

Part A. Determination of a Rate Law

Q1. Record the time required for each of your solutions trials in Part A to turn blue:

	KIO ₃ in 50 mL	Buffer pH	KI in 50 mL	Time (s)
Trial 1	0.010 M	pH 4.70	0.10 M	18.6
Trial 2	0.010 M	pH 5.00	0.10 M	72.0
Trial 3	0.020 M	pH 4.70	0.10 M	9.5
Trial 4	0.010 M	pH 4.70	0.20 M	5.2

Q2. (4 pt) Complete the table below, showing sample calculations for your Trial 1 data:

	[IO ₃ ⁻] in 100 mL final mixture	[H ⁺]	[I ⁻] in 100 mL final mixture	Rate
Trial 1	$2.5 \times 10^{-3} \text{ M}$	$2.0 \times 10^{-5} \text{ M}$	$2.5 \times 10^{-2} \text{ M}$	$-1.3 \times 10^{-4} \frac{\text{M}}{\text{s}}$
Trial 2	$2.5 \times 10^{-3} \text{ M}$	$1.0 \times 10^{-5} \text{ M}$	$2.5 \times 10^{-2} \text{ M}$	$-3.5 \times 10^{-5} \frac{\text{M}}{\text{s}}$
Trial 3	$5.0 \times 10^{-3} \text{ M}$	$2.0 \times 10^{-5} \text{ M}$	$2.5 \times 10^{-2} \text{ M}$	$-5.3 \times 10^{-4} \frac{\text{M}}{\text{s}}$
Trial 4	$2.5 \times 10^{-3} \text{ M}$	$2.0 \times 10^{-5} \text{ M}$	$5.0 \times 10^{-2} \text{ M}$	$-4.8 \times 10^{-4} \frac{\text{M}}{\text{s}}$

Sample calculations of [IO₃⁻], [H⁺], [I⁻] and rate (Trial 1):

$$[\text{IO}_3^-]: C_1 V_1 = C_2 V_2 \rightarrow 0.005 \text{ M} \cdot 0.05 \text{ L} = C_2 \cdot 0.10 \text{ L} \rightarrow C_2 = 2.5 \times 10^{-3} \text{ M}$$

$$[\text{H}^+]: -\text{pH} = \log[\text{H}^+] \therefore [\text{H}^+] = 10^{-4.7} = 2.0 \times 10^{-5} \text{ M}$$

$$[\text{I}^-]: C_1 V_1 = C_2 V_2 \rightarrow 0.05 \text{ M} \cdot 0.05 \text{ L} = C_2 \cdot 0.10 \text{ L} \rightarrow C_2 = 0.025 \text{ M}$$

$$\text{Rate} = \frac{[2.5 \times 10^{-3} \text{ M}] - [0.005 \text{ M}]}{18.6 \text{ sec}} = -0.00013 \frac{\text{M}}{\text{s}}$$

- Q3. (1 pt) Determine the reaction order with respect to IO_3^- , showing your work below; round your answer to the nearest half or full integer:

$$\frac{R_1}{R_2} = \frac{k [\text{IO}_3^-]_1^a [\text{H}^+]_1^b [\text{I}^-]_1^c}{k [\text{IO}_3^-]_2^a [\text{H}^+]_2^b [\text{I}^-]_2^c} = \frac{-1.3 \times 10^{-4}}{-3.5 \times 10^{-5}} = \left(\frac{2.0 \times 10^{-5}}{1.0 \times 10^{-5}} \right)^b \rightarrow 3.87 \approx 2.00^b$$

$$\therefore \boxed{b = 2}$$

- Q4. (1 pt) Determine the reaction order with respect to H^+ , showing your work below; round your answer to the nearest half or full integer:

$$\frac{R_1}{R_3} = \frac{k [\text{IO}_3^-]_1^a}{k [\text{IO}_3^-]_3^a} = \frac{-1.3 \times 10^{-4}}{-5.3 \times 10^{-4}} = \left(\frac{.0025}{.005} \right)^a = 0.26 \approx 0.5^a \therefore \boxed{a = 2}$$

- Q5. (1 pt) Determine the reaction order with respect to I^- , showing your work below; round your answer to the nearest half or full integer:

$$\frac{R_1}{R_4} = \frac{k [\text{I}^-]_1^c}{k [\text{I}^-]_4^c} = \frac{-1.3 \times 10^{-4}}{-4.8 \times 10^{-4}} = \left(\frac{.025}{.05} \right)^c \rightarrow 0.28 \approx 0.5^c$$

$$\therefore \boxed{c = 2}$$

- Q6. (1 pt) Write the full rate law for this reaction:

$$\text{Rate} = k [\text{IO}_3^-]^2 [\text{H}^+]^2 [\text{I}^-]^2$$

- Q7. (2 pt) Calculate the rate constant, k , including units, for each of your four trials, then calculate the average value of k ; show all work:

A = IO_3^-
B = H^+
C = I^-

$$k_1 = \frac{R_1}{[\text{A}]_1^2 [\text{B}]_1^2 [\text{C}]_1^2} = \frac{-1.3 \times 10^{-4}}{(.0025)^2 (2 \times 10^{-5})^2 (.025)^2} = -8.6 \times 10^{13} \frac{1}{\text{M}^5 \cdot \text{s}}$$

$$k_2 = \frac{R_2}{[\text{A}]_2^2 [\text{B}]_2^2 [\text{C}]_2^2} = \frac{-3.5 \times 10^{-5}}{3.9 \times 10^{19}} = -8.9 \times 10^{13} \frac{1}{\text{M}^5 \cdot \text{s}}$$

$$k_3 = \frac{-5.3 \times 10^{-4}}{6.22 \times 10^{-18}} = -8.5 \times 10^{13}$$

$$k_4 = \frac{-4.8 \times 10^{-4}}{6.22 \times 10^{-19}} = -7.7 \times 10^{13}$$

$$k_{\text{Avg}} = \frac{k_1 + k_2 + k_3 + k_4}{4} = -8.4 \times 10^{13} \frac{1}{\text{M}^5 \cdot \text{s}}$$

Part B. Determination of Activation Energy

Q8. Record the time required for each of your solutions trials in Part B to turn blue:

	KIO ₃ in 50 mL	Buffer pH	KI in 50 mL	Time (s)
25°C	0.010 M	pH 4.70	0.10 M	18.6
35°C	0.010 M	pH 4.70	0.10 M	8.0
45°C	0.010 M	pH 4.70	0.10 M	3.9

Q9. (2 pt) Complete the following table, showing sample calculations for your 25°C data:

	[IO ₃ ⁻]	[H ⁺]	[I ⁻]	rate	k
25°C	0.005 M	2.0 × 10 ⁻⁵ M	0.05 M	-2.7 × 10 ⁻⁴ M/s	-2.7 × 10 ¹² 1/M ³ ·s
35°C	0.005 M	2.0 × 10 ⁻⁵ M	0.05 M	-6.3 × 10 ⁻⁴ M/s	-6.3 × 10 ¹² 1/M ³ ·s
45°C	0.005 M	2.0 × 10 ⁻⁵ M	0.05 M	-1.3 × 10 ⁻³ M/s	-1.3 × 10 ¹³

Sample calculations of rate and k (25°C):

$$R = \frac{\Delta[A]}{\Delta t} = \frac{(0.005 - 0.01) \text{ M}}{18.6 \text{ sec}} = -2.7 \times 10^{-4} \frac{\text{M}}{\text{s}}$$

$$k = \frac{R}{[A]^2[B]^2[C]^2} = \frac{-2.7 \times 10^{-4} \frac{\text{M}}{\text{s}}}{1.0 \times 10^{-16}} = -2.7 \times 10^{12}$$

Q10. (2 pt) Use Excel, Google Sheets or a similar program to an Arrhenius plot from your temperature data, where your graph must include a proper title and labeled axes, as well as the equation of the linear trend line and its R² value. Add a blank page to the end of this lab report, then paste your graph onto that page.

Q11. (1 pt) Calculate the activation energy ($\frac{\text{kJ}}{\text{mol}}$) for your reaction:

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

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

$$E_A = 61974.9 \frac{\text{J}}{\text{mol}} \left[62 \frac{\text{kJ}}{\text{mol}} \right]$$

Rate of Reaction of IO_3^-

