

Please print:

Last name: Answer Key

First name: _____

Chem 1062 Exam 2

Spring 2005

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Thursday, March 3, 2005

Equations:

$$\begin{array}{llll} S = k_H P & P_A = P_A^\circ X_A & \ln \frac{[A]_t}{[A]_0} = -kt_{(1st\ order)} & t_{1/2} = \frac{-\ln \frac{1}{2}}{k} \quad (1st\ order) \\ \text{mass } \% = \frac{\text{mass solute}}{\text{mass solution}} \times 100\% & \Delta P = iP_A^\circ X_B & [A]_t = [A]_0 e^{-kt} \quad (1st\ order) & t_{1/2} = \frac{1}{k[A]_0} \quad (2nd\ order) \\ c_m = \frac{\text{moles solute}}{\text{kg solvent}} & \Delta P = P_A^\circ - P_A & \frac{1}{[A]_t} = kt + \frac{1}{[A]_0} \quad (2nd\ order) & K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \\ M = \frac{\text{moles solute}}{\text{L solution}} & \Delta T_b = iK_b c_m & \ln \frac{k_2}{k_1} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right) & Q_c = \frac{[C]_i^c [D]_i^d}{[A]_i^a [B]_i^b} \\ X_A = \frac{\text{moles A}}{\text{total moles of solution}} & \Delta T_f = iK_f c_m & & \\ & \pi = iMRT & & \end{array}$$

Contents:

I. Multiple choice, 8 questions, 3 points each.	24 points
II. Short answer, 2 questions, worth 16 points.	16 points
III. Problems. 3 questions, worth 20 points each.	<u>60 points</u>
	Total: 100 points

I, _____ have read and understand the directions given above, and pledge that I will follow all regulations with regard to Academic Dishonesty as outlined by this college when taking this exam.

Signature _____ Date and Time _____

I. Multiple choice

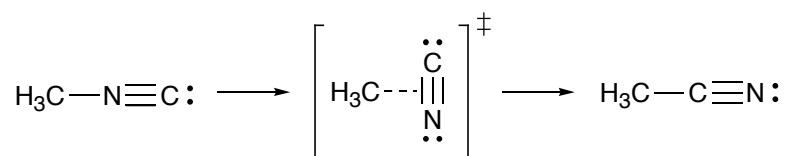
Choose the best answer from the choices given, and clearly **circle** the letter of your choice. (3 pts each)

- Which of the following **increases** the solubility of a gas in a given solvent?
 - increasing the partial pressure of the gas
 - decreasing the temperature of the solvent
 - decreasing the partial pressure of the gas
 - increasing the temperature of the solvent and decreasing the partial pressure of the gas simultaneously
 - both a and b
- All of the following gases have appreciable solubility in water **except**
 - CO₂.
 - SO₃.
 - N₂.
 - NH₃.
 - HCl.
- What is the freezing point of an aqueous 0.750 *m* NH₄I solution? (NH₄I is a strong electrolyte, $K_f = 1.86\text{ }^{\circ}\text{C}/m$)
 - 1.40 $^{\circ}\text{C}$
 - +1.40 $^{\circ}\text{C}$
 - 1.86 $^{\circ}\text{C}$
 - +1.86 $^{\circ}\text{C}$
 - 2.80 $^{\circ}\text{C}$
- Based on the formulas of the following solutions, which compound would have the **smallest** van't Hoff factor, *i*?
 - Ca(NO₃)₂ (aq)
 - MgSO₄ (aq)
 - Th(SO₄)₂ (aq)
 - Al₂(SO₄)₃ (aq)
 - K₂SO₄ (aq)

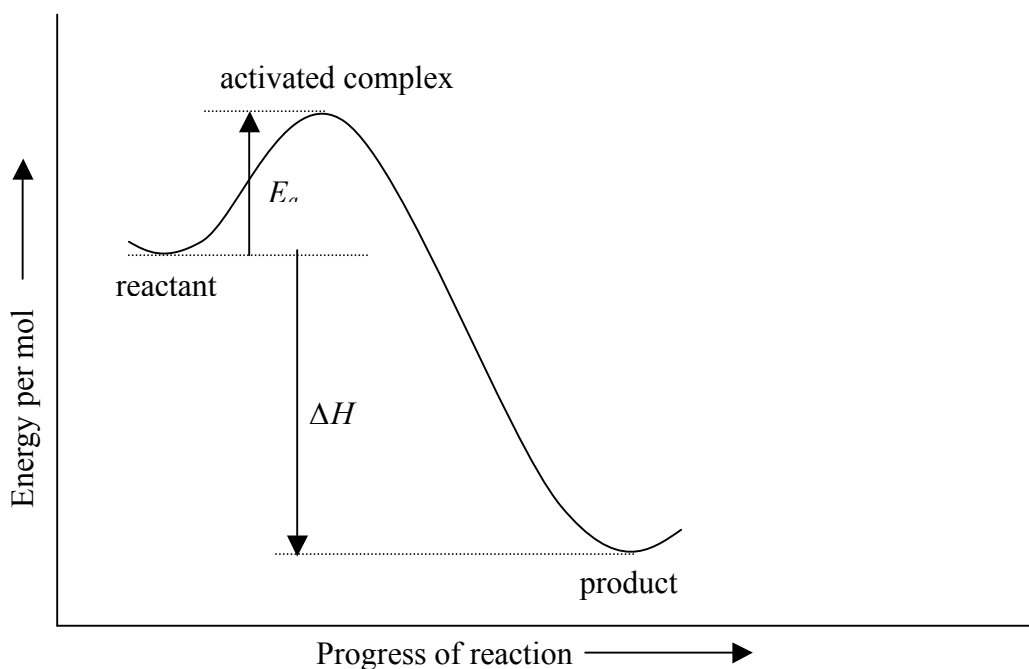
5. Consider the reaction $aA + bB \xrightarrow{C} dD + eE$ where C is a catalyst. The rate law is $\text{Rate} = k[A]^m[B]^n[C]^p$. Which of the following statements is **false**?
- The exponents m , n , and p are often integers.
 - The exponents m and n are always equal to the coefficients a and b , respectively.
 - The exponent p must be determined experimentally.
 - The symbol k represents the rate constant.
 - The overall reaction order is $m + n + p$.
6. All of the following would be expected to affect the rate of a chemical reaction **except**
- adding more reactants.
 - removing some products.
 - increasing the temperature.
 - decreasing the temperature.
 - adding a catalyst.
7. Which of the following statements is always **true**?
- Exothermic reactions have lower activation energies than endothermic reactions.
 - The rate for a reaction depends on the concentrations of all the reactants.
 - The rate of a catalyzed reaction is independent of the concentration of the catalyst.
 - The rate constant is independent of the concentrations of the reacting species.
 - The rate law can be determined from the stoichiometric equation.
8. The following solution is at equilibrium: $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$. If additional $\text{PCl}_5(g)$ is then added to the equilibrium mixture, which of the following is true at the instant the $\text{PCl}_5(g)$ is added?
- $Q_c > K_c$; the reaction will shift to the right.
 - $Q_c > K_c$; the reaction will shift to the left.
 - $Q_c < K_c$; the reaction will shift to the right.
 - $Q_c < K_c$; the reaction will shift to the left.
 - $Q_c = K_c$; the reaction is at equilibrium.

II. Short answer

9. Methyl isonitrile rearranges into acetonitrile by way of an activated complex, as shown below:



The reaction is exothermic (ΔH is negative). Draw a potential energy curve for this process, labeling locations of the *reactant*, *activated complex*, and *product*. Then, draw arrows to indicate the magnitude of the activation energy, E_a , and ΔH for the reaction. (10 pts)



10. Describe a very simple technique that could be used to determine whether an aqueous mixture is a solution or a colloid. Which observation(s) would be made? (6 pts)

Simply shining a light through the solution will identify it as a colloid or solution. Colloids will scatter the light and the beam will be visible, where solutions will not scatter the light.

III. Problems

Choose 3 of the following problems to complete. If you answer more than three, only the first three will be graded. (20 points each)

I would like questions numbered _____, _____, and _____ to be graded.

11. Lauryl alcohol is a nonelectrolyte obtained from coconut oil and is used to make detergents. A solution of 5.00 g of lauryl alcohol in 0.100 kg of benzene freezes at 4.1 °C. What is the approximate molecular weight (in either amu or g/mol) of lauryl alcohol?

$$T_f(\text{benzene}) = 5.5\text{ }^{\circ}\text{C}$$

$$K_f(\text{benzene}) = 5.12\text{ }^{\circ}\text{C}/m$$

$$\Delta T_f = T_f(\text{benzene}) - T_f(\text{solution}) = 5.5\text{ }^{\circ}\text{C} - 4.1\text{ }^{\circ}\text{C} = 1.4\text{ }^{\circ}\text{C}$$

$$\Delta T_f = iK_f c_m$$

$$c_m = \frac{\Delta T_f}{iK_f} = \frac{1.4\text{ }^{\circ}\text{C}}{(1)(5.12\text{ }^{\circ}\text{C}/m)} = 0.2734375\text{ }m = \frac{0.2734375\text{ mol solute}}{1\text{ kg benzene}}$$

$$0.100\text{ kg benzene} \times \frac{0.2734375\text{ mol solute}}{1\text{ kg benzene}} = 0.02734375\text{ mol solute in this solution}$$

$$\text{MW} = \frac{g}{\text{mol}} = \frac{5.00\text{ g solute}}{0.02734375\text{ mol solute}} = 180\text{ g/mol}$$

12. The following data were collected for the rate of disappearance of NO in the reaction $2\text{NO}(g) + \text{O}_2(g) \longrightarrow 2\text{NO}_2(g)$:

Experiment	[NO] (M)	[O ₂] (M)	Initial rate (mol/L·s)
1	0.0126	0.0125	1.41×10^{-2}
2	0.0252	0.0250	1.13×10^{-1}
3	0.0252	0.0125	5.64×10^{-2}

- a. Based on the above information and using the method of initial rates, determine the reaction order with respect to NO, O₂, and overall.

Divide experiment 3 by experiment 1.

$$\frac{E3}{E1} = \frac{5.64 \times 10^{-2}}{1.41 \times 10^{-2}} = \left(\frac{0.0252}{0.0126} \right)^m \left(\frac{0.0125}{0.0125} \right)^n$$

$$4 = 2^m$$

$m = 2$; The reaction is 2nd order in NO.

Divide experiment 2 by experiment 3.

$$\frac{E2}{E3} = \frac{1.13 \times 10^{-1}}{5.64 \times 10^{-2}} = \left(\frac{0.0252}{0.0252} \right)^m \left(\frac{0.0250}{0.0125} \right)^n$$

$$2 = 2^n$$

$n = 1$; The reaction is 1st order in O₂.

$2 + 1 = 3$; The reaction is 3rd order overall.

- b. Write a rate law expression for this reaction.

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

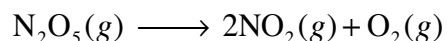
- c. Calculate the rate constant for this reaction, with correct units.

Substituting the values from experiment 1, we get:

$$\left(1.41 \times 10^{-2} \frac{\text{mol}}{\text{L} \cdot \text{s}} \right) = k \left(0.0126 \frac{\text{mol}}{\text{L}} \right)^2 \left(0.0125 \frac{\text{mol}}{\text{L}} \right)$$

$$k = 7110 \frac{\text{L}^2}{\text{mol}^2 \cdot \text{s}}$$

13. Consider the decomposition of N_2O_5 .



The **first-order** rate constant for this reaction at 70°C is $6.82 \times 10^{-3} \text{ s}^{-1}$. The reaction starts with 0.0250 mol of $\text{N}_2\text{O}_5(g)$ in a 2.0 L container.

- a. What is the molar concentration **and** number of moles of N_2O_5 remaining after 2.5 min?

$$[\text{A}]_0 = \frac{0.0250 \text{ mol}}{2.0 \text{ L}} = 0.0125 \text{ M}$$

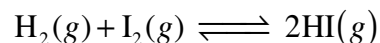
$$\begin{aligned} [\text{A}]_t &= [\text{A}]_0 e^{-kt} \\ &= (0.0125 \text{ M}) e^{(-6.82 \times 10^{-3} \text{ s}^{-1})(150 \text{ s})} \\ &= 0.00449 \text{ M at } 150 \text{ s} \end{aligned}$$

$$2.0 \text{ L} \times 0.00449 \frac{\text{mol}}{\text{L}} = 0.00898 \text{ mol at } 150 \text{ s}$$

- b. What is the half-life (in seconds) of this reaction?

$$t_{1/2} \text{ (1st order)} = \frac{-\ln \frac{1}{2}}{k} = \frac{-\ln \frac{1}{2}}{6.82 \times 10^{-3} \text{ s}^{-1}} = 102 \text{ s}$$

14. A closed system initially containing $1.000 \times 10^{-3} M$ H_2 and $2.000 \times 10^{-3} M$ I_2 at a certain given temperature is allowed to reach equilibrium, and the following equilibrium reaction takes place:



Analysis of the equilibrium mixture shows that the concentration of HI is $1.87 \times 10^{-3} M$.

- a. Using an ICE table, find the equilibrium molar concentrations of all components of this reaction.

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	$2HI(g)$
Initial (M)	$1.00 \times 10^{-3} M$		$2.00 \times 10^{-3} M$		$0 M$
Change (M)	$-x$		$-x$		$+2x$
Equilibrium (M)	$6.500 \times 10^{-5} M$		$1.06 \times 10^{-3} M$		$1.87 \times 10^{-3} M$

$$0 M + 2x = 1.87 \times 10^{-3} M$$

$$x = \frac{1.87 \times 10^{-3} M}{2} = 9.35 \times 10^{-4} M$$

$$[H_2]_{eq} = 1.000 \times 10^{-3} M - 9.35 \times 10^{-4} M = 6.500 \times 10^{-5} M$$

$$[I_2]_{eq} = 2.00 \times 10^{-3} M - 9.35 \times 10^{-4} M = 1.065 \times 10^{-3} M$$

- b. Determine the equilibrium constant, K_c , for this reaction.

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1.87 \times 10^{-3})^2}{(6.5 \times 10^{-5})(1.065 \times 10^{-3})} = 50.5$$

-END OF EXAM-