# Explaining Reaction Rates

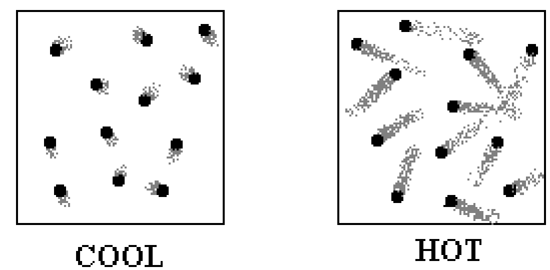
Chapter 6.3

# Collision Theory

- **Collision theory** is the theory that chemical reactions can occur only if reactants collide with proper orientation and with enough kinetic energy to break reactant bonds and form product bonds

## Concepts of the Collision Theory

- A chemical system consists of *particles* (atoms, ions, or molecules) that are in constant random motion at various speeds. The average kinetic energy of the particles is proportional to the temperature of the sample. **Figure 1** shows the distribution of kinetic energies among particles in a sample at two different temperatures.




Kinetic Energy Distribution at Two Temperatures

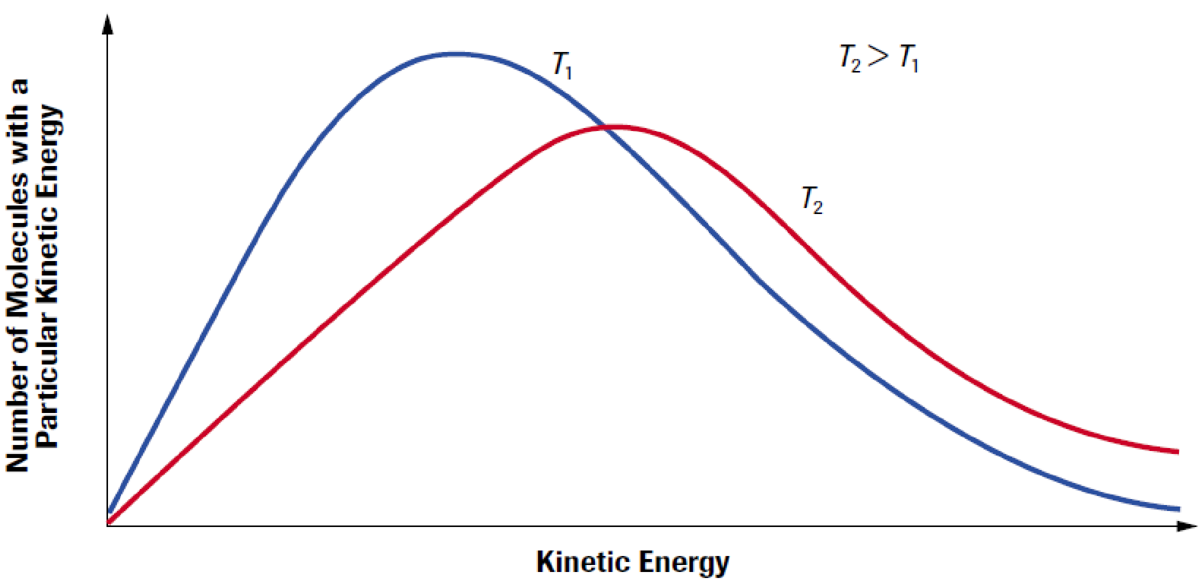
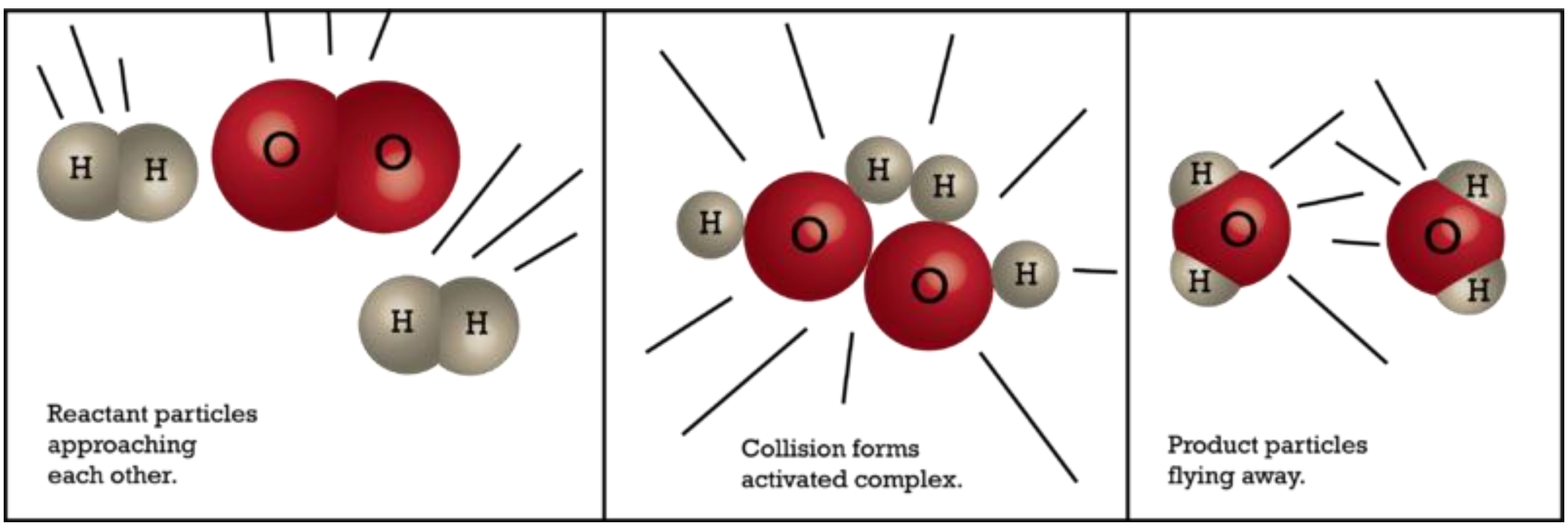



**Figure 1**
Temperature is a measure of the average kinetic energy of the particles. This graph shows how the distribution of kinetic energies changes when a substance is heated or cooled. At any temperature there are some particles with low kinetic energy and some with high kinetic energy. The higher the temperature, the more particles there are with higher kinetic energies.

- A chemical reaction must involve *collisions of particles* with each other or the walls of the container.

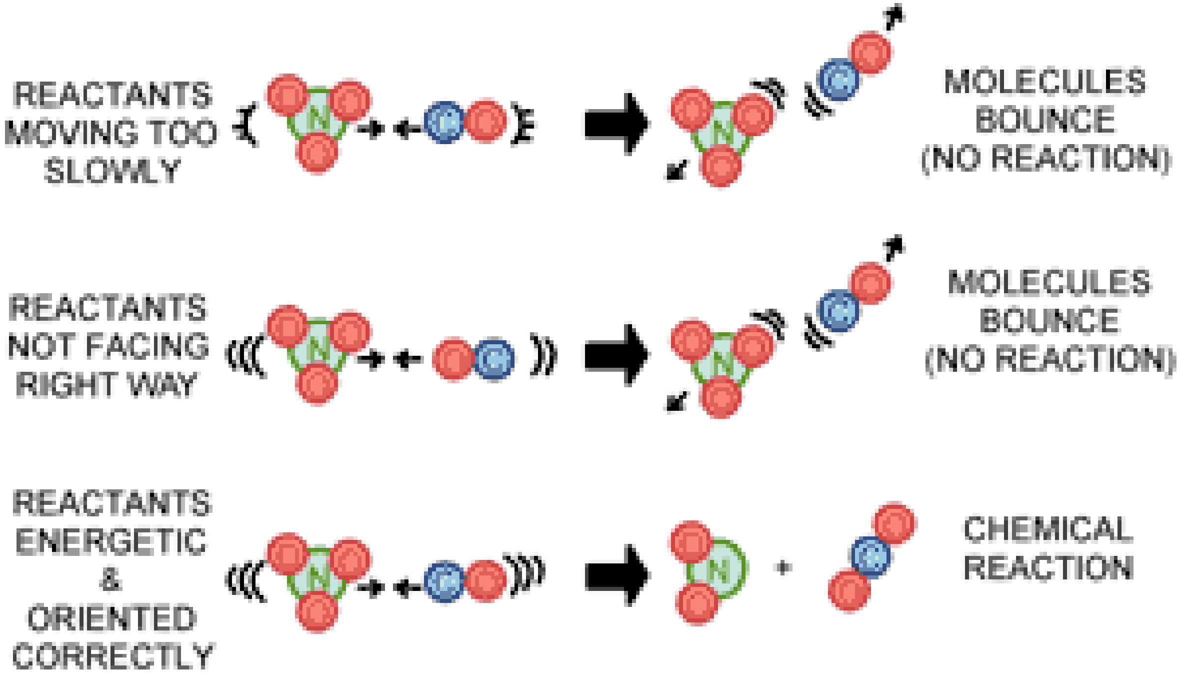
$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$



# Collisions

- Statistics tells us that…
  - In gases there is an average of $10^{30}$ collisions per second
  - If every collision resulted in a reactant molecule forming a product molecule, the rate of chemical reaction should be around $10^6$ M/s
- However…
  - Actual reaction rates for gases are on the order of $10^4$ M/s
- What does this mean?

- An *effective collision* is one that has sufficient energy and correct orientation (alignment or positioning) of the colliding particles so that bonds can be broken and new bonds formed.
- *Ineffective collisions* involve particles that rebound from the collision, essentially unchanged in nature.



# Activation Energy

- **Activation Energy ($E_a$)** is the minimum energy that reactant molecules must possess for a collision to be effective

An Analogy for Activation Energy

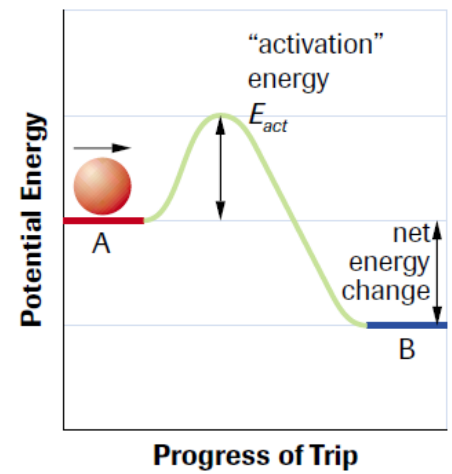
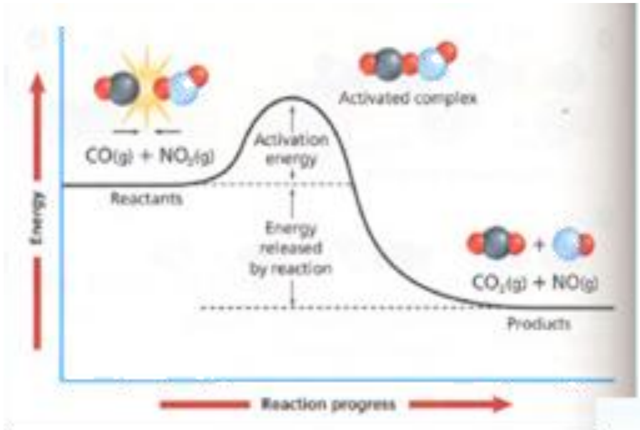


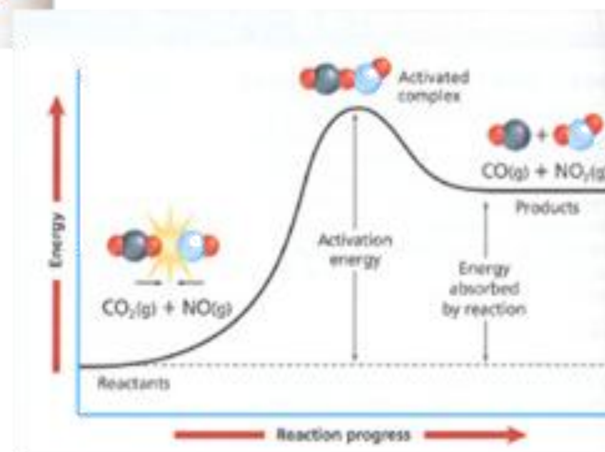

**Figure 3**
On a trip from A to B there is a net decrease in overall (net) energy, but there must be an initial increase in potential energy (activation energy) for the trip to be possible.

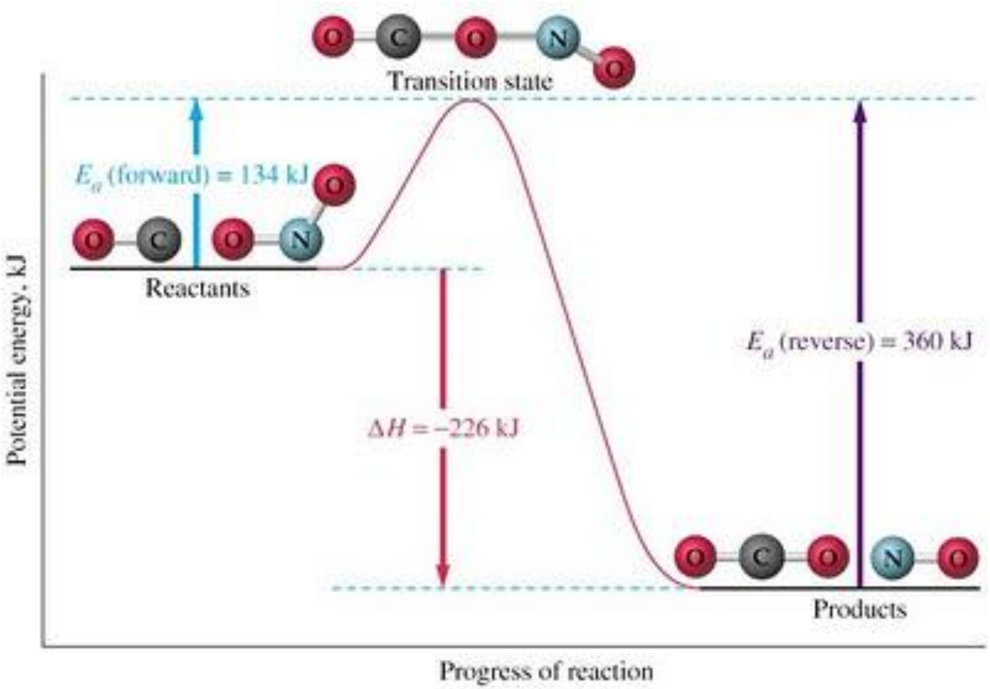
An **activated complex** or **transition state** is an unstable arrangement of atoms containing partially formed and partially broken bonds that represents the maximum potential energy point in the change




*Exothermic*

*Endothermic*



- The rate of a given reaction depends on the *frequency* of collisions and the *fraction* of those collisions that are effective.

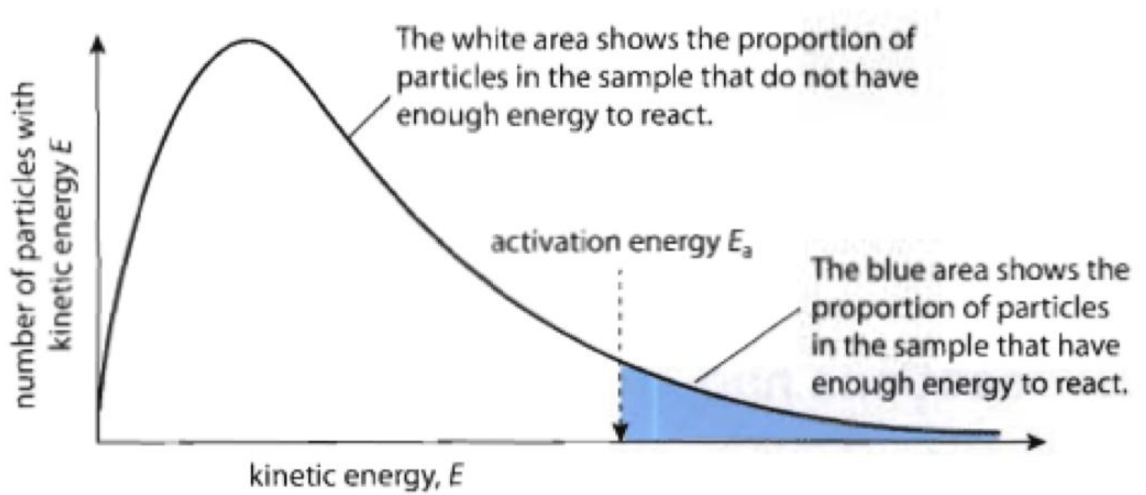
$$\text{rate} = \text{frequency of collisions} \times \text{fraction of collisions that are effective}$$

Increasing either the collision frequency or the fraction of effective collisions will increase the reaction rate

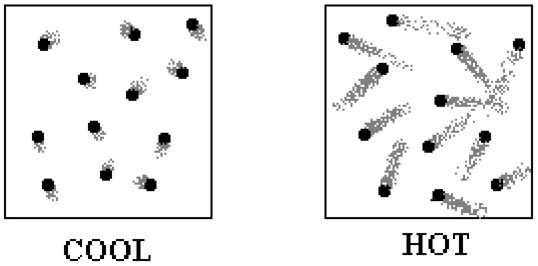
# Temperature

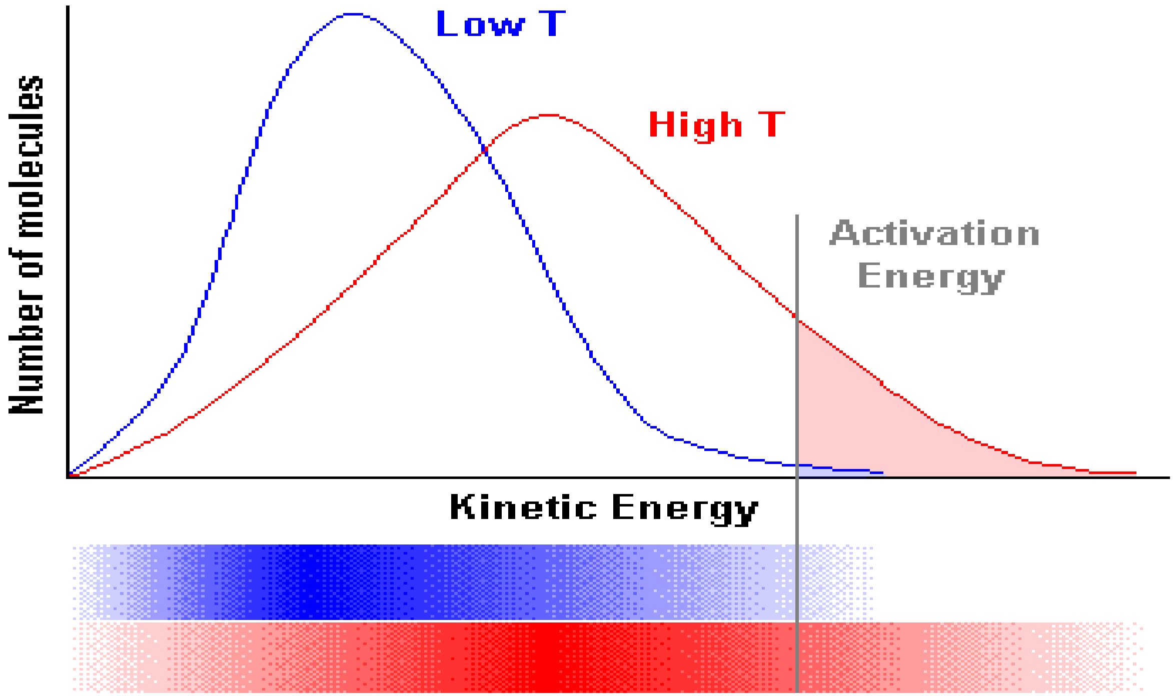
- Increased temperature increases the rate of reaction by increasing both the collision frequency and the fraction of effective collisions

Maxwell-Boltzmann Distribution



# Temperature



# Catalysts

- Catalysts increase the rate of reaction by providing an alternate pathway for the reaction with a lower activation energy

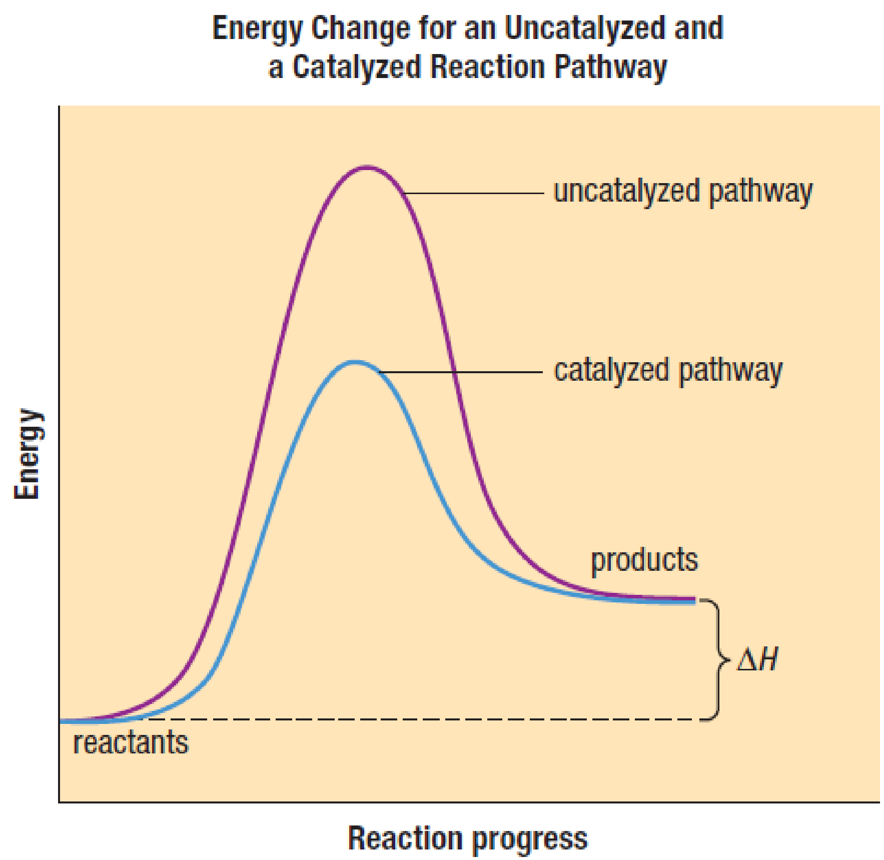



**Figure 10** Energy plots for a catalyzed and an uncatalyzed pathway for a given reaction

# Catalysts

- The lowered activation energy increases the fraction of collisions that are effective

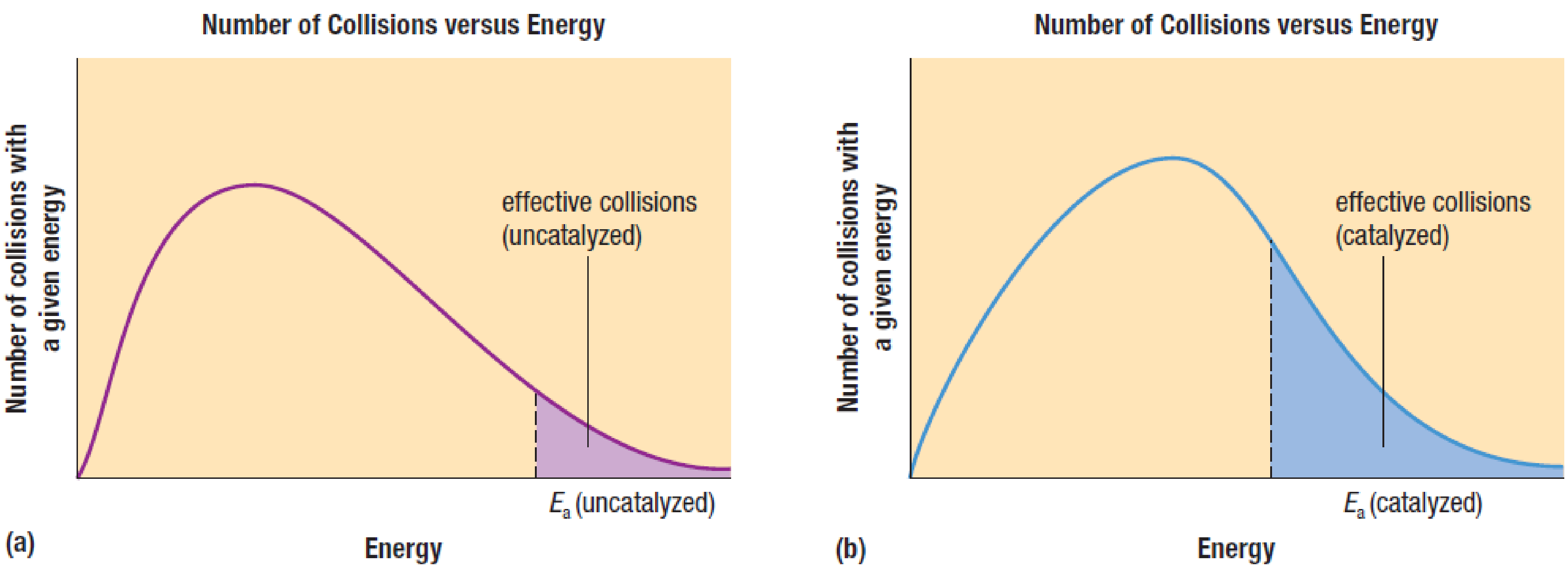
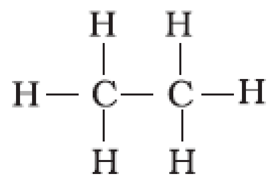



**Figure 9** The effect of a catalyst on the number of reaction-producing collisions. Since a catalyst provides a reaction pathway with a lower activation energy, a much greater fraction of the collisions are successful for the catalyzed pathway (b) than for the uncatalyzed pathway (a) (at a given temperature). This allows reactants to become products at a much higher rate, even if the temperature is not increased.

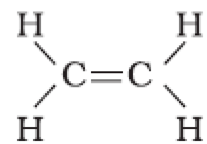
# Chemical Nature of Reactants

- For any reactant, the activation energy required for a successful collision depends on:

- the bond type



*Ethane*

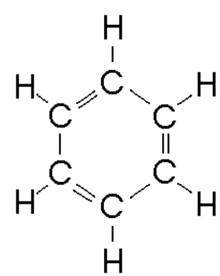


*Ethene*

- the bond strength

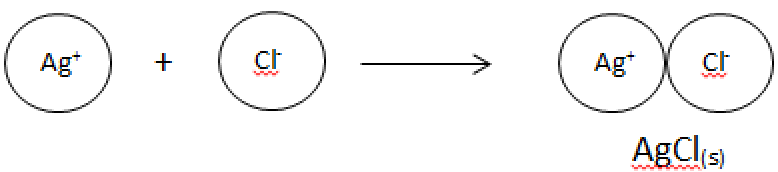
| | |
|---|---|
| H–F | 565 |
| H–Cl | 427 |

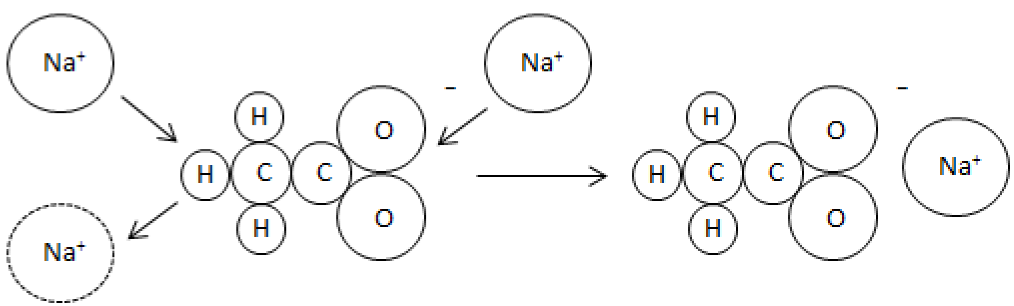
- the  number of bonds



# Chemical Nature of Reactants

- The size and shape of a reactant molecule can also affect the collision orientation



# Concentration

- Increased concentration increases the rate of reaction by increasing the collision frequency

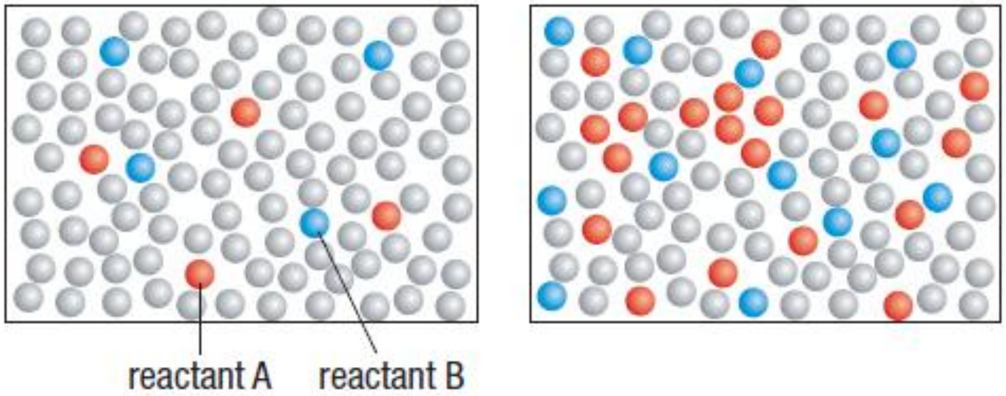
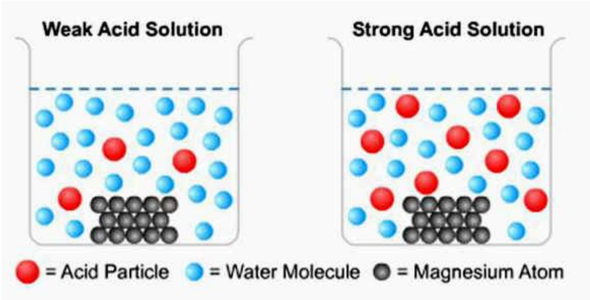



**Figure 7** Two reactions occurring in aqueous solution, one with lower concentrations of the reactants and the other with higher concentrations. Identify the number of reactant A–reactant B collisions in each sample.



# Surface Area

- Increased surface area also increases the rate of reaction by increasing the collision frequency

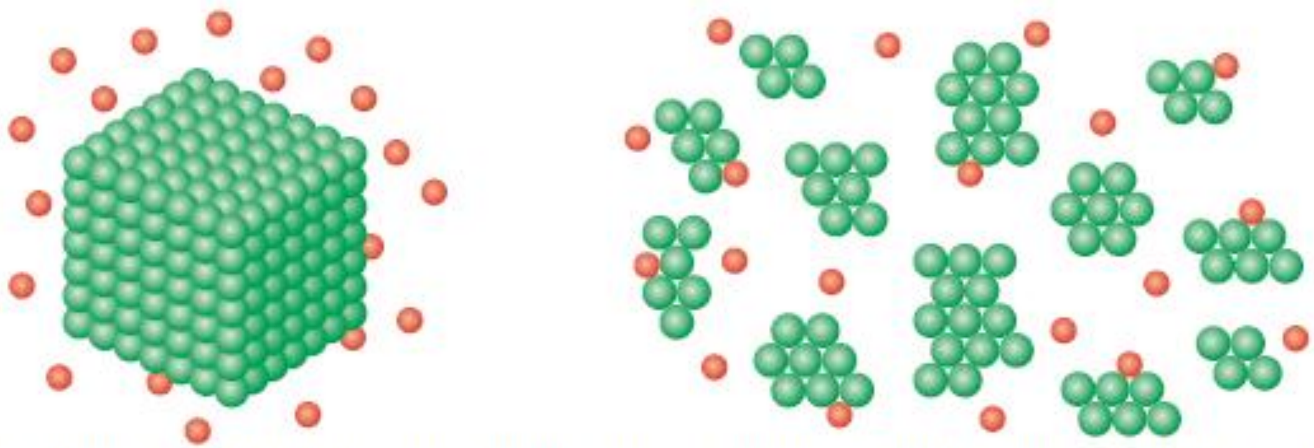
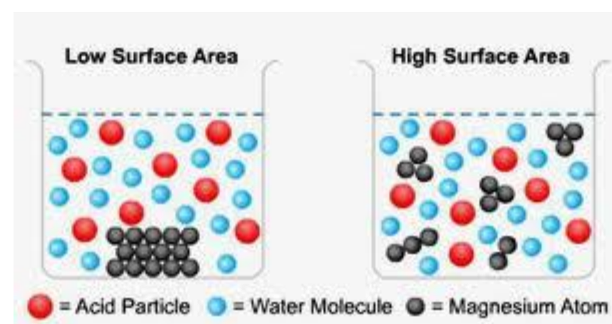


**Figure 8** The entities in a solid structure have fewer potential collision sites than the same number of entities split into smaller bits, increasing the total surface area.



# Summary

$$\text{rate} = \text{frequency of collisions} \times \text{fraction of collisions that are effective}$$

# HOMEWORK

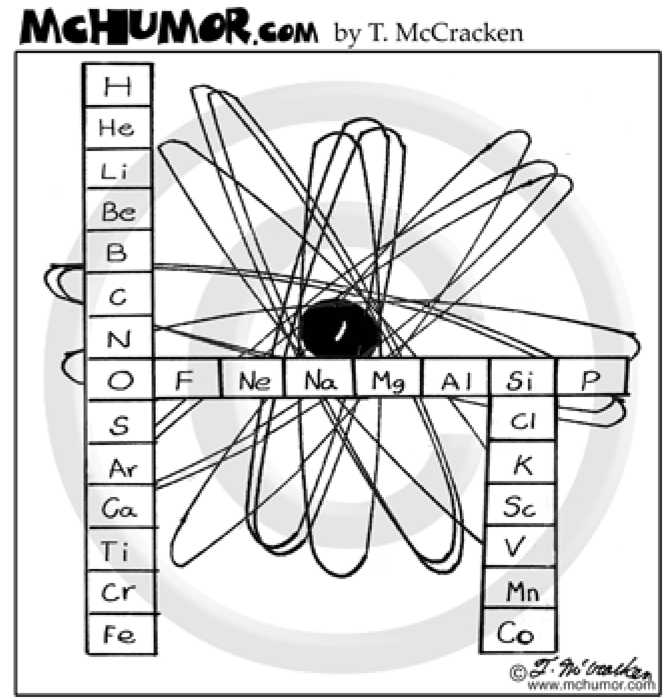
Required Reading:

p. 366-372

(remember to supplement your notes!)

Questions:

p. 372 #1-6




Ideas that never caught on:
The Periodic Chair Of The Elements.