CHM152 Quiz 1 Fall 2018 25 Pts 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)

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

1. (4 Pts) The reaction 2A → *products* is <u>second</u> order with respect to A. If the concentration of A drops from 1.05 M to 0.815 M in a time of 15.0 min, what is the rate constant for this reaction (the same time units may be used)?

2nd order ∴ $\frac{1}{[0.815]} = k(15.0\,min) + \frac{1}{[1.05]}$

$k = 0.0183\ M^{-1} \cdot min^{-1}$

2. (4 Pts) a.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?

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

$M^{-2} \cdot s^{-1}$

b. What would be the effect on the rate if the concentration of A and B were both doubled?

$[2][2]^2$ = 8 fold rate increase

3. (4 Pts) Determine the rate law that corresponds to the data shown for the reaction 2A + B → C?

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

General rate law rate $= k[A]^x[B]^y$

experiments 2 & 3 show [A] to be 0 order

show [B] to be 2nd order

experiment 1 & 2

$\frac{rate_2}{rate_1} = \frac{\cancel{k}\,\cancel{[A]^0}}{\cancel{k}\,\cancel{[A]}}\left(\frac{0.030}{0.015}\right)^x = \frac{0.500}{0.12}$

$2^x = 4$

$x = 2$

Rate $= k[B]^2$

MORE ON BACK

Quiz 1 Key

4. (4 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?

90% left $\quad \ln\left[\frac{90}{100}\right] = (-9.2\times10^{-3}\,s^{-1})(t)$

$t = 11$ sec

5. (6 Pts) At a certain temperature, the data below were collected for the reaction below.

$2ICl + H_2 \rightarrow I_2 + 2HCl$.

a. Determine the rate law for the reaction.

Initial concentrations (M)		Inital Rate of Formation of I_2
[ICl]	[H_2]	Mol/L·s
0.10	0.10	0.0015
0.20	0.10	0.0030
0.10	0.050	0.00075

rate $= k[ICl]^x[H_2]^y$

$\left(\frac{2}{1}\right)$ [ICl] : $\left(\frac{0.20}{0.10}\right)^x = \frac{0.0030}{0.0015}$

$2^x = 2 \quad x = 1$

$\left(\frac{1}{3}\right)$ [H_2] $\left(\frac{0.10}{0.050}\right)^y = \frac{0.0015}{0.00075}$

$2^y = 2$

$y = 1$

rate $= k[ICl]^1[H_2]^1$

b. Determine the value and the units of the rate constant.

$k = \frac{\text{rate}}{[ICl][H_2]} : \frac{0.0015\,M/s}{0.10\,M \quad 0.10\,M} = 0.15\ M^{-1}, s^{-1}$

$0.15\ \frac{L}{mol \cdot s}$

6. (3 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?

$t_{1/2} = \frac{\ln 2}{k} \qquad k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{0.90\,hr}$

$0.77\ hr^{-1}$