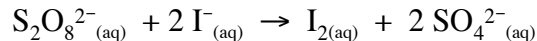


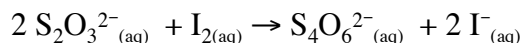
Chemical Kinetics: The Iodine-Clock Reaction:



To measure the rate of this reaction we must measure the rate of concentration change of one of the reactants or products. Here, it is convenient to carry out a *clock reaction* involving the product I_2 . To do this, you will include (to the reacting $\text{S}_2\text{O}_8^{2-}$ and I^-)

- i) a small (but accurately known) amount of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, and
- ii) some starch indicator.

The added $\text{Na}_2\text{S}_2\text{O}_3$ does not interfere with the rate of above reaction, but it does consume the I_2 as soon as it is formed (see below):



This reaction is *much* faster than the previous, so the conversion of I_2 back to I^- is essentially instantaneous.

In a typical experiment, known concentrations of $\text{S}_2\text{O}_8^{2-}$ and I^- are mixed with a small amount of $\text{S}_2\text{O}_3^{2-}$ and starch. As reaction the first proceeds, $\text{S}_2\text{O}_8^{2-}$ and I^- react to produce I_2 , but nothing appears to be happening because the second reaction is consuming the I_2 as quickly as it is formed.

The important moment is when the $\text{S}_2\text{O}_3^{2-}$ runs out, for then the I_2 is no longer consumed (see the second reaction), and the dark blue starch- I_2 complex forms. The stoichiometry of the second reaction indicates that one I_2 molecule must have been generated for every 2 ions of $\text{S}_2\text{O}_3^{2-}$ initially present in the solution.

$[\text{I}_2]$ generated(from first reaction) = $[\text{I}_2]$ consumed(in second reaction) = $\{1/2[\text{S}_2\text{O}_3^{2-}]$ initially present}

Note that the change in concentration of I_2 (before the blue complex forms) is fixed by amount of $\text{S}_2\text{O}_3^{2-}$ included in the initial reaction mixture.

$$\text{rate} = \frac{d[\text{I}_2]}{dt} = \frac{\frac{1}{2}[\text{S}_2\text{O}_3^{2-}]}{t}$$

Where t = time for blue color to appear.

In this experiment you will:

A. Determine Rate Law and Rate Constant

{k is constant at constant temperature}

B. Vary temperature and estimate E_A

{Using the Arrhenius Equation}

A. Determining the rate constant (k) and the rate law for the iodine-clock reaction:

We know $[S_2O_8^{2-}]$ and $[I^-]$, because we know initial volume and concentration of each solution and the final volume for each solution, (Remember $M_1V_1 = M_2V_2$)

e.g for $[S_2O_8^{2-}]$ in trial 1: $[S_2O_8^{2-}] = (0.100 \text{ M})(0.00100\text{L})/(0.0100\text{L}) = 0.0100 \text{ M}$

and we know the general rate law for the iodine-clock reaction:

$$\text{rate} = k[S_2O_8^{2-}]^\alpha [I^-]^\beta$$

So we will:

1. Experimentally determine **rate** of the reaction (stopwatch, color change, $[S_2O_3^{2-}]$).

$$\text{rate} = \frac{d[I_2]}{dt} = \frac{1}{2} \frac{[S_2O_3^{2-}]}{t}$$

2. Empirically determine α and β .

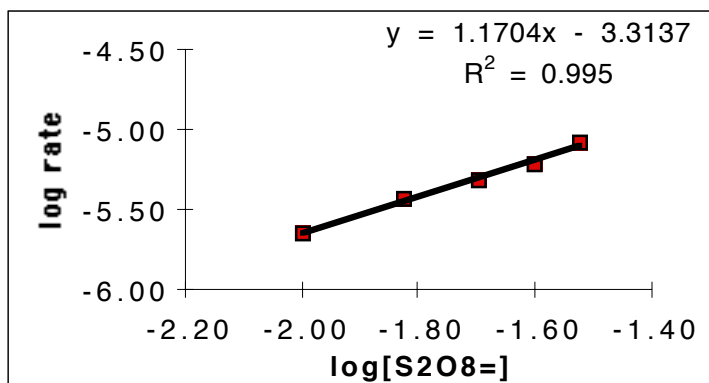
$$\log(\text{rate}) = \log(k) + \alpha \log[S_2O_8^{2-}] + \beta \log[I^-]$$

a. If temperature is constant, k is constant.

If $[I^-]$ is constant, $\log[I^-]$ is constant

$$\log(\text{rate}) = \alpha \log[S_2O_8^{2-}] + C \quad (\text{Where } C = \log(k) + \beta \log[I^-])$$

y = m x + b



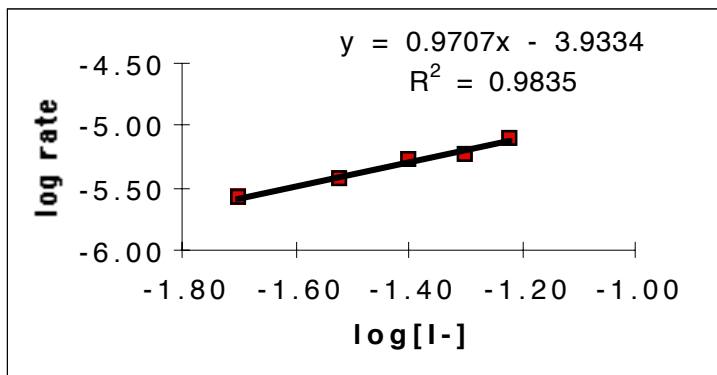
For trials 1 -5, you should plot $\log(\text{rate})$ vs $\alpha \log[S_2O_8^{2-}]$
slope (rounded to the nearest integer) = α

- b. If temperature is constant, k is constant.
 If $[S_2O_8^{2-}]$ is constant, $\log[S_2O_8^{2-}]$

$$\log(\text{rate}) = \beta \log[I^-] + C \quad (\text{Where } C = \log(k) + \alpha \log[S_2O_8^{2-}])$$

$$y = m x + b$$

For trials 6 - 10, you should plot $\log(\text{rate})$ vs $\beta \log[I^-]$:



slope (rounded to the nearest integer) = β

After α and β have been determined:

3. Solve for k @ constant temperature ($^{\circ}K$) for trials 1 - 10 using the rates and α and β that you have determined from trials 1 - 10. (Find mean k , don't forget units units for k : $M^{-1}s^{-1}$)

4. Determine the rate law for the iodine-clock reaction (use mean k from #3: k)

$$\text{rate} = k[S_2O_8^{2-}]^{\alpha} [I^-]^{\beta}$$

B. Varying the temperature, measure rate of reaction, calculate k and estimate E_A

5. Vary the temperature and solve for k (@ each temp).
6. Using k and T, construct an Arrhenius plot and solve for E_A .

Arrhenius Equation: $\ln k = \frac{-E_a}{R} \left(\frac{1}{T}\right) + A$

k = rate constant

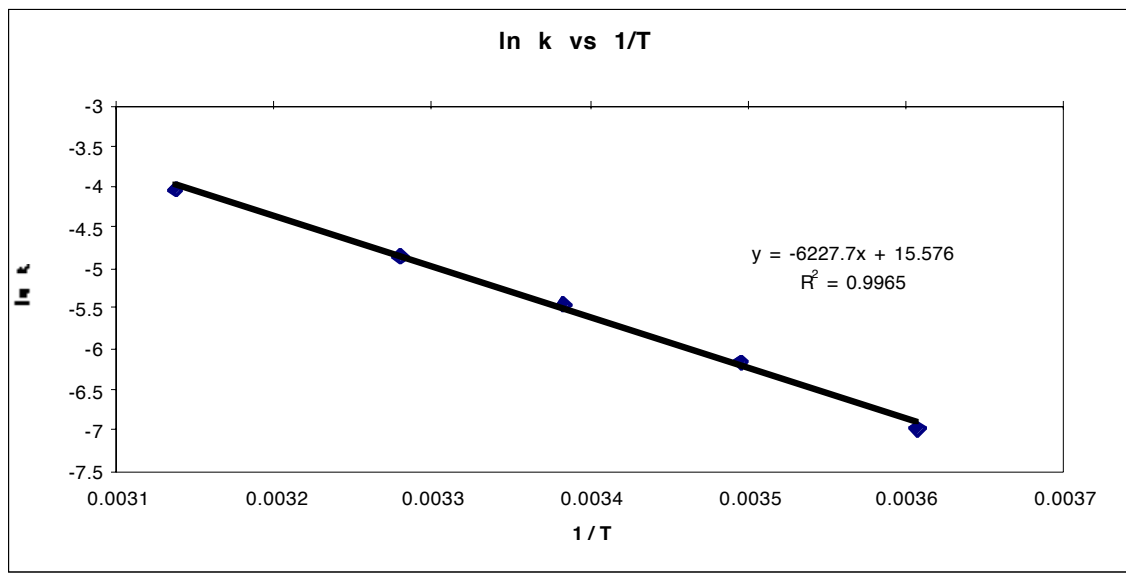
E_A = activation energy

T = temperature ($^{\circ}\text{K}$)

R = Ideal Gas Constant ($8.314 \text{ J}\cdot\text{mol}^{-1}\cdot^{\circ}\text{K}^{-1}$)

A = reaction specific constant

For trials 8, 11 - 14 you should plot $\ln(k)$ vs $1/T$



slope = $-E_A/R$

In this example : $-E_A = (R)(-6227.7 \text{ }^{\circ}\text{K})$

$E_A = (8.314 \text{ J}\cdot\text{mol}^{-1}\cdot^{\circ}\text{K}^{-1})(6227.7 \text{ }^{\circ}\text{K})$

$E_A = 51777 \text{ J/mol}$ or 51.777 kJ/mol