Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Please write all answers to the short items in the space to the left of the question.

( 1 through 18 are worth two points each)

\_\_\_\_\_1. A certain reaction has a rate constant of 3.5 x 10--4 atm–3/s . It is most likely that the

overall reaction order of this reaction is A) second B) third C) fourth D) fifth

If the unit is atm-3, then the reaction must be FOURTH order

\_\_\_\_\_2. For the second order reaction 2 A → A2 , the instantaneous rate of appearance of

A2 is 0.040 M/sec when [A] is 0.20 molar. What is the rate of appearance of A2 at the

same temperature, when the concentration of A is 0.60 molar?

A) 0.12 M/sec B) 0.24 M/s C) 0.36 M/s D) 0.48 M/s

It is second order, so when the concentration of A2 triples, the rate must be NINE time faster. 0.040 x 9 = choice C. I have sometimes asked “What is the rate of disappearance of A. “ That would be 0.72 M/s

Information for questions 4 to 6.

 The gas phase reaction 2 A + 3 B → D + 2 E is studied. It is found that when 2.00 moles of A and 2.00 moles of B are placed in a 1.00 liter container, the initial rate of formation of D is 0.0020 M/sec. The rate law is determined to be rate = k[A]2[B]2

Substituting into the rate law, The rate, 0.0020M/s = 16 k. k = choice A

\_\_\_\_4. What is the value of the rate constant, k ? A) 1.25 x 10–4 B) 0.00100

C) 0.0080 D) 0.0320

\_\_\_\_5. In the reaction above, what is the rate of disappearance of B, when [A] and [B] are both

4.00 molar? A) 0.032 M/s B) 0.096 M/s C) 0.010 M/s D) 0.024 M/s

Rate = 1.25 x 10-4 ( from question 4) [ 4.00]2 [ 4.00]2 16 x 16 x 1.25 x 10-4 = 0.032 M/s But the correct answer is choice B!!! This is a tricky one; the rate of disappearance of B, because of its coefficient of 3, is three times the rate of reaction.

\_\_\_\_6. What effect would it have on the instantaneous rate of this reaction if the volume of the

container was suddenly halved ? A) Rate would be 4 times greater

B) Rate would be 8 times greater C) Rate would be 16 times greater.

D) Rate would not change. All the concentrations would double. The reaction would be 16 times faster ! ( C)

\_\_\_\_7. What is the rate constant of a first order system that has a half life of 600. seconds?

A) 416 /s B) 866/s C) 0.00116 /s D) 0.00167 HL = .693/k k = .693/600. Choice C

\_\_\_\_\_\_16. A reaction that occurs at a constant rate is most likely to be

A) 0 order B) first order C) second order D) third order A

\_\_\_\_\_\_17. A reaction that has a constant half life is

A) 0 order B) first order C) second order D) third order. B

\_\_\_\_\_\_18. A reaction in which a plot of concentration of reactant against time produces a straight line is A) 0 order B) first order C) second order D) third order.

 A

Fill in questions.

19. The reaction 2 H2O2(aq) → 2 H2O(*l*) + O2(g) is catalyzed by the addition of an

aqueous solution of HBr. Describe the effect of the catalyst on...

(2 points each part)

A. the reaction rate. \_\_\_\_\_faster\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

B. the activation energy of the reaction \_\_\_ smaller\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Problem I. A certain reaction occurs through the following two step mechanism.

1. NO(g) + NO(g) → N2O2(g)

2. N2O2(g) + H2(g) → N2O(g) + H2O(g)

A. Write the overall balanced equation for the net reaction. (2)

 2 NO + H2 → N2O+ H2O

B. Write a rate law for each of the two elementary steps shown above. (4)

1. Rate = k [NO]2

2. Rate = k [ N2O2][H2]

C. Identify any intermediate in the mechanism. (2) N2O2

D. If the observed rate law is Rate = k[NO]2[H2] , which of the two steps is the slow, rate determining step? (2)

 Step 2.

II. In the following mechanism, the letters A through F each represent a gaseous substance.

1. A → B + C (fast )

2. B → D + E (slow)

3. D + C → F (fast)

a. Write the overall reaction. b, Write the rate law for this reaction.

 A → E + F. Rate = k[A]/[C] \*\*\*\* We won’t deal with -1 rate orders on the exam.

III. If the decomposition of H2O2 were found to be first order, with a rate constant of

0.0500 M–1/s, how long would it take for a 2.00 molar solution of H2O2 to decompose

to a concentration of 0.500 molar? (show work) (4 pts)

 Ln( 2.00) – Ln ( .500 ) = 0.0500 t. t = 27.7 seconds. ( You can ALSO solve this using half live)

V. Consider the following reaction between HgCl2(aq) and C2O42–

2 HgCl2(aq) + C2O42– (aq) → 2 Cl– (aq) + 2 CO2(g) + Hg2Cl2(s)

The initial rate was determined for several concentrations of HgCl2(aq) and C2O42– , and

the following rate data were obtained for the **rate of disappearance of C2O42– :**

|  |  |  |  |
| --- | --- | --- | --- |
| Experiment | [HgCl2](M) | [C2O42–](M) | Initial Rate (M/s) |
| 1 | 0.150 | 0.300 | 4.5x10–5 |
| 2 | 0.150 | 0.100 | 5.0x10–6 |
| 3 | 0.450 | 0.100 | 1.5x10–5 |
| 4 | 2.00 | ?? | 4.2 x 10–2 |
| 5 | 0.100 | 0.300 | ?? |

a. Write the rate law for the reaction. ( 3 pts) Rate = k[ C2O42-]2[HgCl2]

b. Find the value of the rate constant k. Include correct units. (3 pts) 3.33 x 10-3 M-2/s

c. Find the initial [C2O42– ] in experiment 4. (3 pts) 2.51 M.

d. Find the initial rate in experiment 5. ( 3pts )

0.100 M ( 0.300 M )2( 3.33 x 10-3 M-2/s) = 3.00 x 10-5 But there is a much faster, easier way to do this!! Do you see it?

(Compare experiment 1 with experiment 5…)