**Chemistry 30 notes – Chemical reactions pathways: Activation energy**

- **Reaction rate** is the speed at which a chemical reaction occurs. It is the change in the amounts of reactants consumed (or products produced) over time.

- Chemical reactions occur as a result of collisions between reactant molecules. However, not all collisions result in a reaction. For a collision to be successful (produce a product), two criteria must be satisfied:
  1. Correct orientation of reactants (collision geometry).
  2. The collision must contain enough energy to allow the bonding to change (remember that chemical bonds are stable – it takes energy to break them).

- The activation energy \( (E_a) \) of a reaction is the minimum amount of energy required for a successful reaction.

- Increasing the temperature of the reactants will increase the reaction rate, as a greater number of reactant molecules will collide with energy greater than the activation energy (remember that temperature is kinetic energy).

- Thus increasing temperature increases the force of the collisions, which increases the chances of a reaction occurring.

  ![Effective Collisions and Temperature](image)

- The activation energy is the minimum amount of collision energy required to force both reactant molecules together into a single entity called an **activated complex**.

- From the activated complex, bonding rearrangement occurs to form the products.

- The activated complex is a chemical entity that exists only at the top of the activation energy barrier \( (E_a) \). It is neither product nor reactant. It contains partial bonds and is highly unstable.

- Example:

  \[
  \text{Br-CH-H} + \text{OH}^+ \rightarrow \text{Br-CH-OH} \rightarrow \text{H-C-OH} + \text{Br}^+ \\
  \text{reactants} \quad \text{activated complex} \quad \text{products}
  \]
Even for exothermic reactions, activation energy must be supplied to start the reaction. The energy released by product bond formation supplies $E_a$ for the rest of the molecules present.

In the graphs above $\Delta H = E_{a\text{ forward}} - E_{a\text{ reverse}}$

In all chemical reactions, the species go through two transition states: one as they start to form the activated complex, the other as the activated complex turns into the products.

**Figure 11.5** A potential energy diagram for an exothermic reaction

**Figure 11.6** A potential energy diagram for an endothermic reaction

**Figure 11.8** As the reactants collide, chemical bonds break and form.
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- The rate of a chemical reaction depends on the magnitude of the activation energy ($E_a$).
- The greater the $E_a$ the slower the reaction.

**Catalysts**

- A **catalyst** is a substance that increases the rate of a chemical reaction without being consumed in the reaction.
- Catalysts provide an alternative reaction pathway with a lower activation energy.
- Catalysts do not change the overall enthalpy change ($\Delta H$).

**Figure 11.10** A catalyst provides an alternative pathway for a chemical reaction. The alternative pathway has a lower activation energy. A catalyst also increases the rate of the reverse reaction. What effect does a catalyst have on $\Delta H$?

- Different reactions require different catalysts. Examples:
  - Ammonia production: $N_2 (g) + 3 H_2 (g) \xrightarrow{Rh \text{ or } Pt} 2 NH_3 (g)$
  - Hydrogen peroxide decomposition: $2 H_2O_2 (aq) \xrightarrow{MnO_2} 2 H_2O (l) + O_2 (g)$
  - Automotive catalytic converter:
    - $NO_x (g) \xrightarrow{Pt \text{ or } Rh} N_2 (g)$ and $CO (g) \xrightarrow{Pt \text{ or } Rh \text{ or } Pd} CO_2 (g)$
- Biological catalysts are called **enzymes** which contain an active site that precisely fits the shape of the target molecule (called the substrate).

- Read pages 404-418
- Do problems p. 409 #1-5 (answers p. 733); p. 410 # 6-8; p. 418 #6