CHM152 Quiz #1a 25 Pts Spring 13 Name: Key

Show work where possible

1. (6 Pts) What is the rate law that corresponds to the data shown for the reaction $2A + B \rightarrow C$? (Determine the order with respect to each reactant)

Exp.	Initial [A]	Initial [B]	Initial rate
1	0.015	0.022	0.125
2	0.030	0.044	0.500
3	0.060	0.044	0.500
4	0.060	0.066	1.125

*A

*A: Since there is no rate change for 2 & 3, the order of A is 0.

B: picking any pair where B changes, i.e. 2 over 1.

$$\frac{0.500}{0.125} = \left(\frac{0.044}{0.022}\right)^x \qquad \text{rate} = k[B]^2$$

$$4 = 2^x \qquad x = 2$$

2. (11 Pts) Aspirin, $C_9H_8O_4$, slowly decomposes at room temperature by reacting with water in the atmosphere to produce acetic acid, $HC_2H_3O_2$, and 2-hydroxybenzoic acid, $C_7H_6O_3$ (this is why old bottles of aspirin often smell like vinegar):

$$C_9H_8O_4 + H_2O \rightarrow HC_2H_3O_2 + C_7H_6O_3 \qquad \text{rate} = k[C_9H_8O_4]^x[H_2O]^y$$

Concentration and rate data for this reaction are given below.

Exp	$[C_9H_8O_4]$ (M)	$[H_2O]$ (M)	Rate (M/s)
1	0.0100	0.0200	2.4×10^{-13}
2	0.0100	0.0800	9.6×10^{-13}
3	0.0200	0.0200	4.8×10^{-13}

Determine the rate law for this reaction and calculate k (be sure to include the correct units).

for $C_9H_8O_4$ use $\frac{\text{Exp 3}}{\text{Exp 2}}$: $\frac{4.8 \times 10^{-13}}{2.4 \times 10^{-13}} = \left(\frac{0.0200}{0.0100}\right)^x \qquad x = 1$

For H_2O use $\frac{\text{Exp 2}}{\text{Exp 1}}$: $\frac{9.6 \times 10^{-13}}{2.4 \times 10^{-13}} = \left(\frac{0.0800}{0.0200}\right)^y \qquad y = 1$

$$\text{rate} = k[C_9H_8O_4][H_2O]$$

$$k = \frac{2.4 \times 10^{-13}}{[0.0100][0.0200]} = 1.2 \times 10^{-9}\ M^{-1} \cdot s^{-1}$$

3 (4 Pts) Given the rate law for a reaction, rate = $k[A][B]^2$, where rate is measured in units of $M\ s^{-1}$, what are the units for the rate constant k?

$$k = \frac{\text{rate}}{[A][B]^2} = \frac{M}{s} \Big/ M \cdot M^2 = s^{-1} \cdot M^{-2}$$

4. (4 Pts) Which one of the following is not a valid expression for the of the reaction below? **Explain your answer**.

$$4NH_3 + 7O_2 \rightarrow 4NO_2 + 6H_2O$$

a. $-\frac{\Delta[O_2]}{7\Delta t}$ b. $-\frac{\Delta[NO_2]}{4\Delta t}$ c. $\frac{\Delta[H_2O]}{6\Delta t}$ d. $-\frac{\Delta[NH_3]}{4\Delta t}$ e. **rate = $k[NH_3]^x[O_2]^y$**

NO_2 is a product, should be positive

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SHOW ALL WORK TO RECEIVE CREDTI

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)

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

1. (5 Pts) The reaction 2A + B → *products* is second order with respect to A and zero-order with respect to B. Starting with 0.135 M of A, what is the concentration of A after 35 min if the rate constant is 0.11 $M^{-1}s^{-1}$?

$\text{rate} = k[A]^2[B]^0 \Rightarrow \text{rate} = k[A]^2$

2nd order use:

$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$

$\frac{1}{[A]_t} = 0.11(35 \times 60) + \frac{1}{[0.135]}$

$[A]_t = 4.2 \times 10^{-3}\ M$

2. (5 Pts) Nitrogen pentoxide decomposes by a first-order process yielding N_2O_4 and oxygen.

$2N_2O_5 \rightarrow 2N_2O_4 + O_2$

At a given temperature, the half-life of N_2O_5 is 0.90 hr. What is the first-order rate constant for N_2O_5 decomposition?

$t_{1/2} = \frac{\ln 2}{k}$

$k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{0.90} = 0.77\ hr^{-1}$

$2.1 \times 10^{-4}\ s^{-1}$

3. (5 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

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

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

$E_a = 59200$ J/mol

$E_a = 59.2$ kJ/mol

4. (5 Pts) The rate constant for the first-order decomposition of C_4H_8 at 500°C is $9.2 \times 10^{-3}\ s^{-1}$. How long will it take for 10.0% of a 0.100 M sample of C_4H_8 to decompose at 500°C?

$\ln\frac{[A]_t}{[A]_0} = -kt$

$\ln\frac{90}{100} = -9.2 \times 10^{-3}(t)$

t = 11.4 = 11 seconds

5. (5 Pts) Given that E_a for a certain biological reaction is 48 kJ/mol and that the rate constant is $2.5 \times 10^{-2}\ s^{-1}$ at 15°C, what is the rate constant at 37°C?

$\ln\frac{k_1}{2.5 \times 10^{-2}} = \frac{48 \times 10^3}{8.314}\left(\frac{1}{288} - \frac{1}{310}\right)$

$\ln\frac{k}{2.5 \times 10^{-2}} = 1.4226\ldots$

$k = 0.104\ s^{-1}$