CHM151 Quiz 2B 25 Pts Spring 2011 Name: Key

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) $1/[A]_t = kt + 1/[A]_0$

$t_{1/2} = [A]_0/2k$ $t_{1/2} = 0.693/k$ $t_{1/2} = 1/k[A]_0$ $\ln\frac{k_2}{k_1} = \frac{E_a}{R}(\frac{1}{T_1} - \frac{1}{T_2})$ **SHOW ALL WORK**

1 (6 Pts). For the first-order reaction: $\frac{1}{2}N_2O_4(g) \rightarrow NO_2(g)$; $\Delta H = 28.6$ kJ
the rate constant is $k = 4.64 \times 10^5$ s^{-1} at –3°C, and the activation energy is 53.7 kJ/mol. What is the rate constant at 22°C?

$$\ln\frac{k_2}{4.64\times10^5} = \frac{53.7\times10^3\ \text{J/mol}}{8.314\ \frac{\text{J}}{\text{mol}\cdot\text{K}}}\left(\frac{1}{270\text{K}} - \frac{1}{295\text{K}}\right) = 2.0273\ldots$$

$$\frac{k_2}{4.64\times10^5} = e^{2.0273\ldots}$$

$$k_2 = 3.5\times10^6\ \text{s}^{-1}$$

2 (6 Pts) The rate constant for a first-order reaction is 1.6×10^{-2} s^{-1} at 694 K and 4.5×10^{-2} s^{-1} at 892 K. What is the activation energy?

$$\ln\frac{1.6\times10^{-2}}{4.5\times10^{-2}} = \frac{E_a}{8.314}\left(\frac{1}{892} - \frac{1}{694}\right)$$

$$E_a = 26900\ \frac{\text{J}}{\text{mole}} = 26.9\ \frac{\text{kJ}}{\text{mol}}$$

3 (4 Pts). A first-order chemical reaction is observed to have a rate constant of 21 min^{-1}. What is the corresponding half-life for the reaction?

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{21\ \text{min}^{-1}} = \underline{0.033\ \text{min}}$$

4 (9 Pts). For the reaction A + B + C → products, the following initial-rate data were obtained.

	$[A]_0$ (mol/L)	$[B]_0$ (mol/L)	$[C]_0$ (mol/L)	Initial Rate (mol/(L · s))
1.	0.40	0.40	0.20	0.0160
2.	0.20	0.40	0.40	0.0080
3.	0.60	0.10	0.20	0.0015
4.	0.20	0.10	0.20	0.0005
5.	0.20	0.20	0.40	0.0020

rate = $k[A]^x[B]^y[C]^z$

What are the reaction orders with respect to A, B, and C, respectively?

for A: $\frac{\text{Exp 3}}{\text{Exp 4}} \Rightarrow \left(\frac{0.60}{0.20}\right)^x = \frac{0.0015}{0.0005}$ $(3)^x = 3$ $x = 1$

For B: $\frac{\text{Exp 2}}{\text{Exp 5}} \Rightarrow \left(\frac{0.40}{0.20}\right)^y = \frac{0.0080}{0.0020}$ $(2)^y = 4$ $y = 2$

for C use any pair with C changing since we know A + B orders.

i.e. $\frac{\text{Exp 2}}{\text{Exp 1}} \Rightarrow \frac{0.0080}{0.0160} = \cancel{\frac{k}{k}}\left(\frac{0.20}{0.40}\right)^1\cancel{\left(\frac{0.40}{0.40}\right)^2}\left(\frac{0.40}{0.20}\right)^z$

$0.5 = 0.5\,(2)^z$ $2^z = 1$ $z = 0$

yellow

CHM151 Quiz 2a 25 Pts Spring 2011 Name: Key

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) $1/[A]_t = kt + 1/[A]_0$

$t_{1/2} = [A]_0/2k$ $t_{1/2} = 0.693/k$ $t_{1/2} = 1/k[A]_0$ $\ln\frac{k_2}{k_1} = \frac{E_a}{R}(\frac{1}{T_1} - \frac{1}{T_2})$ **SHOW ALL WORK**

1 (4 Pts). A first-order chemical reaction is observed to have a rate constant of 36 min^{-1}. What is the corresponding half-life for the reaction?

$$t_{1/2} = \frac{0.693}{36\ min^{-1}} = 0.0193\ min$$

2 (6 Pts). For the first-order reaction: $\frac{1}{2}N_2O_4(g) \rightarrow NO_2(g)$; $\Delta H = 28.6$ kJ
the rate constant is $k = 2.45 \times 10^5\ s^{-1}$ at –10°C, and the activation energy is 53.7 kJ/mol. What is the rate constant at 14°C?

$$\ln\frac{k_2}{2.45 \times 10^5} = \frac{53.7 \times 10^3\ J/mol}{8.314\ \frac{J}{mol \cdot K}}\left(\frac{1}{263K} - \frac{1}{287K}\right) = 2.0537...$$

$$\frac{k_2}{2.45 \times 10^5} = e^{2.0537...} \quad k_2 = 1.91 \times 10^6\ s^{-1}$$

3 (9 Pts). For the reaction A + B + C → products, the following initial-rate data were obtained.

	$[A]_0$ (mol/L)	$[B]_0$ (mol/L)	$[C]_0$ (mol/L)	Initial Rate (mol/(L · s))
1.	0.40	0.40	0.20	0.0160
2.	0.20	0.40	0.40	0.0080
3.	0.60	0.10	0.20	0.0015
4.	0.20	0.10	0.20	0.0005
5.	0.20	0.20	0.40	0.0020

What are the reaction orders with respect to A, B, and C, respectively?

See 2B Key

4 (6 Pts). The rate constant for a first-order reaction is $1.4 \times 10^{-2}\ s^{-1}$ at 712 K and $4.8 \times 10^{-2}\ s^{-1}$ at 898 K. What is the activation energy?

$$\ln\frac{1.4 \times 10^{-2}}{4.8 \times 10^{-2}} = \frac{E_a}{8.314}\left(\frac{1}{898} - \frac{1}{712}\right)$$

$$E_a = 35200\ \frac{J}{mol} = 35.2\ \frac{kJ}{mol}$$