Catalysis

Catalysis is the processes of accelerating the chemical reaction using the catalyst. A catalyst is the substance that increases the rate of chemical reaction without actually taking part in the chemical reaction meaning remains without consumption during the chemical reaction. However, its physical appearance may change; solid when added at the beginning of the reaction may turn into powder at the end of the reaction. In addition, it may react to form an intermediate substance. When that happens, it will subsequently regenerate without any loss.

I call the catalyst as the "Real Estate Agent," who can speed up your home buying process; otherwise the process could be slow if you do it on your own.

In the laboratory, the oxygen gas is prepared by heating the potassium chlorate according to the following equation.

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

However, this thermal decomposition is very slow, but the rate of chemical reaction (liberation of oxygen gas) can be increased substantially by adding the small amount manganese dioxide ($MnO_2(s)$), a catalyst. Note that no MnO_2 is consumed in the reaction. And all the MnO_2 can be recovered at the end of the chemical reaction.

Now the question how catalyst does it?

According to Arrhenius equation, the rate constant k (which measures the rate of reaction) depends on the frequency factor A as well as on the activation energy E_A(activation energy is the minimum energy required to start the chemical reaction) ; larger the A or smaller the E_A, the greater the rate of reaction. In many instances, a catalyst increases the rate of reaction by lowering the activation energy. Let us consider the following equation in absence of catalyst and in presence of catalyst and try to understand the action of a catalyst in terms of activation energy.

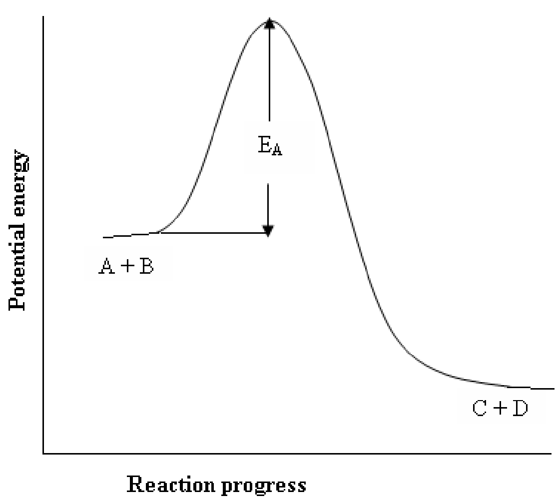
Let assume the following reaction with a certain rate constant k and certain activation energy E_A;

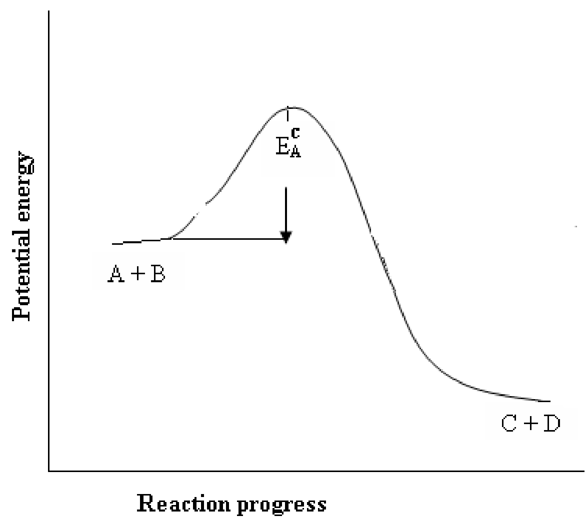
$$A + B \xrightarrow{k} C + D \qquad \text{(absence of a catalyst)}$$

In the presence of a catalyst, k and E_A are not the same. Let us call k as k_c (catalytic rate constant) and E_A as E^c_A (activation energy in presence of catalyst). Thus above reaction is written as

$$A + B \xrightarrow{kc} C + D \qquad \text{(presence of a catalyst)}$$

Potential energy profiles for these two reactions are shown below; the left diagram represents the absence of a catalyst and the right diagram indicates the presence of a catalyst.





Comparing both diagram, it is obvious that

$$E^c_A \quad < \quad E_A$$

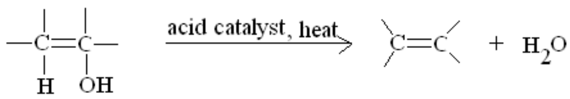
Which means that activation energy of catalyzed reaction is smaller that the activation energy of unanalyzed reaction. Lowering activation energy facilitates increase in rate of reaction because more number of reactant molecules can now pass over the energy barrier towards the product side [It is like a high jump sport in Olympic. If the height of the horizontal bar is lowered, more number of athletes can jump over the bar. Think that the activation energy barrier is like horizontal bar]. Note that the total energy of reactants (A+B) and products (C+D) remains the same.

Type of Catalysis

Generally, the catalysis is classified into two types, depending on the nature of the catalyst; **homogeneous catalysis** and **heterogeneous catalysis.**

Homogeneous Catalysis

In this type, the reactants and catalyst exist in a single phase. There are many organic reactions that require catalysts, for example, synthesis of alkenes from dehydration of alcohols. This reaction requires an acid catalyst and proceeds as follows.



Both reactant and catalyst exist in the same phase.

Another example would be destruction of ozone (O_3) in the upper atmosphere according to the following equation:

$$O_3 (g) + O (g) \rightarrow 2\ O_2(g)$$

where one molecule of ozone collides with an oxygen atom to create two oxygen molecules. This reaction proceeds fairly slowly because of high activation energy (remember higher the activation energy slower the reaction). Due to this, the above reaction proceeds very slowly causing no alarming concern. The ozone layer protects us from the harmful ultraviolet light from the sun. However, the presence of CL atoms that are produced from man-made chlorofluorocarbons(CFCs) (present in aerosols, cleaning agents, refrigerants, and foam plastics) catalyzes the destruction of ozone layer according to the following steps:

Step 1: The Cl atom combines with ozone molecules to form one molecule of CLO and oxygen:

$$CL\ (g) + O_3 (g) \rightarrow CLO(g) + O_2(g)$$

Step 2: The CLO then combines with oxygen atom to produce one CL atom and oen oxygen molecules:

$$CLO\ (g) + O (g) \rightarrow CL\ (g) + O_2(g)$$

The overall reaction is:

$$\cancel{CL\ (g)} + O_3 (g) \rightarrow \cancel{CLO\ (g)} + O_2 (g)$$

$$\cancel{CLO\ (g)} + O\ (g) \rightarrow \cancel{CL\ (g)} + O_2 (g)$$

$$O_3 (g) + O (g) \rightarrow 2\ O_2 (g)$$

Heterogeneous Catalysis

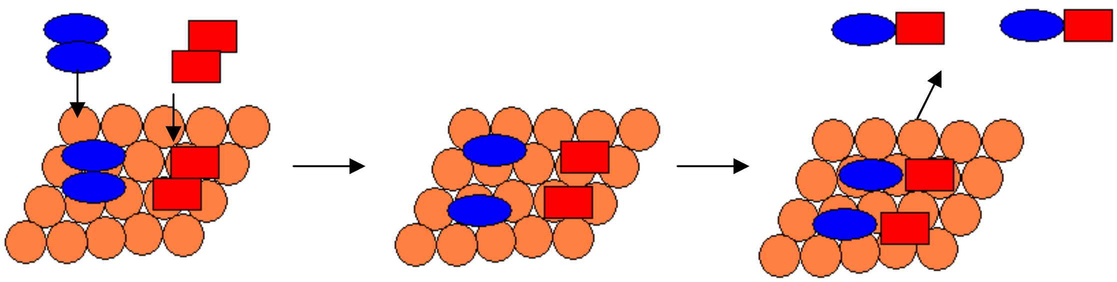
In this type, the reactants and catalyst exist in different phases. Most of the time, the catalyst is solid and the reactants could be either liquids or gases, for example, *Haber process*, which was invented by German chemist Fritz Haber who discovered that reaction of nitrogen gas with hydrogen gas to form ammonia as shown in the following equation

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

is much faster at 500^0 C in presence of a catalyst (mixture of iron, potassium oxide, and aluminum oxide) . Otherwise, it is a slow reaction. Here catalyst is in solid phase and reactants are in gaseous phase.

Importance of Ammonia: Very valuable inorganic substance that is used as the starting material in making fertilizers, explosives, and many other industrial applications.

Action of a Catalyst: In heterogeneous catalysis, the surface of the solid catalyst is usually the site of the reaction means that the catalyst provides a suitable ground for reactants to meet. Once they meet, they react with one another and form products. Therefore, a solid catalyst is a kind of *match-maker*. The entire mechanism shown in the following diagram:



Can you guess what is happening here?

Your Car and Catalytic Converters

When you run your car, you burn gasoline that contains mainly organic substances know as hydrocarbons (substance containing hydrogen and carbon). When these are burned, you produce various gases like carbon dioxide ($CO_2(g)$), carbon monoxide ($CO(g)$), water vapor ($H_2O(g)$), and unburned hydrocarbons (C_xH_y), which eventually discharge in the air through exhaust tail pipe.

hydrocarbon + heat $\rightarrow$ $CO_2(g)$ + $CO (g) + H_2O(g)$ + unburned hydrocarbons

Out of these gases, CO (g) and unburned hydrocarbon (C_xH_y) are harmful to health and create air pollution.

The temperature inside the car engine is high enough for nitrogen (N_2 (g)) and oxygen (O_2 (g)) gases react with one another to produce nitric oxide (NO (g)):

$N_2 (g) + O_2 (g) \rightarrow 2NO (g)$

Nitric oxide, when discharged into the atmosphere, combines with oxygen gas in the air and forms nitrogen dioxide (NO_2 (g)), which also air pollutant.

$$2NO (g) + O_2 (g) \rightarrow 2NO_2 (g)$$

All cars (in the United States anyway) are equipped with two catalytic converters for the purpose of converting harmful gases, mentioned above, into harmless gases. The first converter contains a catalyst like platinum (Pt) or palladium (Pd) or a transition metal oxide such as copper (II) oxide (CuO) or chromium (III) oxide (Cr_2O_3), which oxidizes CO and unburned hydrocarbons to CO_2 and H_2O. The second converter operates at lower temperature than the first one and contains a different catalyst like a transition metal or metal oxide that dissociates NO into N_2 and O_2 before the exhaust gas is thrown out in the air. This process is shown schematically in the following diagram.