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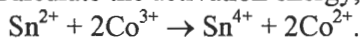
rate = k      rate = k[A]      rate = k[A]<sup>2</sup>      [A]<sub>t</sub> = -kt + [A]<sub>0</sub>      ln[A]<sub>t</sub> = -kt + ln[A]<sub>0</sub>      R = 8.314 J/(mol•K)

1/[A]<sub>t</sub> = kt + 1/[A]<sub>0</sub>      t<sub>1/2</sub> = [A]<sub>0</sub>/2k      t<sub>1/2</sub> = 0.693/k      t<sub>1/2</sub> = 1/k[A]<sub>0</sub>      ln  $\frac{k_1}{k_2} = \frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)$       e=mc<sup>2</sup>

1. (6 Pts) The following mechanism has been suggested for the reaction:  $\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{I}^- \rightarrow \text{I}_2 + 2\text{H}_2\text{O}$
- $\text{H}_2\text{O}_2 + \text{I}^- \rightarrow \text{HOI} + \text{OH}^-$       Slow
- $\text{OH}^- + \text{H}^+ \rightarrow \text{H}_2\text{O}$       Fast
- $\text{HOI} + \text{H}^+ + \text{I}^- \rightarrow \text{I}_2 + \text{H}_2\text{O}$       Fast

- a. What is the rate law for the reaction?      rate = k [H<sub>2</sub>O<sub>2</sub>] [I<sup>-</sup>]
- b. Identify all intermediates included in this mechanism.      HOI and OH<sup>-</sup>

2. (6 Pts) Calculate the activation energy, in kJ/mol, for the redox reaction



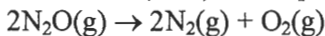
Temp (°C)	k (1/M•s)
2	3.12 × 10 <sup>3</sup>
27	27.0 × 10 <sup>3</sup>

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

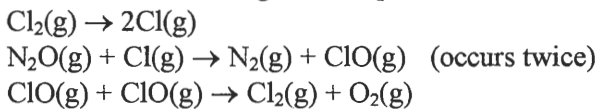
$\ln \frac{3.12 \times 10^3}{27.0 \times 10^3} = \frac{E_a}{8.314} \left( \frac{1}{300} - \frac{1}{275} \right)$

$E_a = 59207 \text{ J/mol} = 59.2 \text{ kJ/mol}$

3. (3 Pts) Nitrous oxide (N<sub>2</sub>O) decomposes at 600°C according to the balanced equation



A reaction mechanism involving three steps is shown below. Identify all any catalysts in the following mechanism.



4. (4 Pts) A certain reaction, reaction A → products, is first order with respect to A. Starting with [A] = 0.250M, it takes 45 min to reduce the concentration of A to 0.110M. What is its rate constant for this reaction?

$\ln [A]_t = -kt + \ln [A]_0$

$\ln [0.110] = -k(45 \text{ min}) + \ln [0.250]$

$k = 1.8 \times 10^{-2} \text{ min}^{-1}$

5. (6 Pts) The activation energy for the following first-order reaction is 102 kJ/mol.  $\text{N}_2\text{O}_5(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g})$
- The value of the rate constant (k) is  $1.35 \times 10^{-4} \text{ s}^{-1}$  at 35°C. What is the value of k at 0°C?

$\ln \frac{k_2}{1.35 \times 10^{-4}} = \frac{102 \times 10^3 \text{ J}}{8.314} \left( \frac{1}{308} - \frac{1}{273} \right)$

$\ln \frac{k_2}{1.35 \times 10^{-4}} = -5.1067 \dots$

$k_2 = 8.17 \times 10^{-7} \text{ s}^{-1}$