Quiz 1 (Chapter 12)

Name: Key.

	Zero-Order	First-Order	Second-Order
rate law	rate = k	rate = $k[A]$	rate = $k[A]^2$
units of rate constant	$M\,s^{-1}$	s^{-1}	$M^{-1}\,s^{-1}$
integrated rate law	$[A] = -kt + [A]_0$	$\ln[A] = -kt + \ln[A]_0$	$\frac{1}{[A]} = kt + \left(\frac{1}{[A]_0}\right)$

Arrhenius Equation: $k = Ae^{-E_a/RT}$

1. Consider the following reaction in aqueous solution:

A + 2B → C + D

(4) a) Write the equation that relates the rate expressions for this reaction in terms of the disappearance of A and the disappearance of B.

$$-\frac{\Delta[A]}{\Delta t} = -\frac{1}{2}\frac{\Delta[B]}{\Delta t}$$

(4) b) If the rate of disappearance of A at a particular moment during the reaction is 1.4×10^{-4} mol L^{-1} s^{-1}, what is the rate of B at that moment?

$$\frac{\Delta[B]}{\Delta t} = 2\frac{\Delta[A]}{\Delta t} = 2(1.4 \times 10^{-4}\ mol/Ls)$$

$$= 2.8 \times 10^{-4}\ mol/Ls$$

2. The following data have been determined for the reaction:

$$2NO + Br_2 \rightarrow 2NOBr_2$$

	[NO] initial (M)	$[Br_2]$ initial (M)	Rate ($mol\ L^{-1}\ s^{-1}$)
1	0.02	0.02	9.6×10^{-2}
2	0.04	0.02	3.8×10^{-1}
3	0.02	0.04	1.9×10^{-1}

Determine 1) the rate law and 2) the rate constant for this reaction.

① Rate Law

[NO] trial 1 & 2

when [NO] ×2, rate ×4, thus $rate = [NO]^2$

$[Br_2]$ trial 1 & 3

$$\frac{rate_1}{rate_2} = \frac{k[NO]^2[Br_2]}{k[NO]^2[Br_2]} = \frac{9.6 \times 10^{-2}}{1.9 \times 10^{-1}} = \frac{[0.02]^x}{[0.04]^x}$$

$$= 0.505 = \left(\frac{1}{2}\right)^x$$

(0.5)

$x = 1$

$$rate = k[NO]^2[Br_2]^1$$

② Rate Constant

using trial 1

$$\frac{M}{s} \quad 9.6 \times 10^{-2} = k\,[0.02]^2[0.02]^1 \quad (M^3)$$

$$k = 12000\ \frac{1}{M^2 s}$$

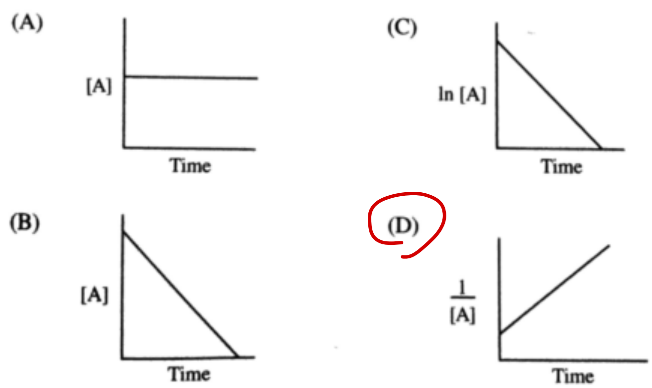
3. Which of the following graphs may have been created using the data gathered from the following reaction? Assume this is a single step reaction:

(3)

$$2A \rightarrow B + C$$

⟹ rate = $k[A]^2$

2nd order



Second-Order
rate = $k[A]^2$
$M^{-1}\,s^{-1}$
$\frac{1}{[A]} = kt + \left(\frac{1}{[A]_0}\right)$

4. Dinitrogen pentoxide gas decomposes according to the equation: $2\ N_2O_{5(g)} \rightarrow 4\ NO_{2(g)} + O_{2(g)}$. The first-order reaction was allowed to proceed at 40 °C. The initial concentration of N2O5 was 0.400 M and after 20 minutes, the concentration changed to 0.289 M.

(4)

(a) Calculate the rate constant for the reaction.

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

$\ln[0.289] = -k(20\ \text{min}) + \ln[0.400]$

$-0.325037.. = -k(20\ \text{min})$

$k = 0.0163\ \text{1/min}$

(3)

(b) After how many minutes will $[N_2O_5]$ be equal to 0.350 M?

$\ln[0.350] = -(0.0163)(x) + \ln[0.400]$

$x = 8.19\ \text{min}$

(5)

5. The rate constant at 550 °C for the decomposition reaction $2H_2O2 \rightarrow 2H_2O + O_2$ is $6.0 \times 10^{-7}\ s^{-1}$, and the frequency factor (A) is $1.2 \times 10^{12}\ s^{-1}$. Determine the activation energy for the reaction.

$$k = A e^{-Ea/RT}$$

$$6.0 \times 10^{-7}\ s^{-1} = (1.2 \times 10^{12}\ s^{-1})\ e^{-\left(\frac{Ea}{8.314 \cdot (550 + 273.15\ k)}\right)}$$

$$\ln 5 \times 10^{-19} = \ln e^{-\frac{Ea}{6843.67}}$$

$$-42.14 = -\frac{Ea}{6843.67}$$

$$Ea = 288390\ J/mol$$

6. At 600 K, compound A decomposes to form compounds B and C via a first-order reaction. Discuss the effect of each of the following conditions on the half-life of A:

(2)

(a) Increasing the initial concentration of A

None Half life is based on rate constant, not concentration.

ex) 100 M → 50 M (20 min) 50 M → 25 M 20 min) Same time.

(2)

(b) Increasing the temperature at which the reaction occurs

Shorter half life

increasing temp increases the reaction rate. time needed to spend 50% will be shorter.

7. Consider the following:

$2NO_{(g)} \leftrightarrow N_2O_{2(g)}$ (fast, k_1 represents the forward rate constant, k_{-1} the reverse rate constant)

$N_2O_{2(g)} + H_{2(g)} \rightarrow N_2O_{(g)} + H_2O_{(g)}$ (slow, k_2 the rate constant)

$N_2O_{(g)} + H_{2(g)} \rightarrow N_{2(g)} + H_2O_{(g)}$ (fast, k_3 the rate constant)

3

(a) Write the overall reaction.

$2NO + N_2O_2 + H_2 + N_2O + H_2$

$\rightarrow N_2O_2 + N_2O + H_2O + N_2 + H_2O$

$2NO + 2H_2 \rightarrow 2H_2O + N_2$

2

(b) Identify all intermediates.

N_2O_2 & N_2O

3

(c) Write the overall rate law.

rate = $k_1[NO]^2 = k_{-1}[N_2O_2]$

$[N_2O_2] = \frac{k_1[NO]^2}{k_{-1}}$

rate = $k_2[N_2O_2][H_2]$

$= k_2\frac{k_1}{k_{-1}}[NO]^2[H_2] = K[NO]^2[H_2]$