Basic knowledge

Catalytic activation

Many reactions are too slow for technical applications at ambient temperature because the required activation energy is very high. Catalysts lower the required activation energy and accelerate the chemical reaction. Thus, some reactions would not be possible without a catalyst reducing the energy required for production.

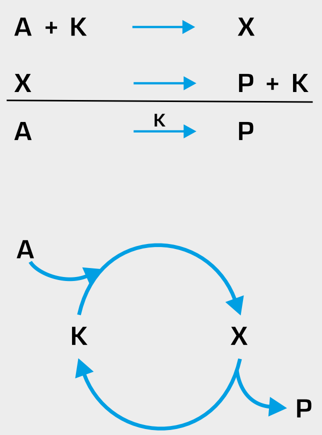
According to Wilhelm Ostwald, a catalyst is any substance that changes the speed of a reduction without appearing in the end product. Catalysis can be understood as the acceleration of a chemical reaction by means of a catalyst. Catalysts are used in approximately 80% of all industrial chemical processes.

In the simple case of the reaction of a reactant **A** to a product **P** by means of a catalyst **K**, one can imagine that the catalysis occurs via an intermediate product **X**. The reactant and the catalyst thus first form an intermediate product. In a second step, the catalyst is released and the intermediate product is converted to form the product **P**. The catalyst is unchanged after the reaction and is available again for further reactions.

One possible explanation of catalysis is the theory of the transition state. This theory assumes that the reactants involved in the reaction have to cross an energy barrier for the reaction to take place. The molecular state at the maximum of the energy barrier $\mathbf{E_1}$ is referred to as activated complex. The products form directly from this molecular state. During catalysis, the activated complex is formed from the reactants and the catalyst. The energy $\mathbf{E_2}$, which is required to form the complex with the catalyst, is lower than the energy $\mathbf{E_1}$ which would be required without the catalyst. This lower energy requirement means that a larger number of reactants react per time unit to form products, i.e. the reaction rate is higher.

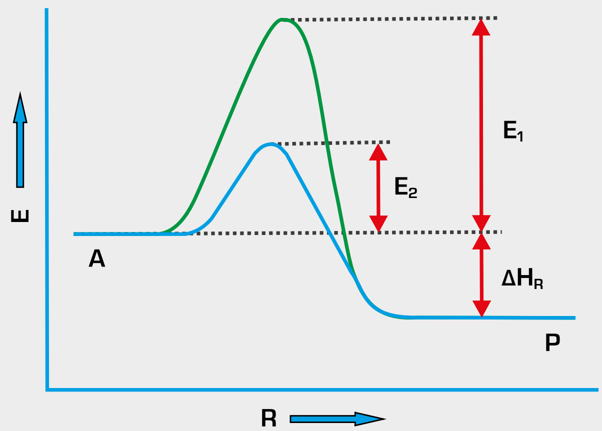
There are two types of catalysis:

- **Homogeneous catalysis**
  The catalyst and the starting substances of the chemical reaction are in the same phase. This means that the reaction takes place either in the liquid or in the gaseous phase. In the liquid phase, the properties of the solvent (e.g. viscosity) also influence the reaction rate in addition to the type of reactants and catalyst.
- **Heterogeneous catalysis**
  The catalyst is in the solid phase in most cases. The starting substances of the reaction are in the liquid or gaseous phase. In addition to the actual chemical reaction between reactants and catalyst, processes such as diffusion inside the solid catalyst and sorption processes have a significant influence on the reaction rate.



Reaction schematic of a simple catalytic reaction as a schematic (top) and cycle (bottom):

A reactant, **K** catalyst, **X** intermediate product,
P product



Energy change with and without catalyst (exothermic):

E energy, **R** reaction coordinate,
$\mathbf{E_1}$ energy required to form an activated complex without catalyst,
$\mathbf{E_2}$ energy required to form an activated complex with catalyst,
$\mathbf{\Delta H_R}$ reaction enthalpy