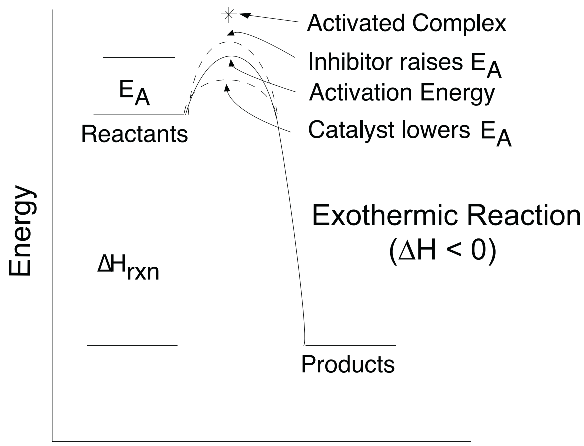
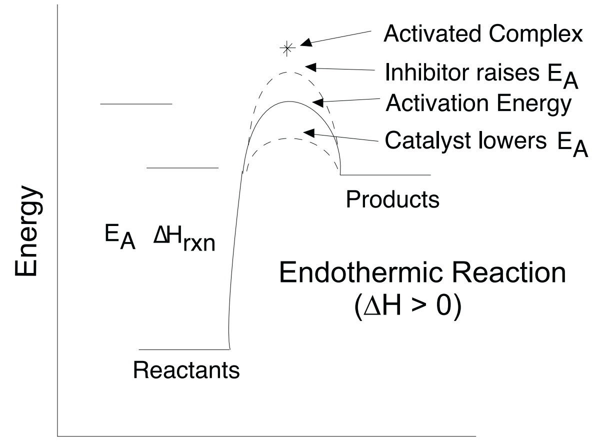
Kinetics: Collision Theory: Student Review Notes

When molecules collide, sometimes they react and sometimes they don’t. It’s explained in terms of an activation energy that is required to trigger the reaction and anintermediate molecule, called the activated complex





* With the idea of activation energy, temperature really comes into play. The higher the temperature, the more energy that is in the system and the closer the molecules are to getting over the activation energy hump.
* The important equation here is the Arrhenius equation that relates the rate constant to temperature.

Arrhenius Equation

$$k = A e^{-\frac{E_A}{RT}}$$

Linearized Arrhenius Equation

$$\ln k = \ln A - \frac{E_A}{RT}$$

k = rate constant

A = frequency factor

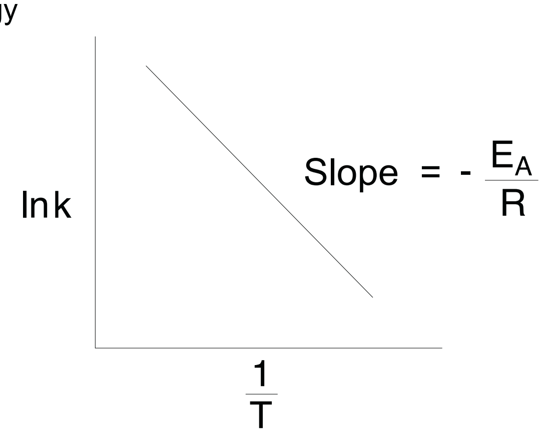
E_A = Activation Energy (Joules)

R = Universal gas constant (8.31 J/molK)

T = Temperature (Kelvin)

So, $\ln k = \ln A - \frac{E_A}{RT}$ is a linear equation that can be used to find a number of things.

A. Plot the natural log of the rate constant, ln (k) vs. inverse temperature, $\frac{1}{T}$ and the slope of the line will give you the activation energy



B. Or, if you know E_A and k at one temperature (Activation energy doesn't change with temperature) then you can find k at a different temperature using the following equation:

$$\ln \frac{k_2}{k_1} = \frac{E_A}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$$

Also, if you know the rate constant and temperature at two points you can use the above equation to solve for the activation energy.