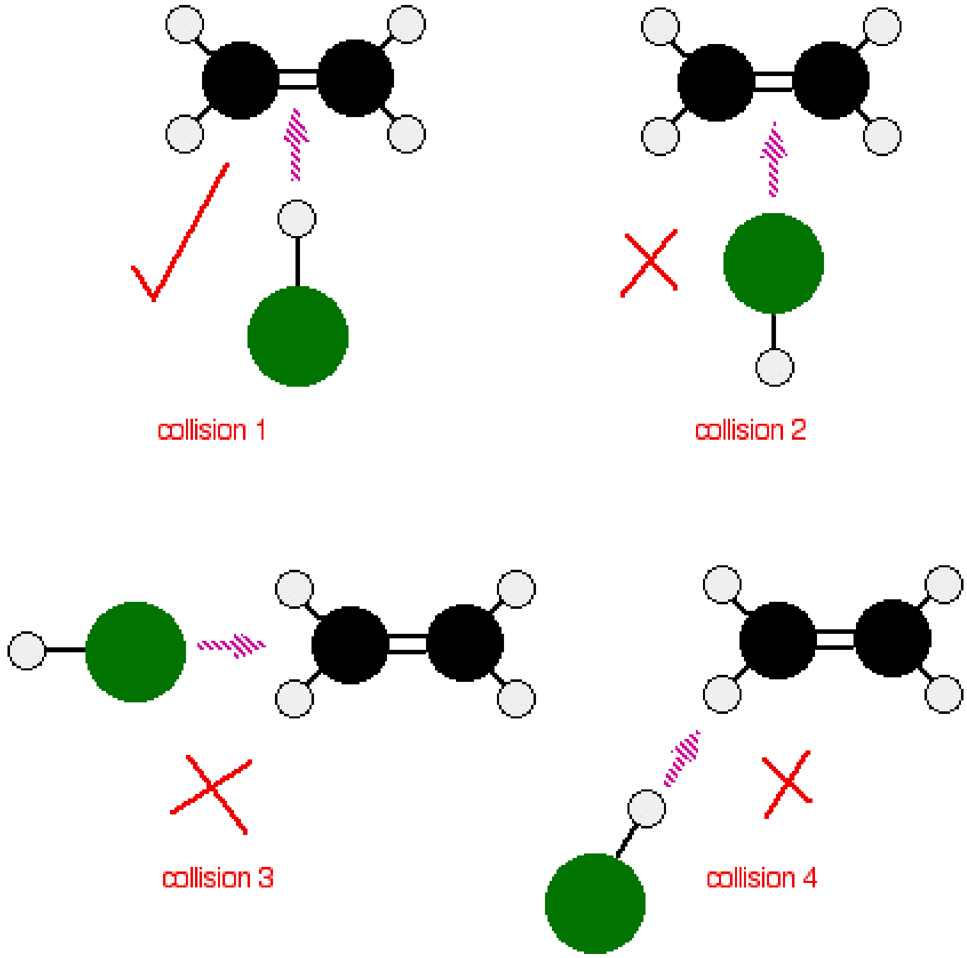
5.2. Collision Theory and Activation Energy

Collision theory explains the requirements for a chemical reaction to occur at the molecular level.

Two Conditions for a Successful Reaction

For a collision between particles to result in a reaction, two criteria must be met:

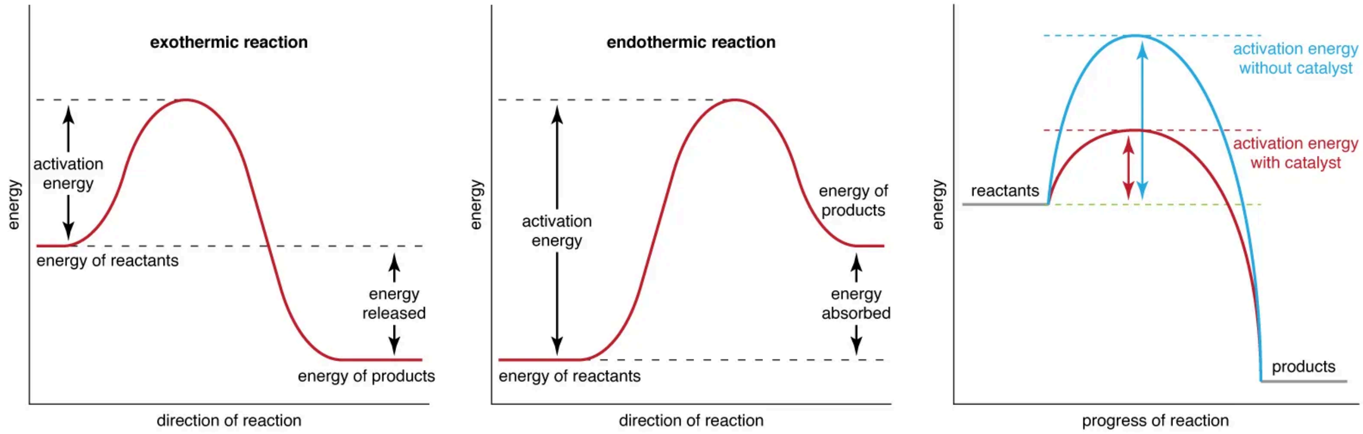
- **Sufficient Energy:** Colliding particles must possess enough kinetic energy to overcome the repulsion between their electrons and break existing chemical bonds.
- **Correct Orientation:** The particles must collide in a specific spatial alignment that allows the atoms to rearrange and form new bonds.



Activation Energy (E_a):

This is the **minimum amount of energy** required for a collision to be effective and lead to a reaction.

- **High E_a:** Fewer particles have this energy, so the reaction is slower.
- **Low E_a:** More particles have this energy, so the reaction is faster.

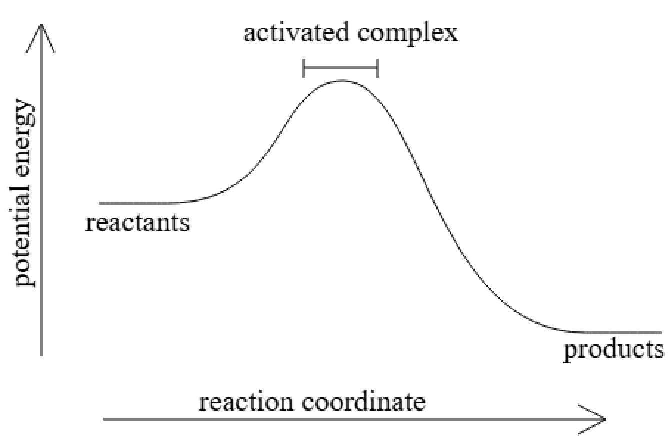


Activated Complex (Transition State):

- During an effective collision, particles form a temporary, high-energy, and unstable species called the **activated complex**.
- It is a transitional structure where old bonds are breaking and new bonds are beginning to form.
- It exists for a very short time before breaking apart to form the final products.

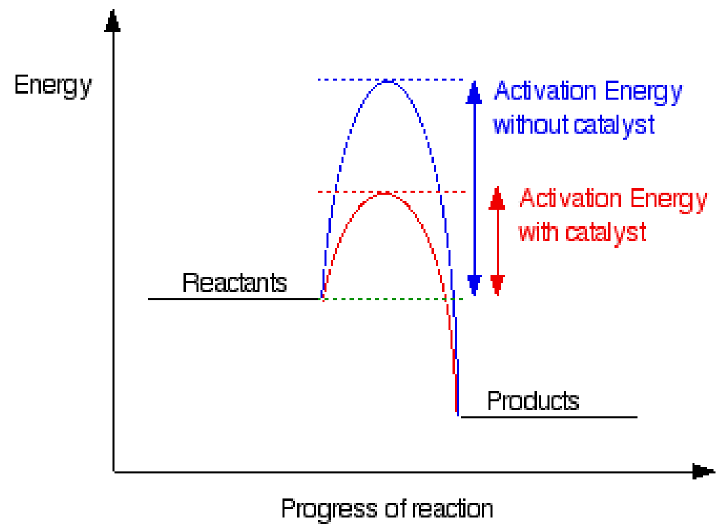
Example: $H_2 + Cl_2 \rightarrow 2HCl$

1. **Reactants:** H_2 and Cl_2 molecules collide effectively.
2. **Activated Complex:** A high-energy arrangement (e.g., H--Cl--Cl) forms momentarily.
3. **Products:** The complex breaks down to form two stable HCl molecules



5.3. Catalysts and their Role in Reaction Kinetics

A catalyst is a substance that speeds up a chemical reaction without being consumed or used up in the process.It provides an alternative pathway for the reaction that has a lower activation energy.



How Does a Catalyst Work? (The "Hill" Analogy)

- **Without a Catalyst:** Reacting particles must overcome a high energy "hill" (the original activation energy). Few particles have enough energy, so the reaction is slow.
- **With a Catalyst:** The catalyst provides a different, **lower hill** (lower activation energy). This allows **more particles** to have the required energy to react, leading to a faster reaction.

Properties of Catalysts

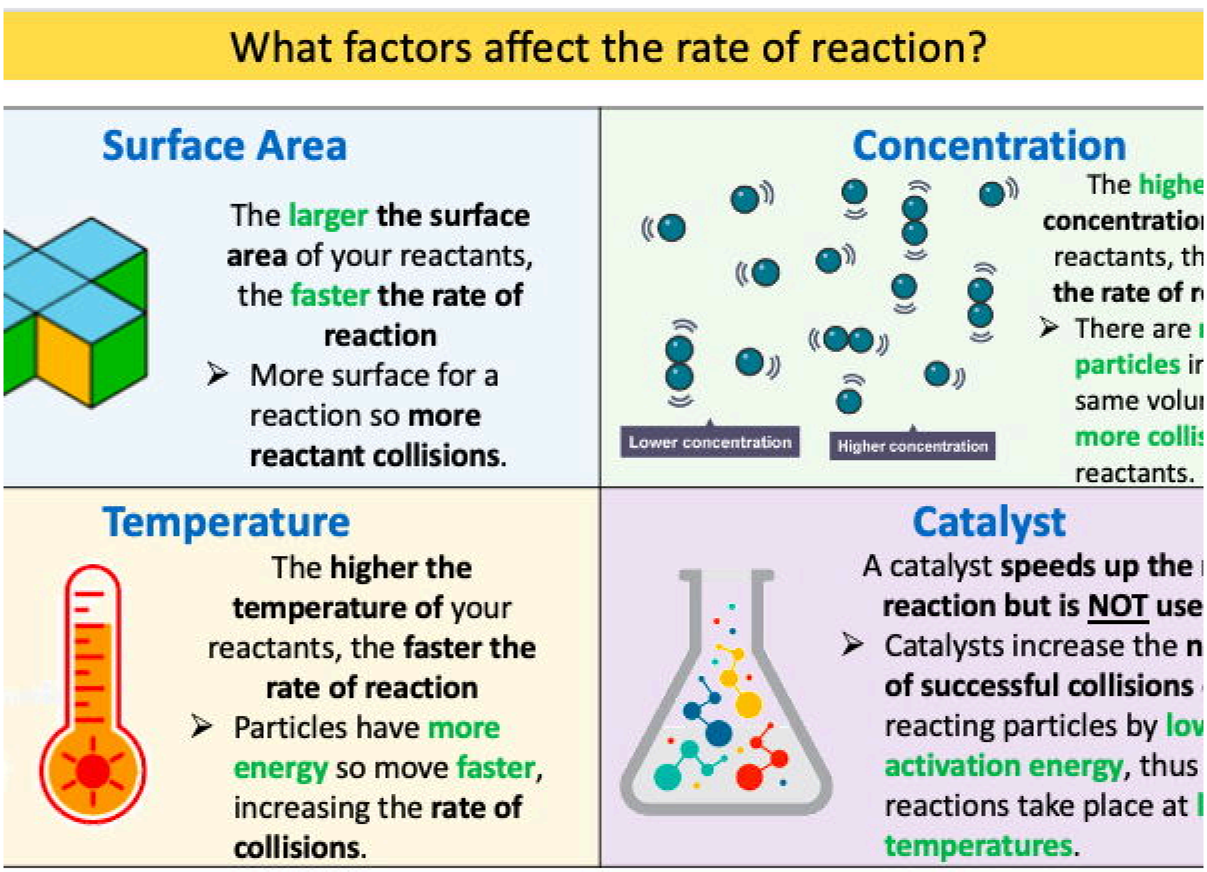
- **Not Consumed:** A catalyst remains **unchanged in mass and composition** at the end of the reaction.
- **Does Not Initiate Reactions:** A catalyst **cannot** make a non-spontaneous reaction happen. It only speeds up a reaction that is already feasible.
- **No Overall Energy Change:** A catalyst does not alter the **total energy change** (ΔH) of the reaction between reactants and products.

5.3.1. Physical Parameters that Affect the Rate of Reaction

Several measurable physical changes indicate that a reaction is occurring and can be used to track its rate.

1. Change in Mass

As reactants are used up, their **mass decreases**. As products are formed, their **mass increases**. Monitoring mass change over time is a direct way to measure the reaction rate.



2. Formation of a Gas

In an **open system**, if a gas is produced and escapes, the **mass of the reaction mixture will decrease**. In a **closed system**, the gas cannot escape. As more gas is produced, the **number of gas particles increases**, leading to an **increase in pressure** (if volume is constant). This pressure increase can be measured to track the rate.

3. Temperature

Increasing the temperature increases the reaction rate. At higher temperatures, particles have **more kinetic energy**. They move faster, leading to more frequent collisions. A greater proportion of particles have energy equal to or greater than the activation energy, leading to more successful collision

5.3.2. Factors Affecting Rate of Reactions

Several factors influence the reaction rate by affecting the frequency and **effectiveness** of collisions between particles.

1. Concentration of Reactants

Higher concentration leads to a faster reaction rate. More reactant particles in a given volume increase the frequency of collisions.

Examples:

- Stronger (higher H^+ concentration) in acid rain damages marble faster.
- Two antacid tablets neutralize acid faster than one because more reacting particles are present.
- Doubling the concentration of H_2 or Cl_2 gas doubles the reaction rate.

2. Surface Area (for Solids)

A larger surface area leads to a **faster reaction rate**. Breaking a solid into smaller pieces or powder exposes more particles to the other reactant, increasing the **number of sites where collisions can occur**.

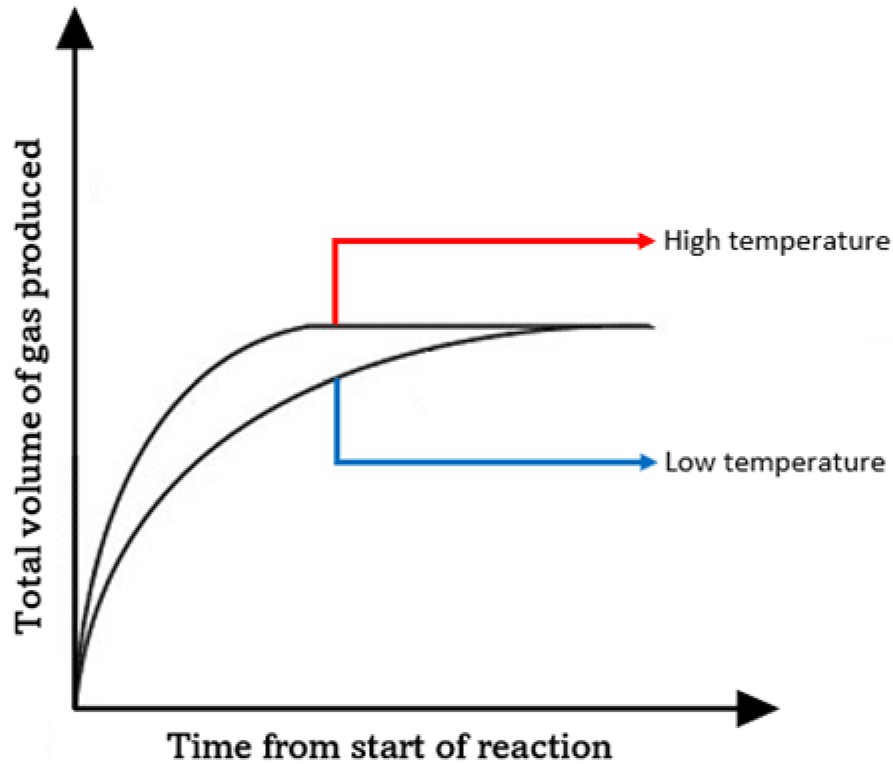
Examples:

- Zinc powder reacts with dilute HCl faster than a zinc lump.
- Powdered aluminum reacts quickly with NaOH, while aluminum foil reacts slowly.

3. Temperature

Higher temperature leads to a **much faster reaction rate**.

- **More Frequent Collisions:** Particles move faster, leading to more collisions per second.
- **More Energetic Collisions:** A greater proportion of particles have kinetic energy equal to or greater than the **activation energy (E_a)**.



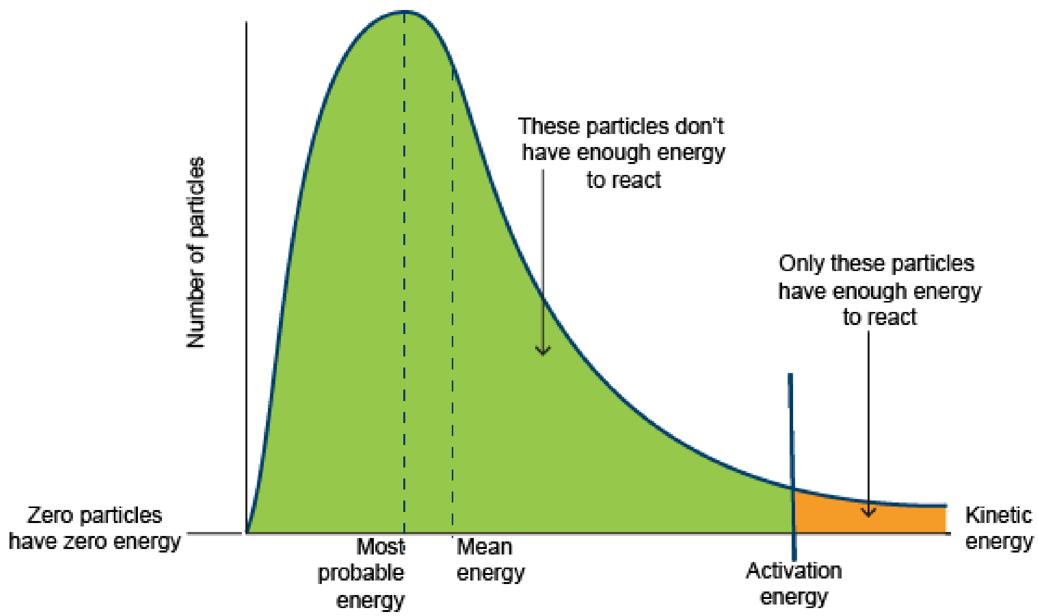
4. Pressure (for Gaseous Reactions)

Higher pressure leads to a **faster reaction rate**.Increasing the pressure of a gas is equivalent to increasing its **concentration** in a given volume. This forces gas particles closer together, resulting in more frequent collisions.

Example: Doubling the partial pressure of H_2 or Cl_2 in their reaction mixture doubles the reaction rate.

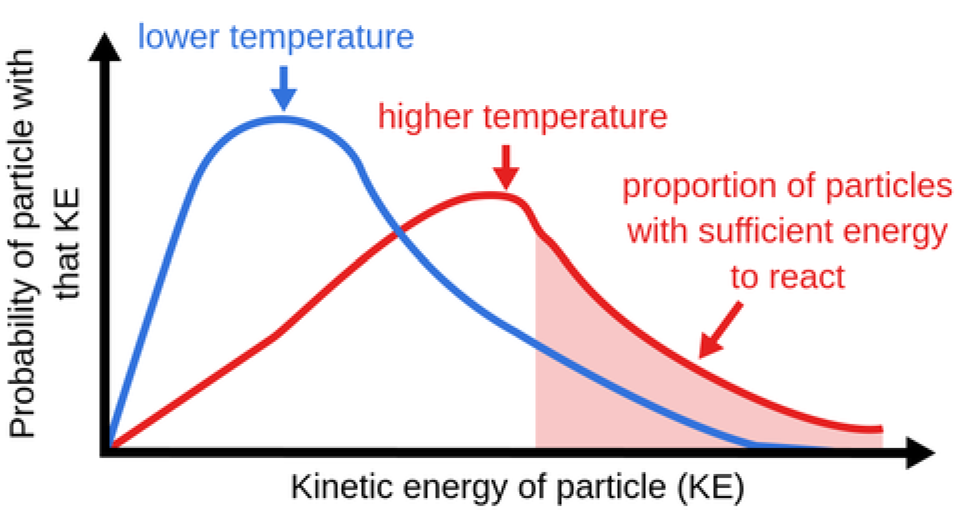
Maxwell-Boltzmann Energy Distribution

This concept explains *why* temperature has such a significant effect on reaction rate. The Maxwell-Boltzmann curve shows the distribution of kinetic energies among particles in a gas at a specific temperature. No particles have zero energy. Only a **small fraction** of particles have energy greater than or equal to the activation energy (E_a) at a given temperature (T_1).



Effect of Increasing Temperature (to T_2):

The entire curve shifts to the right and flattens, meaning the **average particle energy increases**. Crucially, the area under the curve beyond the activation energy (E_a) **significantly increases**. This means a **much larger proportion of particles** now possess the required energy to react upon collision.



Conclusion: An increase in temperature doesn't just cause more collisions; it dramatically increases the number of **effective collisions** (those with energy ≥ E_a), which is the primary reason for the faster reaction rate.

5.3.4. Enzymes

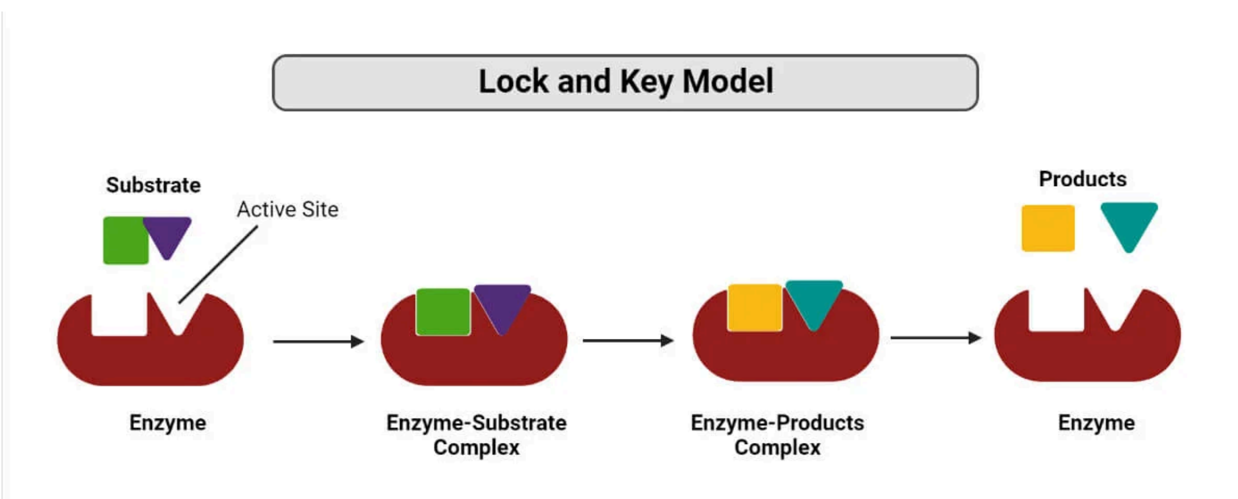
Enzymes are **biochemical catalysts** that regulate the vast majority of chemical reactions within living organisms. **Nature:** They are specialized **proteins**.

Properties:

- **Not Consumed:** Like catalysts, they are not used up in the reactions they facilitate.
- **Highly Specific:** Each enzyme catalyzes **only one specific reaction**.
- **Extremely Efficient:** They can speed up reactions by a factor as high as $\mathbf{10^{20}}$.
- **Location:** Found inside cells and in extracellular fluids (e.g., saliva, gastric juice).

How They Work (Lock and Key Model):

Enzymes have a specific three-dimensional shape. They bind to reactant molecules (substrates) and hold them in the precise orientation required for a successful collision. This precise binding dramatically **lowers the activation energy** and increases the reaction rate.

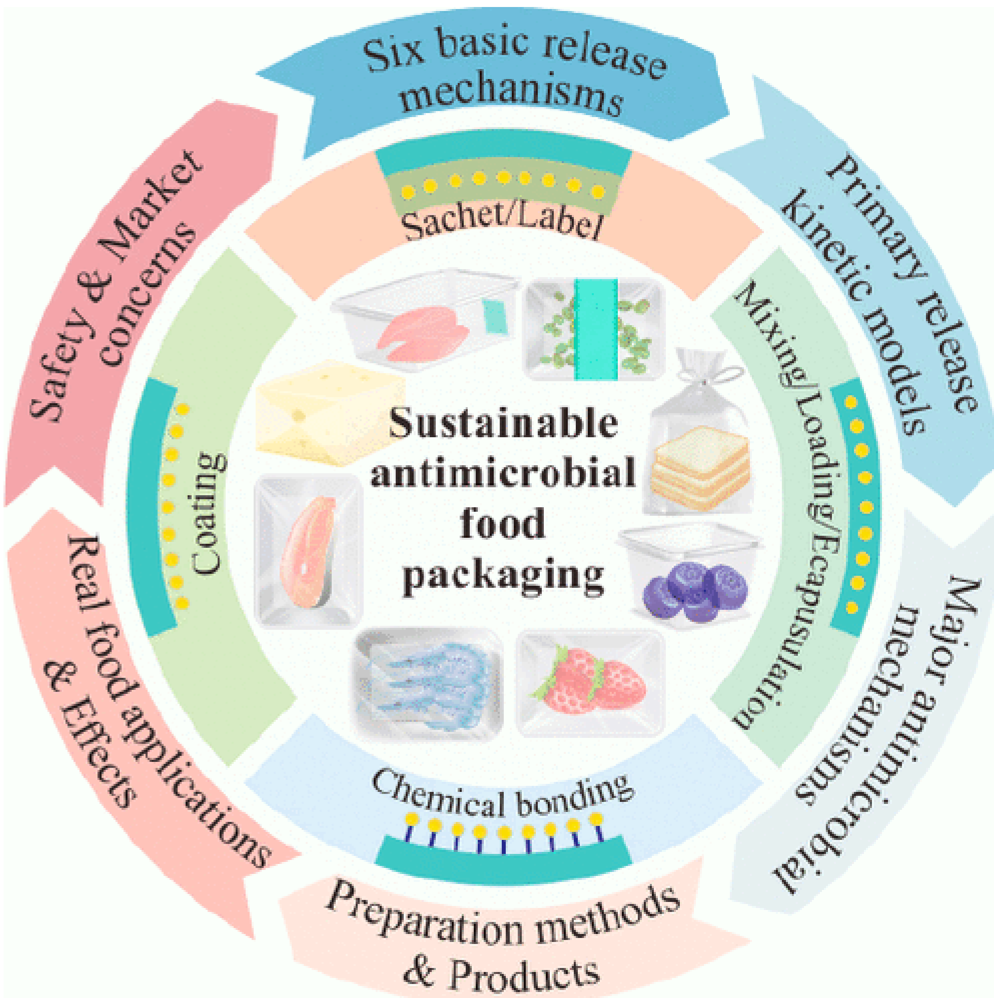


5.4. Role of Chemical Kinetics in the Food Industry

Chemical kinetics is applied in the food industry to control quality, minimize waste, and extend shelf life. Key applications include:

Optimizing Harvest and Transport: Determining the best time to harvest and transport produce to ensure it arrives with optimal taste, texture, and nutritional value.

Minimizing Losses: Estimating harvest time so products reach the market at peak quality, reducing losses from over-ripening during transit.



Understanding Degradation: Identifying factors (like oxidation) that cause food spoilage during transportation and storage.

Improving Storage Methods: Developing storage and transportation conditions (e.g., controlled atmospheres) that preserve nutritional content.

Shelf Life: Using kinetics to find methods (like refrigeration or modified packaging) that **slow down degradation reactions** by controlling temperature and humidity.