

Name KEY

Last 5 digits of Student Number: XXX – X \_\_\_\_ – \_\_\_\_

**Chem 116**  
**Sample Examination #2**

This exam consists of eight (8) pages, including this cover page. Be sure your copy is complete before beginning your work. If this test packet is defective, ask for another one.

A copy of the Periodic Table is attached at the back of the exam. You may remove it and use the back side of the Periodic Table as scratch paper. No work on scratch paper will be graded or collected.

The following information may be useful:

Equations

Arrhenius equation:  $k = A e^{-E_a/RT}$   
 $t_{half} = \frac{0.693}{k}$  for a first-order reaction

Constants of nature

$$R = 8.314 \frac{J}{mol \cdot K} = 0.08206 \frac{L \cdot atm}{mol \cdot K}$$

More Equations

Differential rate laws:

0. rate of change of  $[A] = k [A]^0$
1. rate of change of  $[A] = k [A]^1$
2. rate of change of  $[A] = k [A]^2$

Integrated rate laws:

0.  $[A] = [A]_0 + k t$
1.  $\ln[A] = \ln[A]_0 - k t$
2.  $\frac{1}{[A]} = \frac{1}{[A]_0} + k t$

DO NOT WRITE BELOW THIS LINE

Part I:

- Questions 1-7 \_\_\_\_\_ (maximum 35)  
 Question 8 \_\_\_\_\_ (maximum 7)  
 Question 9 \_\_\_\_\_ (maximum 8)  
 Question 10 \_\_\_\_\_ (maximum 10)

Part II

**Disclaimer:**

This is a copy of a typical Exam 2 given in Chem 116 during the academic year. Your test will be different. This test is being posted to give you a sense of the format, style, scope and level of a typical test on this material. This test may have questions on topics that may not be covered on your exam. Moreover, your test may have questions on topics not covered in this practice exam. Posting this test in no way limits the format, style, scope and level of the test that you will take. Do not limit your preparation to the material in this practice exam.

Total

## Part I. Multiple-Choice or Short Response

There are 10 questions. Questions 1-7 are multiple-choice and are each worth 5 points. Question 8 is worth 7 points. Question 9 is worth 8 points. Question 10 is worth 10 points.

- A 1. An aqueous solution contains 2.05 g of an unknown, non-ionic solute in 50.0 g of water. The solution boils at 100.062 °C. What is the molar mass of the unknown solute? The boiling point constant for water is 0.5121 °C/m.

- A) 340 g/mol  
B) 1290 g/mol  
C) 401 g/mol  
D) 18 g/mol  
E) 80. g/mol

$$\Delta T_{bp} = i K_{bp} m \quad \text{and the solute is non-ionic so } i = 1$$

$$m = \frac{0.062^\circ\text{C}}{(1)(0.5121^\circ\text{C}/m)} = 0.121 \text{ moles solute / kg solvent}$$

$$\text{moles solute} = 0.121 \frac{\text{moles solute}}{\text{kg solvent}} (0.050 \text{ kg solvent}) = 0.00605 \text{ moles}$$

$$\text{molar mass} = \text{grams/moles} = 2.05 \text{ g}/0.00605 \text{ mol} = 340 \text{ g/mol}$$

- B 2. A particular ionic compound, when dissolved in water, decreases the freezing point of the water. The ionic compound does not dissociate completely, and the following experiment was done to determine how much it dissociates. When 0.0351 moles of the ionic compound were added to 75.0 g of water, the freezing point of the solution was -1.44 °C. What is the van't Hoff factor for the ionic compound? The freezing point constant for water is -1.86 °C/m.

- A) 1.29  
B) 1.65  
C) 0.774  
D) 2.00  
E) 7.05

$$i = \frac{\Delta T_{fp}}{K_{fp} m} = \frac{-1.44^\circ\text{C}}{(-1.86^\circ\text{C}/m) \left( \frac{0.0351}{0.0750} m \right)} = 1.65$$

- D 3. The half-life of the radioactive gold-198 isotope is 2.7 days. If you begin with a  $6.4 \times 10^{-6}$  g sample of the isotope, how much of the sample remains after 3.6 days?

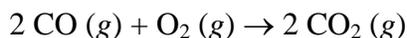
- A)  $6.4 \times 10^{-6}$  g  
B)  $4.8 \times 10^{-6}$  g  
C)  $3.2 \times 10^{-6}$  g  
D)  $2.5 \times 10^{-6}$  g  
E)  $1.6 \times 10^{-6}$  g

$$\text{number of half-lives} = \frac{3.6 \text{ days}}{2.7 \text{ days}} = 1.333$$

$$\text{amount remaining} = \left( \frac{1}{2} \right)^{\text{no. of half-lives}} (\text{original amount})$$

$$= \left( \frac{1}{2} \right)^{1.333} (6.4 \times 10^{-6} \text{ g}) = 2.5 \times 10^{-6} \text{ g}$$

D 4. For the following reaction,



[CO] (mol/L)	[O <sub>2</sub> ] (mol/L)	Initial rate $\left(\frac{\text{mol}}{\text{L} \cdot \text{min}}\right)$
0.20	0.20	$1.84 \times 10^{-4}$
0.20	0.40	$3.68 \times 10^{-4}$
0.60	0.20	$1.66 \times 10^{-3}$

the data above were obtained in laboratory experiments. What is the form of the rate law for this reaction?

A) rate =  $k [\text{CO}][\text{O}_2]$

B) rate =  $k [\text{CO}][\text{O}_2]^2$

C) rate =  $k [\text{CO}]^2$

D) rate =  $k [\text{CO}]^2 [\text{O}_2]$

E) rate =  $k [\text{CO}]^3 [\text{O}_2]$

form must be rate =  $[\text{CO}]^n [\text{O}_2]^m$

comparing experiments 1 and 3: [O<sub>2</sub>] stays constant, [CO] changes  
[CO]×3 causes rate×9  
9 must be 3<sup>n</sup>, so n = 2

comparing experiments 1 and 2: [CO] stays constant, [O<sub>2</sub>] changes  
[O<sub>2</sub>]×2 causes rate×2  
2 must be 2<sup>m</sup>, so m = 1

E 5. Which of the following changes in experimental conditions would not increase the rate of the following reaction:



A) Increasing the temperature

B) Increasing the amount of H<sub>2</sub>O<sub>2</sub> initially present

C) Using a catalyst for the reaction

D) Using less water in the initial reactant mixture

E) Increasing the pressure

Changes that always increase rate of a rxn:

- temperature increase
- use of catalyst
- increase concentration of reactant

Using less water in initial reactant mixture would cause all reactant concentrations to be larger, so rate increase.

Increase in pressure does not do anything for a system that is not gaseous.

A 6. If each of the following systems begins at equilibrium, which disturbance would result in a shift that favors the products as the system returns to equilibrium following the disturbance?

A) Remove some NH<sub>3</sub> (g) from this system:  $3 \text{H}_2 (g) + \text{N}_2 (g) \rightleftharpoons 2 \text{NH}_3 (g)$

B) Increase the volume for this system:  $\text{NO}_2 (g) + \text{CO} (g) \rightleftharpoons \text{NO} (g) + \text{CO}_2 (g)$

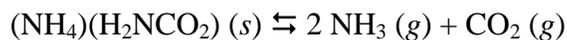
C) Increase the temperature in an exothermic reaction

D) Add more C (s) in this system:  $\text{C}(s) + \text{CO}_2 (g) \rightleftharpoons 2 \text{CO} (g)$

E) Increase the pressure in this system by adding Ar (g):  $2 \text{NO}_2 (g) \rightleftharpoons \text{N}_2\text{O}_4 (g)$

(A) favors products because removing a reactant causes reaction to progress in direction to partially un-do disturbance, hence replace some NH<sub>3</sub>

- C 7. At 43 °C, the reaction below takes place in a closed 2.00 L container.

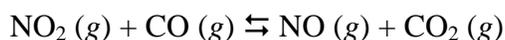


When the system is at equilibrium, there are 0.20 moles of  $\text{NH}_3$  and 0.10 moles of  $\text{CO}_2$  present. What is the value of the equilibrium constant ( $K_c$ ) at this temperature?

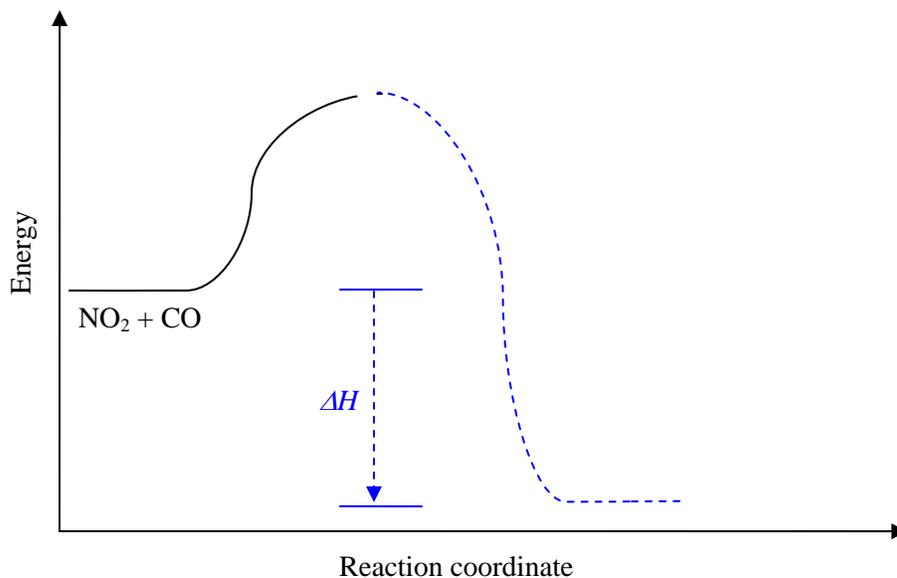
- A)  $K_c = 3.4 \times 10^{-3}$   
 B)  $K_c = 4.0 \times 10^{-3}$   
 C)  $K_c = 5.0 \times 10^{-4}$   
 D)  $K_c = 6.0 \times 10^{-2}$   
 E) there is not enough information to tell

$$K_c = [\text{NH}_3]^2 [\text{CO}_2] = \left( \frac{0.20 \text{ mol}}{2.00 \text{ L}} \right)^2 \left( \frac{0.10 \text{ mol}}{2.00 \text{ L}} \right) = 5.0 \times 10^{-4}$$

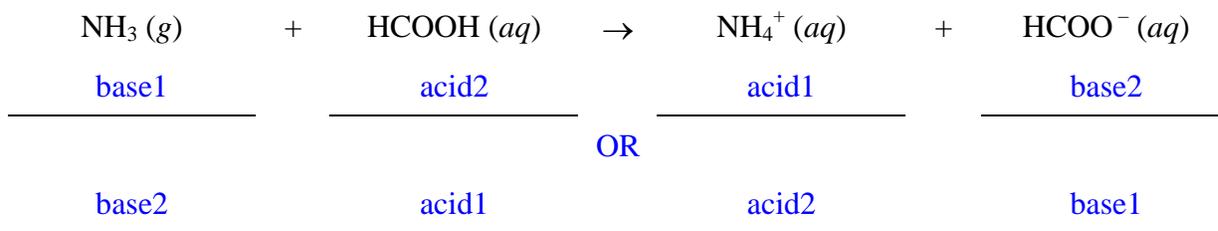
8. The reaction shown below is exothermic, giving off 240 kJ.



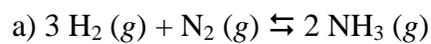
If the activation energy in the forward direction is 130 kJ, appropriately complete the reaction diagram sketch that has been started below. Also, clearly indicate  $\Delta H$  for the reaction on the graph. (Please note: you only need to sketch the rest of the graph – do not worry about making the graph exactly correspond to the values given.)



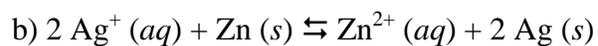
9. For the following reaction, label the conjugate acid-base pairs (acid-1 and base-1; acid-2 and base-2).



10. Write the expressions for the reaction coefficient ( $Q_c$ ) for each of the following reactions in terms of reactant and product concentrations.



$$Q_c = \frac{[\text{NH}_3]^2}{[\text{H}_2]^3 [\text{N}_2]}$$

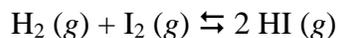


$$Q_c = \frac{[\text{Zn}^{2+}]}{[\text{Ag}^+]^2}$$

## Part II. Problems

Each problem is worth 20 points.

1. At 75 °C, the equilibrium constant  $K_c = 0.522$  for the reaction



Consider a mixture of 0.124 mol of  $\text{H}_2$ , 0.124 mol of  $\text{I}_2$ , and 0.370 mol of  $\text{HI}$  in a 1.00-liter vessel at 75 °C.

- a) What is the value of the reaction quotient for this initial mixture?

$$Q = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{\left(\frac{0.370 \text{ mol}}{1.00 \text{ L}}\right)^2}{\left(\frac{0.124 \text{ mol}}{1.00 \text{ L}}\right)\left(\frac{0.124 \text{ mol}}{1.00 \text{ L}}\right)} = 8.90$$

- b) In what direction will the reaction proceed? In other words, which chemicals will there be more made of?

reaction will proceed toward reactants (to the left)  
not necessary to provide justification, but justification is that  $Q$  is too large, so numerator needs to decrease to approach  $K$

- c) What are the final concentrations of all three chemicals when equilibrium is reached? Show your work.

	$\text{H}_2(g)$	+	$\text{I}_2(g)$	$\rightleftharpoons$	$2 \text{HI}(g)$
Initial	0.124 M		0.124 M		0.370 M
Change	+ $x$		+ $x$		- $2x$
Equilibrium	$0.124 + x$		$0.124 + x$		$0.370 - 2x$

$$K = \frac{[0.370 - 2x]^2}{[0.124 + x][0.124 + x]} = \frac{(0.370 - 2x)^2}{(0.124 + x)^2} = \left(\frac{0.370 - 2x}{0.124 + x}\right)^2 = 0.522$$

Solve for  $x$  by taking square root of both sides and doing algebra  
 $x = 0.103$

Final concentrations:  
 $[\text{H}_2] = [\text{I}_2] = 0.124 + x = 0.227 \text{ M}$   
 $[\text{HI}] = 0.370 - 2x = 0.164 \text{ M}$

Check to make sure values make sense:  $K$  should equal 0.522:

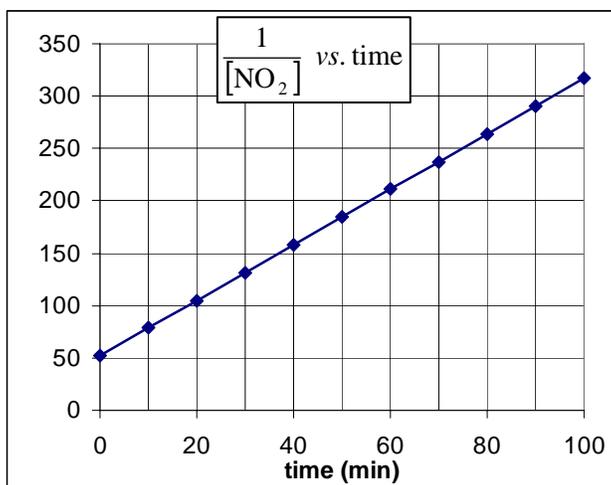
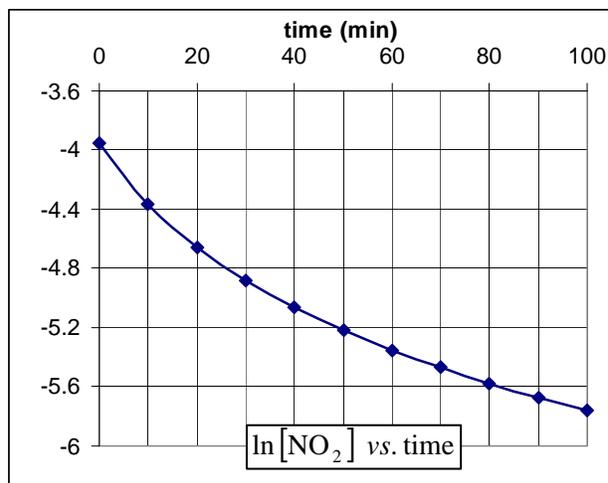
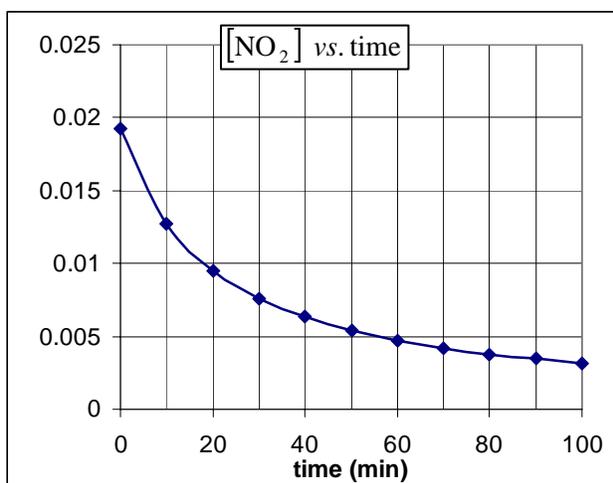
$$\frac{(0.164)^2}{(0.227)(0.227)} = 0.522$$

2. The decomposition of  $\text{NO}_2$  proceeds according to the reaction



In a certain experiment, the reaction took place at 303 K. Data from the experiment are given in the table below and also plotted in the graphs. Use this information to answer the questions on the next page.

Time (minutes)	$[\text{NO}_2]$ (mol/L)	$\ln [\text{NO}_2]$	$[\text{NO}_2]^{-1}$ (L/mol)
0	0.0192	-3.951	52.0
10.	0.0127	-4.363	78.5
20.	0.00952	-4.654	105.
30.	0.00761	-4.879	132
50.	0.00542	-5.218	185
70.	0.00421	-5.470	238
100.	0.00316	-5.759	317



## Problem 2 (continued)

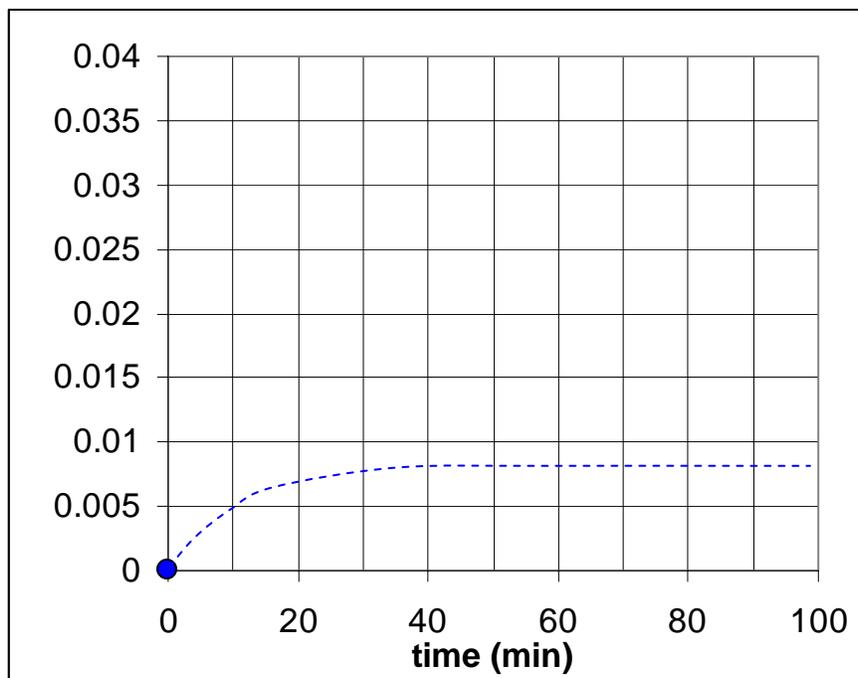
- a) Estimate the initial rate of change of the concentration of  $\text{NO}_2$ . Show your calculations. Be sure to include units in the rate.

$$\frac{\Delta[\text{NO}_2]}{\Delta t} = \frac{0.0127 - 0.0192 \text{ mol/L}}{10. - 0 \text{ min}} = -6.5 \times 10^{-4} \frac{\text{mol}}{\text{L} \cdot \text{min}}$$

- b) What is the order of this reaction with respect to  $\text{NO}_2$ ? Justify your answer.

Second order  
Third graph ( $1/[\text{NO}_2]$  vs. time) is a straight line and corresponds to form for second order rate law

- c) There is no oxygen ( $\text{O}_2$ ) present at the start of the experiment, but it gets produced during the experiment. Using the graph at the right, sketch a plot of the concentration of oxygen,  $[\text{O}_2]$ , versus time, as oxygen is produced during the experiment. The starting point is already shown on the graph.



- d) Write the rate law for this reaction.

$$\text{rate} = k [\text{NO}_2]^2$$

- e) Calculate the rate constant,  $k$ , for this reaction. For extra credit (2 pts maximum), you may specify the units of the rate constant.

$k$  is the slope of the third graph

$$k = \frac{\Delta\left(\frac{1}{[\text{NO}_2]}\right)}{\Delta t} = \frac{78.5 - 52.0 \text{ L/mol}}{10. - 0 \text{ min}} = 2.7 \frac{\text{L}}{\text{mol} \cdot \text{min}}$$