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	2.5×10^{-3} M	2.0×10^{-5} M	2.5×10^{-2} M	-1.3×10^{-4} $\frac{M}{s}$
Trial 2	2.5×10^{-3} M	1.0×10^{-5} M	2.5×10^{-2} M	-3.5×10^{-5} $\frac{M}{s}$
Trial 3	5.0×10^{-3} M	2.0×10^{-5} M	2.5×10^{-2} M	-5.3×10^{-4} $\frac{M}{s}$
Trial 4	2.5×10^{-3} M	2.0×10^{-5} M	5.0×10^{-2} M	-4.8×10^{-4} $\frac{M}{s}$

Sample calculations of $[IO_3^-]$, $[H^+]$, $[I^-]$ and rate (Trial 1):

$[IO_3^-]$: $C_1V_1 = C_2V_2 \rightarrow 0.005M \cdot 0.05\,L = C_2 \cdot 0.10\,L \rightarrow C_2 = 2.5\times10^{-3}\,M$

$[H^+]$: $-pH = \log[H^+]$ $\therefore [H^+] = 10^{-4.7} = 2.0\times10^{-5}$

$[I^-]$: $C_1V_1 = C_2V_2 \rightarrow 0.05M \cdot 0.05\,L = C_2 \cdot 0.10\,L \rightarrow C_2 = 0.025\,M$

$$\text{Rate} = \frac{[2.5\times10^{-3}\,M] - [0.005\,M]}{18.6\ \text{sec}} = -0.00013\,\frac{M}{s}$$

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:

$$\frac{R_1}{R_2} = \frac{\cancel{k}\cancel{[IO_3^-]_1^a}[H^+]_1^b\cancel{[I^-]_1^c}}{\cancel{k}\cancel{[IO_3^-]_2^a}[H^+]_2^b\cancel{[I^-]_2^c}} = \frac{-1.3\times10^{-4}}{-3.5\times10^{-5}} = \left(\frac{2.0\times10^{-5}}{1.0\times10^{-5}}\right)^b \rightarrow 3.87 = 2.00^b$$

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

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:

$$\frac{R_1}{R_3} = \frac{k[IO_3^-]_1^a}{k[IO_3^-]_3^a} = \frac{-1.3\times10^{-4}}{-5.3\times10^{-4}} = \left(\frac{.0025}{.005}\right)^a = 0.26 = 0.5^a \quad \therefore \boxed{a = 2}$$

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:

$$\frac{R_1}{R_4} = \frac{k[I^-]_1^c}{k[I^-]_4^c} = \frac{-1.3\times10^{-4}}{-4.8\times10^{-4}} = \left(\frac{.025}{.05}\right)^c \rightarrow 0.28 = 0.5^c$$

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

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

$$\text{Rate} = k[IO_3^-]^2[H^+]^2[I^-]^2$$

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:

$A = IO_3^-$

$B = H^+$

$C = I^-$

$$k_1 = \frac{R_1}{[A]_1^2[B]_1^2[C]^2} = \frac{-1.3\times10^{-4}}{(.0025)^2(2\times10^{-5})^2(.025)^2} = -8.6\times10^{13}\ \frac{1}{M^5\cdot s}$$

$$k_2 = \frac{R_2}{3.9\times10^{7+19}} = \frac{-3.5\times10^{-5}}{2.5\times10^{7}} = -8.9\times10^{13}\ \frac{1}{M^5\cdot s}$$

$$k_3 = \frac{-5.3\times10^{-4}}{6.22\times10^{-18}} = -8.5\times10^{13}$$

$$k_4 = \frac{-4.8\times10^{-4}}{6.22\times10^{-19}} = -7.7\times10^{13}$$

$$k_{Avg} = \frac{k_1+k_2+k_3+k_4}{4} = \boxed{-8.4\times10^{13}\ \frac{1}{s\cdot M^5}}$$

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	$-2.7\times10^{-4}\ \frac{M}{s}$	$-2.7\times10^{12}\ \frac{1}{M^5\cdot s}$
35°C	0.005 M	2.0×10^{-5} M	0.05 M	$-6.3\times10^{-4}\ \frac{M}{s}$	$-6.3\times10^{12}\ \frac{1}{M^5\cdot s}$
45°C	0.005 M	2.0×10^{-5} M	0.05 M	$-1.3\times10^{-3}\ \frac{M}{s}$	-1.3×10^{13}

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

$$R = \frac{\Delta[A]}{\Delta t} = \frac{(0.005-0.01)\,M}{18.6\ sec} = -2.7\times10^{-4}\ \frac{M}{s}$$

$$k = \frac{R}{[A]^2[B]^2[C]^2} = \frac{-2.7\times10^{-4}\ \frac{M}{s}}{1.0\times10^{-16}} = -2.7\times10^{12}$$

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 R^2 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 $\left(\frac{kJ}{mol}\right)$ for your reaction:

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_A}{R}\left(\frac{1}{T_1}-\frac{1}{T_2}\right) \qquad E_A = \frac{R\cdot\ln\left(\frac{k_2}{k_1}\right)}{\left(\frac{1}{T_1}-\frac{1}{T_2}\right)} = \frac{8.314\ \frac{J}{mol\,K}\cdot\ln\left(\frac{-1.3\times10^{13}}{-2.7\times10^{12}}\right)}{\left(\frac{1}{298.15}-\frac{1}{318.15}\right)K}$$

$$E_A = 61974.9\ \frac{J}{mol} \quad \boxed{62\ \frac{kJ}{mol}}$$