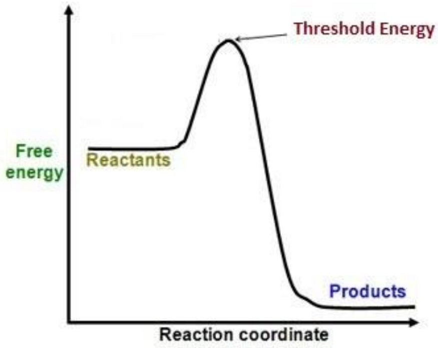
Collision Theory

1. If two molecules are to react together, they must collide together.

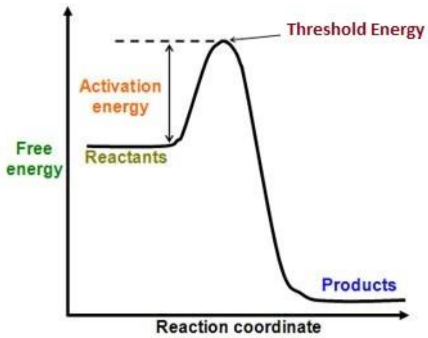
2. **Threshold Energy:** All collisions do not lead to chemical reactions: only those collisions give rise to chemical reaction in which the molecule acquires energy **level greater than or equal the threshold energy.**



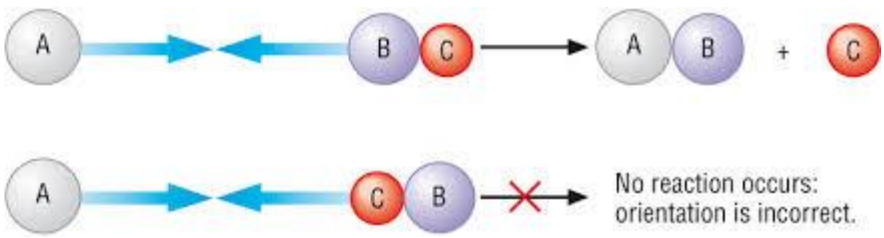
3. Thus, only those collisions result in product formation in which the colliding molecules cross the minimum energy level barrier i.e. threshold energy level. The collisions which result in the formation of product are called effective collisions. Collisions among molecules possessing energy less than threshold energy are not the effective collisions and thus do not result in the formation of products.

4. **Activation Energy:** Thus, colliding molecules must possess certain minimum energy (Threshold energy) (E_T) to make the collision effective, but most molecules called normal molecules have lesser energy than the threshold energy. The additional energy required by the molecule to attain threshold energy is called activation energy, which is attained by the molecules because of interchange of energies during the collisions.

   Hence, **Activation Energy = threshold Energy − Energy of Colliding Molecules.**



5. **Orientation of Molecules**: The orientation of the colliding molecules partially determines whether a reaction between the two molecules will occur. If the collision does take place with the correct orientation, there is still no guarantee that the reaction will proceed to form carbon dioxide.



Conclusion: Collision theory explains why most reaction rates increase as concentrations increase. With an increase in the concentration of any reacting substance, the chances for collisions between molecules are increased because there are more molecules per unit of volume. More collisions mean a faster reaction rate, assuming the energy of the collisions is adequate.

The Activated Complex theory or Transition State Theory

This theory is based on the idea that bond breaking and bond making involved in a chemical reaction must occur continuously or simultaneously. For example, reaction between one hydrogen and one iodine molecule to form 2 hydrogen iodide molecules.

$$H_2 + I_2 \longrightarrow 2HI$$

At some state in the process the H−H and I−I bonds must be ruptured, while the H−I bonds are being established. Thus, if we represent a partially ruptured or established bond by a dotted line, we can write,

H−H	+	H– – –H	$\longrightarrow$	H + H
I – I		I – – –I		I I
Reactants		Activated complex		Products

The intermediate product with partially formed bond is called **activated complex or transition state,** and the energy of activation is the energy required to form the activated complex or intermediate. **The energy of the activated complex will be higher than that of the reactants and products.** Hence, the reactants are not converted directly into the products. There is an energy barrier or activated complex between the reactants and products. The reactants must cross this energy barrier before converting into products. **The height of the barrier determines the threshold energy.**