Every type of reaction follows a specific and consistent pathway with a unique activated complex. When a catalyst is introduced to the reaction, the pathway changes and a different activated complex is formed with a lower \( E_a \). Because of this lower \( E_a \), a higher % of collisions are effective, thereby increasing the reaction rate.

Catalysts are involved in creating a different, lower energy activated complex, but remain unaltered at the end of the reaction.

An inhibitor forms a new activated complex that has a larger activation energy, thereby decreasing reaction rate.

Catalysts lower the \( E_a \), thereby increasing the rate, but do they alter \( \Delta H \) for the reaction? (look at the curve on above)

A catalyst is written above the reaction arrow in a reaction. However, sometimes, it’s written as a reactant and a product (as it does not get used up in a reaction).

Example
Uncatalyzed (very slow):
\[
2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) + \text{energy}
\]

- Catalyzed (fast)

\[
\text{MnO}_2(s) + 2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) + \text{MnO}_2(s) + \text{energy}
\]

Draw a PE diagram for the uncatalyzed and catalyzed decomposition of \( \text{H}_2\text{O}_2 \) (hydrogen peroxide).
A heterogeneous catalyst is in a different phase than reactants, and usually adsorbs the reactant (holds it on its surface) to allow better geometry with other reactants.

A homogeneous catalyst is in the same phase as reactants, and usually alters the regular reaction pathway to a new, lower energy reaction pathway.