THE COLLISION THEORY

- In order for a reaction to take place, a collision between the reacting particles must take place
- Anything that can be done to increase the rate of collisions (number of collisions per second) will increase the rate of reaction
- An effective collision must have:
    - sufficient energy called activation energy
        - the minimum amount of energy with which particles must collide before a reaction will take place
        - only particles with energy ≥ Ea are able to react

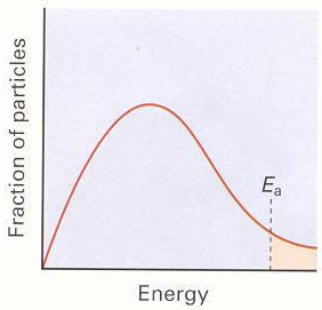

Figure 6.10 The area under a Maxwell-Boltzmann distribution graph represents the distribution of the kinetic energy of particles at a constant temperature. At a given temperature, only a certain fraction of the molecules in a sample have enough kinetic energy to react.

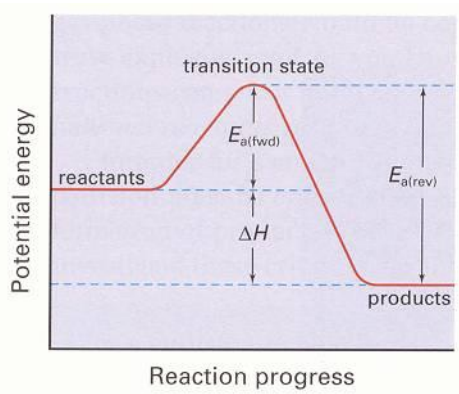

Figure 6.12 A potential energy diagram for an exothermic reaction

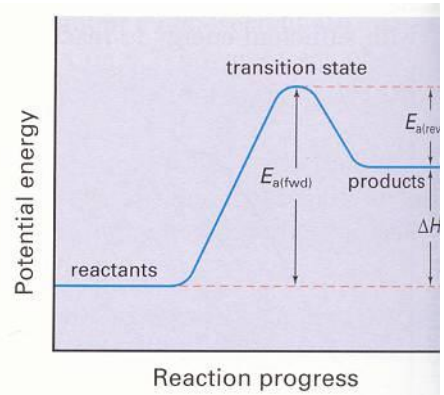

Figure 6.13 A potential energy diagram for an endothermic reaction

- the top of the activation "hill" on a potential energy graph represents the ***transition state***

- the substance that is found at the transition state is called the ***activated complex***
    - it has partially broken and partially made bonds and is highly unstable (doesn't last long)

    - a correct orientation of the colliding particles

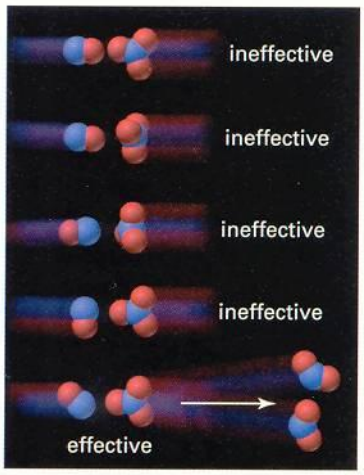

Figure 6.9 Only one of these five possible orientations of NO and NO_3 will lead to the formation of a product.

<u>The Effect of Concentration</u>

- If there are move particles in a given amount of space, there will be more collisions per unit time
- Of these collisions there will be a higher percentage with the right amount of energy and orientation to result in a reaction
- This can be seen by drawing a Maxwell-Boltzmann graph and then drawing a larger graph each with the same activation energy
- There are more particles above Ea in the larger graph

<u>The Effect of Temperature</u>

- If the temperature rises, the particles will be moving faster
- This will result in more collisions per unit time
- Since the particles are moving faster, they will collide harder (more energy involved in each collision)
- You can see in the figure below that at the higher temperature, there are more particles with
  energy ≥ Ea

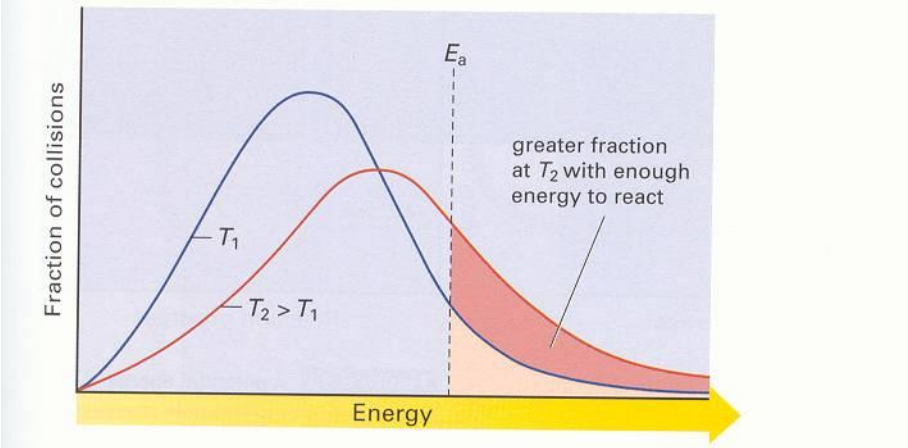

Figure 6.11 At increased temperatures, more particles have enough energy to react.

<u>The Effect of Surface Area</u>

- This factor only applies to heterogeneous reactions
- With greater surface area, more particles are available to collide with other particles therefore the number of collisions increase

<u>The Effect of Catalysts</u>

- Catalysts accelerate a reaction by lowering the activation energy of a reaction so that a larger fraction of the reactants have sufficient energy to react
- It lowers the Ea by providing an alternative pathway for the reaction
- They can be divided into 2 categories:
    - ***Homogeneous catalysts*** exist in the same phase as the reactants
    - ***Heterogeneous catalysts*** exist in a different phase from the phase of the reactants

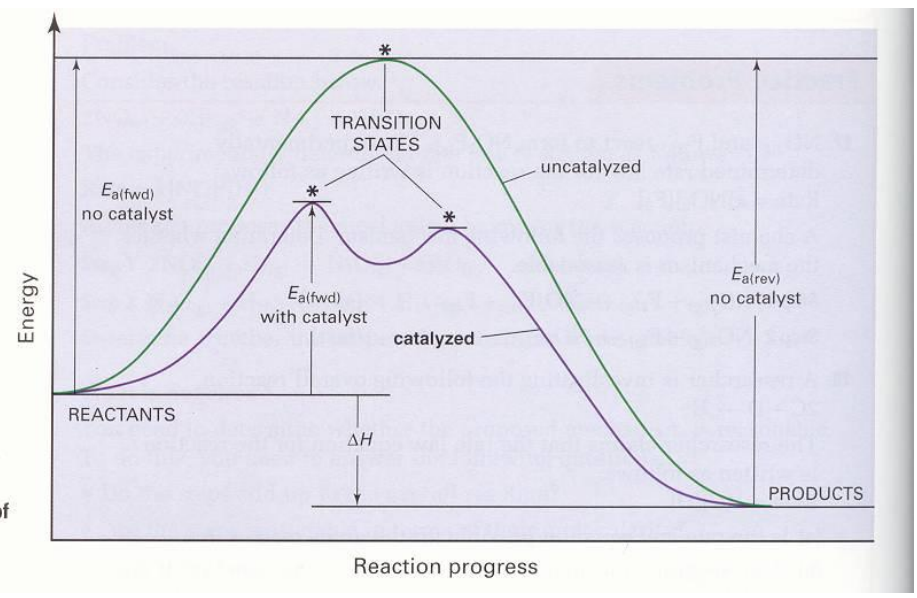

Figure 6.18 A catalyst lowers the activation energy of a reaction by providing an alternative mechanism. A catalyst also increases the rate of the reverse reaction. What effect does a catalyst have on ΔH of a reaction?