CHM 152/54 Quiz #2a 25 Pts Fall 2013 Name: Key

SHOW ALL WORK TO RECEIVE CREDIT

rate = k rate = k[A] rate = k[A]^2 $[A]_t = -kt + [A]_0$ $\ln[A]_t = -kt + \ln[A]_0$ R= 8.314 J/(mol•K)

$1/[A]_t = kt + 1/[A]_0$ $t_{1/2} = [A]_0/2k$ $t_{1/2} = 0.693/k$ $t_{1/2} = 1/k[A]_0$ $\ln\frac{k_1}{k_2} = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$ $e=mc^2$

1. (6 Pts) The following mechanism has been suggested for the reaction: $H_2O_2 + 2H^+ + 2I^- \rightarrow I_2 + 2H_2O$

   $H_2O_2 + I^- \rightarrow HOI + OH^-$ *Slow*
   $OH^- + H^+ \rightarrow H_2O$ *Fast*
   $HOI + H^+ + I^- \rightarrow I_2 + H_2O$ *Fast*

   a. What is the rate law for the reaction? rate = $k[H_2O_2][I^-]$

   b. Identify all intermediates included in this mechanism. HOI and OH^-

2. (6 Pts) Calculate the activation energy, in kJ/mol, for the redox reaction
   $Sn^{2+} + 2Co^{3+} \rightarrow Sn^{4+} + 2Co^{2+}$.

Temp (°C)	k (1/M·s)
2	3.12×10^3
27	27.0×10^3

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

$\ln\frac{3.12\times10^3}{27.0\times10^3} = \frac{E_a}{8.314}\left(\frac{1}{300} - \frac{1}{275}\right)$

$E_a = 59207$ J/mol = 59.2 kJ/mol

3. (3 Pts) Nitrous oxide (N_2O) decomposes at 600°C according to the balanced equation
   $2N_2O(g) \rightarrow 2N_2(g) + O_2(g)$
   A reaction mechanism involving three steps is shown below. Identify all any catalysts in the following mechanism.
   $Cl_2(g) \rightarrow 2Cl(g)$
   $N_2O(g) + Cl(g) \rightarrow N_2(g) + ClO(g)$ (occurs twice)
   $ClO(g) + ClO(g) \rightarrow Cl_2(g) + O_2(g)$

   Cl_2

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. $N_2O_5(g) \rightarrow 2NO_2(g) + {}^1/_2O_2(g)$
   The value of the rate constant (k) is $1.35 \times 10^{-4}\ s^{-1}$ at 35°C. What is the value of k at 0°C?

$\ln\frac{k_2}{1.35\times10^{-4}} = \frac{102\times10^3\ \text{J}}{8.314}\left(\frac{1}{308} - \frac{1}{273}\right)$

$\ln\frac{k_2}{1.35\times10^{-4}} = -5.1067\ldots$

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