Temperature Dependence on the Rate Constant

1. Experiment

Determine the activation energy for the reaction

$$S_2O_8^{2-} + 2\,I^- \rightarrow 2\,SO_4^{2-} + I_2$$

The rate law for the reaction is

$$\text{Rate} = k[S_2O_8^{2-}][I^-]$$

Experimental Setup:

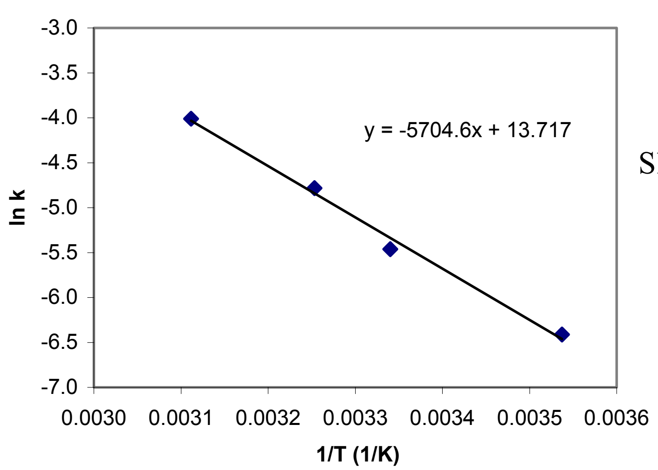
	Temp (°C)	Volumes of Reagents (mL) 0.20 M NaI	0.010 M $Na_2S_2O_3$	2% Starch Indicator	Water	0.20 M $K_2S_2O_8$	Reaction Time (s)
1	**26.2**	**4.0**	**4.0**	**2.0**	**4.0**	**4.0**	**134**
2	**34.2**	**4.0**	**4.0**	**2.0**	**4.0**	**4.0**	**68**
3	**48.2**	**4.0**	**4.0**	**2.0**	**4.0**	**4.0**	**31**
4	**9.5**	**4.0**	**4.0**	**2.0**	**4.0**	**4.0**	**349**
5							

For reasons we won't go into here,

$$\text{Rate} = \frac{\Delta[S_2O_8^{2-}]}{t_{\text{rxn}}} = \frac{1}{2}\frac{[S_2O_3^{2-}]_0}{t_{\text{rxn}}}$$

	Rate (M/s)	Rate Constant
1	**8.21×10^{-6}**	**4.24×10^{-3}**
2	**1.62×10^{-5}**	**8.36×10^{-3}**
3	**3.55×10^{-5}**	**1.83×10^{-2}**
4	**3.23×10^{-6}**	**1.64×10^{-3}**
5		

Prepare an Arrhenius plot:



$$\text{Slope} = -\frac{E_a}{R} \qquad E_a = 5704.6\ \text{K} \times 8.314\,\tfrac{\text{J}}{\text{mol}\cdot\text{K}} = 47{,}400\,\tfrac{\text{J}}{\text{mol}}$$

2. The reaction to produce ethyl alcohol from ethyl iodide

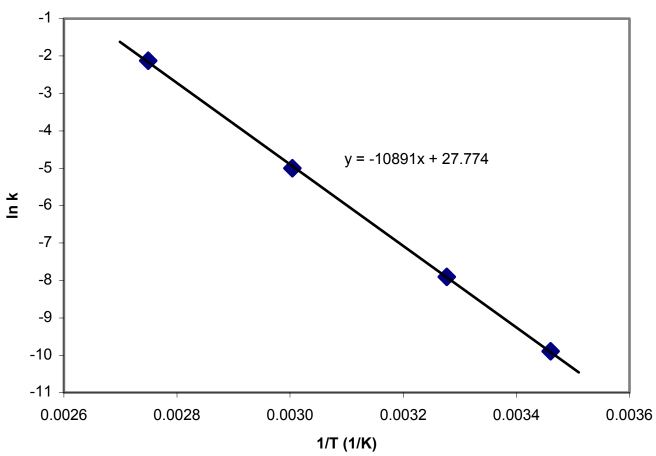
$$C_2H_5I + OH^- \rightarrow C_2H_5OH + I^-$$

was studied at several temperatures. The following rate constants for the reaction were determined:

Temperature (°C)	Rate Constant ($M^{-1}s^{-1}$)
15.83	5.03×10^{-5}
32.02	3.68×10^{-4}
59.75	6.71×10^{-3}
90.61	0.119

Determine the activation energy of the reaction graphically and by using the Arrhenius equation.

Construct an Arrhenius plot:



$$\text{slope} = -\frac{E_a}{R} \quad E_a = 10{,}891\ \text{K} \times 8.314 \tfrac{\text{J}}{\text{mol·K}} = 90{,}550 \tfrac{\text{J}}{\text{mol}}$$

or use the Arrhenius equation:

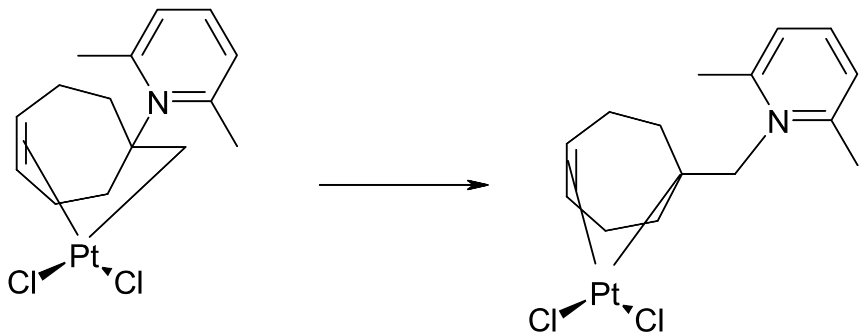
$$\ln\frac{k_1}{k_2} = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$\ln\frac{5.03\times10^{-5}\,\text{M}^{-1}\text{s}^{-1}}{3.68\times10^{-4}\,\text{M}^{-1}\text{s}^{-1}} = \frac{E_a}{8.314\frac{\text{J}}{\text{mol·K}}}\left(\frac{1}{305.17\text{K}} - \frac{1}{288.98\text{K}}\right)$$

$$-1.9901 = \frac{E_a}{8.314\frac{\text{J}}{\text{mol·K}}}\left(-1.8359\times10^{-4}\,\text{K}^{-1}\right)$$

$$E_a = 90{,}124\tfrac{\text{J}}{\text{mol}}$$

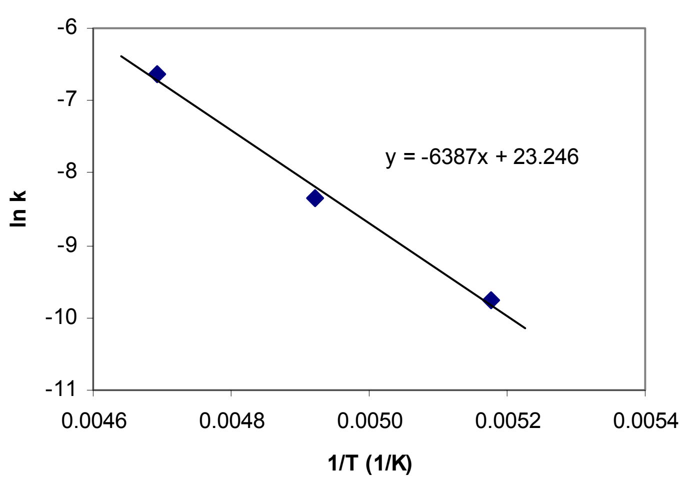
3. The following reaction is possibly important in the catalytic hydrogenation of alkene hydrocarbons. Determine the activation energy for the reaction.



Determine the activation energy of the reaction. What is the rate constant for the reaction at 25°C?

Data

Temp (°C)	Rate Constant (s^{-1})
-60.0	1.3×10^{-3}
-70.0	2.4×10^{-4}
-80.0	5.8×10^{-5}



$$E_a = 6387\text{K} \times 8.314 \tfrac{\text{J}}{\text{mol}\cdot\text{K}} = 53{,}100 \tfrac{\text{J}}{\text{mol}}$$

$$\ln A = 23.246$$

$$\ln k = \ln A - \frac{E_a}{RT} = 23.246 - \frac{6387\text{K}}{298\text{K}} = 1.813$$

$$k = 6.1\ \text{s}^{-1}$$