Collision Theory: A Model that Explains Reaction Rates

Collision Theory is based on the reality that, for a given reaction to occur, molecules **MUST** Collide !

The rate of a reaction is directly proportional to the rate of reactant collisions.

Two conditions must be met for an effective collision: (i.e., collisions of reactants that result in products 😉)

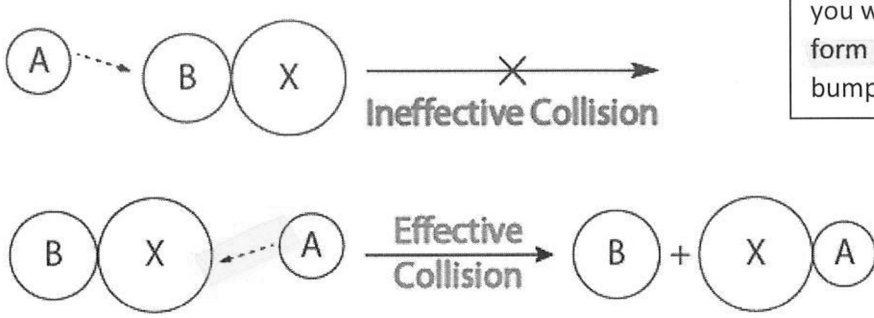
1. Correct orientation: atoms that will form new bonds must bump into each other.
2. Sufficient energy: atoms that will form new bonds must collide with enough energy to overcome the activation energy of the reaction.

Let's take a look at both of these conditions.

Correct orientation: The reacting species must collide in an orientation that allows contact between the atoms that will become bonded together in the product. ~~This can be described in two ways:~~

- Sometimes known as the orientation factor : represents the fraction of collisions that occur with an orientation that allows the reaction to occur.
- $\uparrow$ orientation factor = $\uparrow$ probability of an effective collision

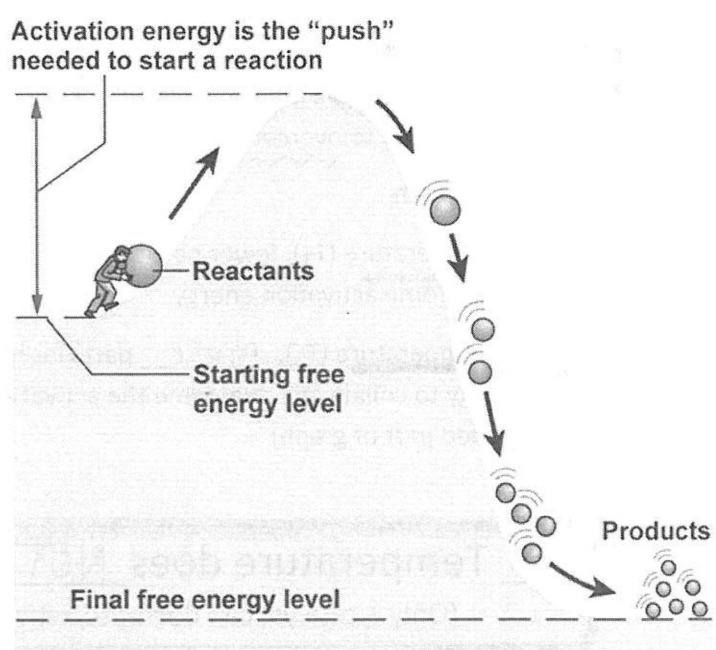
$$\boxed{A + BX \longrightarrow B + AX}$$



For an effective collision, you want the atoms that will form new bonds to bump into each other!

Sufficient energy: The collision must occur with enough energy to overcome the pre-existing attractions in the reactant molecules between each bonded atom and their shared electrons, breaking the preexisting bonds and forming new bonds.

- Must overcome the activation energy!



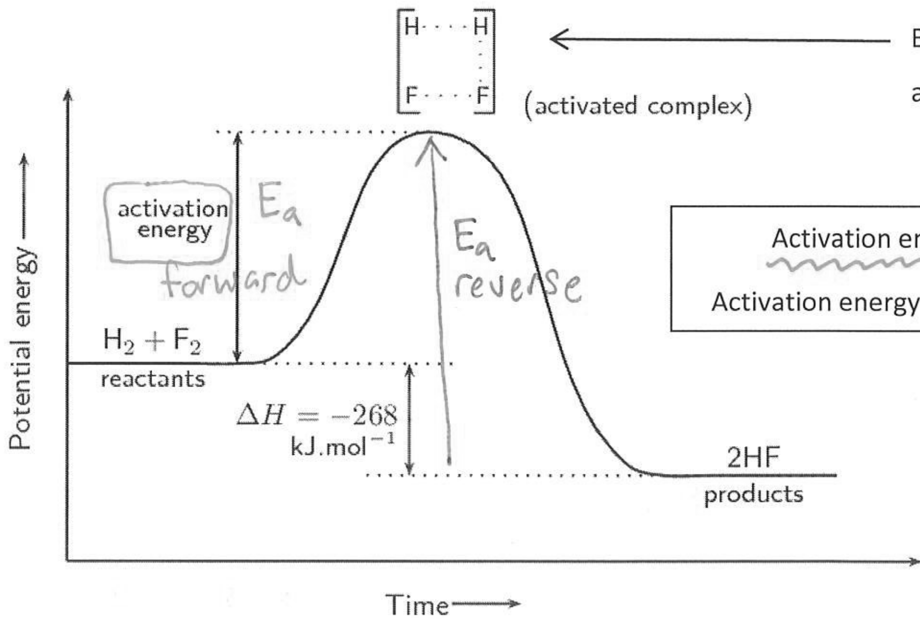
The Reaction Pathway

Activated energy (Ea): the energy barrier (or hump) that must be overcome for reactants to convert into products.

IMPORTANT: Higher activation energy = Slower reaction rate!

Activated complex (or transition state): the high energy transient state that is the collision product of the reactants, with some bonds partially broken and some bonds partially formed. The activated complex can either revert to reactants or proceed to products.

Bonds that are **breaking or forming** are shown as dotted lines!



Activation energy of forward reaction = E_a

Activation energy of reverse reaction = $E_a(fwd) - \Delta H$

Effect of Temperature on Effective Collisions

As temperature increases, by definition the average Kinetic energy of the particles also increases.

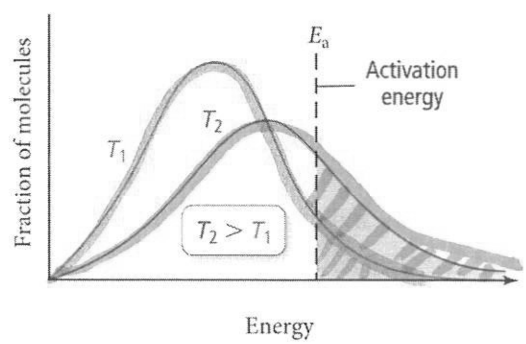
Heat 'em up = Speed 'em up!

Higher temperature = Higher (faster) reaction rate because:

Faster particles =

1. More total collisions → on average, collide with more force
2. Higher energy collisions: a greater fraction of collisions will have enough energy to overcome the activation energy

As temperature increases, the fraction of molecules with enough energy to surmount the activation energy barrier also increases.



As you can see in the graph,

- At a lower temperature (T_1), fewer particles have enough energy to overcome activation energy
- At a higher temperature (T_2), more particles have enough energy to collide and overcome the activation energy (shaded part of graph)

Please note:

↑ Temperature does NOT decrease activation energy!!

(Only a catalyst can decrease activation energy: more on that shortly!)

Which is changing: the rate of reaction, or the rate constant, k?

Only two changes will actually affect the rate constant, k, but lots of things can affect the rate of reaction!

- The **rate constant**, *k*, is the proportionality constant related to the rate of a particular reaction.

Conditions that can change the **Rate Constant, *k***	
Increase *k*	**Decrease *k***
High temperature	Low temperature
Lower activation energy (catalyst)	Higher activation energy (no catalyst)

- The **rate of a reaction** is the change in concentration of reactants or products per unit time.

Conditions that can change the **Reaction Rate**	
High Reaction Rates	**Low Reaction Rates**
High temperature	Low temperature
Low activation energy	High activation energy
Weak reactant bonds	Strong reactant bonds
High orientation factor	Low orientation factor
High reactant concentration	Low reactant concentration

Collision Theory in a Nutshell

Collision theory states that:

1. Particles must collide
2. with sufficient energy
3. in the correct orientation

➔ for an effective collision to occur (where reactants actually turn into products)

Collision theory explains why certain conditions increase the rate of reaction:

1. Increased concentration of reactant particles: More particles means more collisions, and thus more opportunity to collide with correct orientation → faster reaction rate!
2. Increased temperature (i.e. increased kinetic energy): Particles move faster and collide more often, which means greater chance of colliding with effective orientation AND they move faster and collide with greater force, thus having a better chance of overcoming the activation energy → faster reaction rate!
3. Increased pressure for gases: Particles are closer together and therefore more likely to collide, and thus more opportunity to collide with correct orientation → faster reaction rate!