Arrhenius Equation

The temperature dependence of the rate of a chemical reaction is explained by Arrhenius equation.

$$k = Ae^{-E_a/RT}$$

$k \rightarrow$ rate constant

$A \rightarrow$ Arrhenius factor / frequency factor /pre-exponential factor.

It takes into account number of molecules that overcome orientation barrier.

$E_a \rightarrow$ Activation energy

$R \rightarrow$ Gas constant

$T \rightarrow$ temperature

$e^{-E_a/RT} \rightarrow$ It corresponds to the fraction of molecules that have kinetic energy greater than E_a ie the fraction of molecules which cross energy barrier.
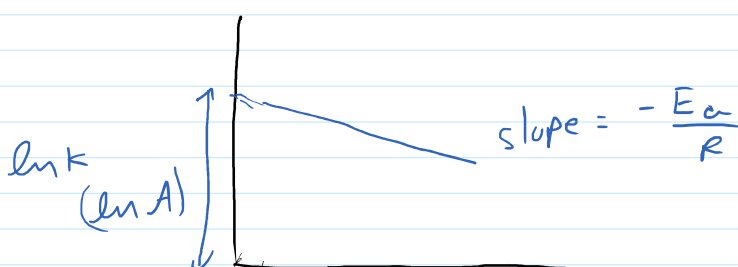
Take natural log on both sides of Arrhenius equation

$$\ln k = \ln Ae^{-E_a/RT}$$
$$= \ln A + \ln e^{-E_a/RT}$$

$$\boxed{\ln k = \ln A - \frac{E_a}{RT}}$$ $(\ln e^x = x \ln e = x)$

$$2.303 \log k = 2.303 \log A - \frac{E_a}{RT}$$

$$\log k = \log A - \frac{E_a}{2.303 \, RT}$$

$\ln k$
$(\ln A)$ slope $= -\frac{E_a}{R}$

$$\frac{1}{T}$$

Let k_1 be rate constant at temperature T_1 and k_2 be rate constant at temperature T_2.

$$\ln k_1 = \ln A - \frac{E_a}{RT_1} \qquad\qquad \ln k_2 = \ln A - \frac{E_a}{RT_2}$$

$$\ln k_1 - \ln k_2 = \left(\ln A - \frac{E_a}{RT_1} \right) - \left(\ln A - \frac{E_a}{RT_2} \right)$$

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$2.303 \log \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$\boxed{\log \frac{k_1}{k_2} = \frac{E_a}{2.303 R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)}$$