KINETICS Practice Problems and Solutions

Name: ______________________________ **AP Chemistry**
Period: _____ **Date:** __________ **Dr. Mandes**

The following questions represent potential types of quiz questions. Please answer each question completely and thoroughly. The solutions will be posted on-line on Monday.

6. Consider the reaction: $P_4 + 6 H_2 \rightarrow 4 PH_3$. A rate study of this reaction was conducted at 298 K. The data that were obtained are shown in the table.

$[P_4]$, mol/L	$[H_2]$, mol/L	Initial Rate, mol/(L · s)
0.0110	0.0075	3.20×10^{-4}
0.0110	0.0150	6.40×10^{-4}
0.0220	0.0150	6.39×10^{-4}

a. What is the order with respect to: P_4 __0___.

H_2 __1___.

b. Write the rate law for this reaction. rate = $k[H_2]$

c. Determine the value and units of the rate constant, k. plug and chug using the rate law & data from exp't 1 and solving for k, we get k = 0.0427 s^{-1}

7. Consider the reaction: $SO_2 + O_3 \rightarrow SO_3 + O_2$. A rate study of this reaction was conducted at 298 K. The data that were obtained are shown in the table.

$[SO_2]$, mol/L	$[O_3]$, mol/L	Initial Rate, mol/(L · s)
0.25	0.40	0.118
0.25	0.20	0.118
0.75	0.20	1.062

a. What is the order with respect to: SO_2 ___2__.

O_3 ___0__.

b. Write the rate law for this reaction. rate = $k[SO_2]^2[O_3]^0$

c. Determine the value and units of the rate constant, k. plug and chug using the rate law & data from exp't 1 and solving for k, we get k = 2.36 mol·L^{-1}· s^{-1}

8. Consider the following mechanism.

$A_2 + B_2 \rightarrow R + C$ (slow)

$A_2 + R \rightarrow C$ (fast)

a. Write the overall balanced chemical equation. $2 A_2 + B_2 \rightarrow 2 C$

b. Identify any intermediates within the mechanism. R

c. What is the order with respect to each reactant? A_2 1st; B_2 1st

d. Write the rate law for the overall reaction. rate = k $[A_2][B_2]$

9. Consider the following mechanism.

$O_3 \rightarrow O_2 + O$ (fast)

$O_3 + O \rightarrow 2 O_2$ (slow)

a. Write the overall balanced chemical equation. $2 O_3 \rightarrow 3 O_2$

b. Identify any intermediates within the mechanism. O

c. What is the order with respect to each reactant? O_3 2nd (once in rds, then once when sub for intermediate)

d. Write the rate law for the overall reaction. rate = k $[O_3]^2$

10. Consider the reaction: 2B → C + 3D. In one experiment it was found that at 300 K the rate constant is 0.134 L/(mol·s). A second experiment showed that at 450 K, the rate constant was 0.569 L/(mol·s). Determine the activation energy for the reaction.

at 300 K: $k_{300} = Ae^{\frac{-Ea}{RT}}$

at 450 K: $k_{450} = Ae^{\frac{-Ea}{RT}}$

$$\ln\left(\frac{k_{450}}{A}\right) = \frac{-E_a}{RT}$$

$$\ln(k_{450}) - \ln(A) = \frac{-E_a}{RT} \quad \textit{where } \ln(A) = \ln(k_{300}) - \frac{-E_a}{RT}$$

so that

$$\ln(k_{450}) - [\ln(k_{300}) - \frac{-E_a}{RT}] = \frac{-E_a}{RT}$$

$$\ln(\frac{k_{450}}{k_{300}}) = \frac{E_a}{R}(\frac{1}{T_{300}} - \frac{1}{T_{450}})$$

plug and solve for Ea, Ea = 10.8 kJ

MORE PROBLEMS>>>>

Determining rate law from mechanisms *(use the rate-determining step to get the orders).*

1. One method for the destruction of ozone in the upper atmosphere is:

	O_3	+	NO	→	NO_2 +	O_2	(slow)
	NO_2	+	O	→	NO +	O_2	(fast)
overall rxn	O_3	+	O	→	$2O_2$		

a. Which species is an intermediate? ______________________

b. Which species is a catalyst? ______________________

c. Which is the rate-determining step (rds)? ______________________

d. Number of times each reactant is used in the rds? ______________________

e. Write the rate law for the reaction. ______________________

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Determining rate law from Initial Rates. *(Use the ratio of initial rates to get the orders).*

2. Consider the table of initial rates for the reaction: $2ClO_2 + 2OH^{1-} \rightarrow ClO_3^{1-} + ClO_2^{1-} + H_2O$.

Experiment	$[ClO_2]_o$, mol/L	$[OH^{1-}]_o$, mol/L	Initial Rate, mol/(L · s)
1	0.050	0.100	5.75×10^{-2}
2	0.100	0.100	2.30×10^{-1}
3	0.100	0. 050	1.15×10^{-1}

a . Order with respect to ClO_2: __________________

b. Order with respect to OH^{1-}: __________________

c. Rate law for this reaction: ________________________________

d. Value and units for the rate constant: ________________________________

3. Consider the table of initial rate for the reaction between hemoglobin (Hb) and carbon monoxide.

Experiment	$[HB]_o$, μmol/L	$[CO]_o$, μmol/L	Initial Rate, μmol/(L · s)
1	2.21	1.00	0.619
2	4.42	1.00	1.24
3	3.36	2.40	2.26

a . Order with respect to HB: __________________

b. Order with respect to CO: __________________

c. Rate law for this reaction: ________________________________

d. Value and units for the rate constant: ________________________________

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Determining rate law from time and concentration data. *(Use the integrated rate laws and graphing to get orders).*

4. The rate of this rxn depends only on NO_2: $NO_2 + CO \rightarrow NO + CO_2$.

The following data were collected.

Time (s)	**$[NO_2]$ (mol/L)**
0	0.500
1200.	0.444
3000.	0.381
4500.	0.340
9000.	0.250
18000.	0.174

a. Order with respect to NO_2: ______________________________

b. Rate law for this reaction: ______________________________

c. $[NO_2]$ at 2.7×10^4 s after the start of the rxn. ______________________________

5. The following data were obtained for the decomposition of N_2O_5 in CCl_4.

The following data were collected.

Time (s)	**$[N_2O_5]$ (mol/L)**
0	1.46
423	1.09
753	0.89
1116	0.72
1582	0.54
1986	0.43
2343	0.35

a. Order with respect to N_2O_5: ______________________________

b. Rate law for this reaction: ______________________________

c. $[N_2O_5]$ at 3.5×10^3 s after the start of the rxn. ______________________________

SOLUTIONS!!!!!! TO "MORE PROBLEMS">>>>>

1.

a. Which species is an intermediate? NO_2

b. Which species is a catalyst? NO

c. Which is the rate-determining step (rds)? slow step

d. Number of times each reactant is used in the rds? O_3 is used once so order is 1
O is used zero times, so order is 0

e. Write the rate law for the reaction. rate = $k[O_3]^1$

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Note that for a free-response question you must show the work (ratio of rate laws), but not for multiple choice

2.

$$\frac{rate_2}{rate_1} = \frac{k[ClO_2]_2^m[OH^-]_2^n}{k[ClO_2]_1^m[OH^-]_1^n} \qquad \frac{rate_2}{rate_3} = \frac{k[ClO_2]_2^m[OH^-]_2^n}{k[ClO_2]_3^m[OH^-]_3^n}$$

$$\frac{0.230}{0.0575} = \frac{0.100^m}{0.0500^m} \qquad \frac{0.230}{0.115} = \frac{0.100^m}{0.0500^m}$$

$$4 = 2^m \qquad 2 = 2^n$$

2 = m, so order is 2   1 = n, so order is 1

a. Order with respect to ClO_2: 2

b. Order with respect to OH^{1-}: 1

c. Rate law for this reaction: so, rate = $k[ClO_2]^2[OH^{1-}]^1$

d. Value and units for the rate constant: k = 230 $\frac{L^2}{mol^2 \cdot s}$

get the value by subbing the data for exp't 1 into the rate law and solving for k

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3.

$$\frac{rate_2}{rate_1} = \frac{k[HB]_2^m[CO]_2^n}{k[HB]_1^m[CO]_1^n} \qquad \frac{rate_3}{rate_1} = \frac{k[HB]_3^m[CO]_3^n}{k[HB]_1^m[CO]_1^n}$$

$$\frac{1.24}{0.619} = \frac{4.42^m}{2.21^m} \text{ s} \qquad \frac{2.26}{0.619} = \frac{(3.36)^1}{(2.21)^1}\frac{(2.41)^n}{(1.00)^n}$$

$$2 = 2^m \qquad 2.4 = 2.4^n$$

1 = m, so the order is 1   1 = n, so the order is 1

a. Order with respect to HB: 1

b. Order with respect to CO: 1

c. Rate law for this reaction: so, rate = $k[HB]^1[CO]^1$

d. Value and units for the rate constant: k = 0.28 $\frac{L}{mol \cdot s}$

get the value by subbing the data for exp't 1 into the rate law and solving for k

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4.

- Graph for zeroeth order: $[NO_2]$ vs. time [y vs. x; y = ax +b]

  slope = -1.72×10^{-5} y-intercept = 0.451 r^2 = 0.901

  General integrated rate law: [A] = -kt + [A]o

  This reaction's integrated rate law: $[H_2O_2] = (-1.72 \times 10^{-5})t + 0.451$ r^2 = 0.901

- Graph for first order: $\ln[NO_2]$ vs. time [y vs. x; y = ax +b]

  slope = -5.78×10^{-5} y-intercept = -0.770 r^2 = 0.971

  General integrated rate law: $\ln[A] = -kt + \ln[A]_o$

  This reaction's integrated rate law: $\ln[NO_2] = (-5.78 \times 10^{-5})t + (-0.770)$ r^2 = 0.971

- Graph for second order: $[NO_2]^{-1}$ vs. time [y vs. x; y = ax +b]

  slope = 2.10×10^{-4} y-intercept = 2.01 r^2 = 0.999 - best so order is 2

  General integrated rate law: $[A]^{-1} = kt + [A]_o^{-1}$

  This reaction's integrated rate law: $[NO_2]^{-1} = 2.10 \times 10^{-4}t + 2.01$ r^2 = 0.999

- Graph with the greatest r^2 value: $[NO_2]^{-1}$ vs. time, *so the order is second order*

a. Order with respect to NO_2: ________2________

b. Rate law for this reaction: ________rate = $k[NO_2]^2$________

c. $[NO_2]$ at 2.7×10^4 s after the start of the rxn. Subbing 2.7×10^4 s for time in "$[NO_2]^{-1} = 2.10 \times 10^{-4}t + 2.01$"

$[NO_2]$ = 0.130 mol/L

5.

- Graph for zeroeth order: $[N_2O_5]$ vs. time [y vs. x; y = ax +b]

  slope = -4.54×10^{-4} y-intercept = 1.31 r^2 = 0.947

  General integrated rate law: [A] = -kt + [A]o

  This reaction's integrated rate law: $[N_2O_5] = (-4.54 \times 10^{-4})t + 1.31$ r^2 = 0.947

- Graph for first order: $\ln[N_2O_5]$ vs. time [y vs. x; y = ax +b]

  slope = -6.05×10^{-4} y-intercept = 0..353 r^2 = 0.999

  General integrated rate law: $\ln[A] = -kt + \ln[A]_o$

  This reaction's integrated rate law: $\ln[N_2O_5] = (-6.05 \times 10^{-4})t + 0.353$ r^2 = 0.999 - best so order is 1

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- Graph for second order: $[N_2O_5]^{-1}$ vs. time [y vs. x; y = ax +b]

  slope = 9.18×10^{-4} y-intercept = 0.517 r^2 = 0.971s

  General integrated rate law: $[A]^{-1}$ = kt - + $[A]_o^{-1}$

  This reaction's integrated rate law: $[N_2O_5]^{-1} = 9.18 \times 10^{-4}t + 0.517$ $r^2 = 0.971$

- Graph with the greatest r^2 value: ln $[N_2O_5]$ vs. time, *so the order is first order*

  Order with respect to N_2O_5:

  Rate law for this reaction:

a. Order with respect to N_2O_5: _______________1____________________

b. Rate law for this reaction: _______________rate = $k[N_2O_5]^1$________

c. $[N_2O_5]$ at 3.5×10^3 s after the start of the rxn. Subbing 3.5×10^3 s for time in "ln$[N_2O_5] = (-6.05 \times 10^{-4})t + 1.31$"

$[N_2O_5]$ = 0.171 mol/L