Name Key (Please Print)

Chem 104 - Section 1
Hour Examination II
Sample Test

This test consists of six (6) 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.

$R = 0.0821$ L·atm/K·mol = 8.314 J/K·mol K = °C + 273

$k = Ae^{-E_a/RT}$ $[A] = [A]_o(½)^h$ $t_{½} = 0.693/k$ $K_p = K_c(RT)^{\Delta n}$

DO NOT WRITE BELOW THIS LINE

This is a copy of a typical second test in Chem 104. 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 will not be covered on the test you take. Moreover, your test may have questions on topics that are not covered on this test. Posting this test in no way limits the format, style, scope, or level of the test that you will take. **Do not limit your preparation to the material on this sample test.**

Name Key

1. (44 points; 4 points each) Circle the best answer to each of the following.

a. As temperature is increased, the rate of a chemical reaction

increases   decreases   remains the same   may either increase or decrease

b. Consider the reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. If the rate of disappearance of $H_2(g)$, $-\Delta[H_2]/\Delta t$, is 2.4 mol/s, what is the rate of appearance of $NH_3(g)$, $+\Delta[NH_3]/\Delta t$?

0.80 mol/s   1.2 mol/s   1.6 mol/s   2.4 mol/s   3.6 mol/s

c. The reaction $A_2 \rightarrow 2A$ has the observed rate law $Rate = k[A_2]$. A plot of $\ln[A_2]$ vs. time gives a straight line whose slope is

$-k$ $-E_a/R$ $\ln A$ $\ln[A_2]_o$ k

d. The reaction $A_2 \rightarrow 2A$ has the observed rate law $Rate = k[A_2]$ and a half-life $t_{½} = 20.0$ s. If the starting concentration of A_2 is 0.96 mol/L, what concentration of A_2 will remain when the reaction is allowed to run for exactly one minute?

0.060 mol/L   0.12 mol/L   0.24 mol/L   0.32 mol/L   0.48 mol/L

e. Which one of the following is *not true* about a catalyst?

It changes the value of E_a.   It changes the value of k.   It changes the value of ΔH.

It changes the mechanism of the reaction.   It is not consumed in the overall reaction.

f. Consider the following mechanism for the hypothetical reaction $A_2 + 2B \rightarrow 2AB$:

$$A_2 \rightarrow 2A$$
$$A + B \rightarrow AB$$

In this mechanism, the species A is a(n)

catalyst   intermediate   product   inhibitor   reactant

Name Key

g. For the reaction $4HBr(g) + O_2(g) \rightarrow 2H_2O(g) + 2Br_2(g)$ the observed rate law is first order with respect to both HBr and O_2. Which one of the following proposed mechanisms is plausible? [All three mechanisms do add to the overall reaction stoichiometry when appropriate multiplications are made for certain steps; you don't have to worry about that.] (Circle just the Roman numeral for your answer.)

I. $O_2 \rightarrow 2O$ slow
$HBr + O \rightarrow HOBr$ fast
$HBr + HOBr \rightarrow H_2O + Br_2$ fast

II. $HBr + O_2 \rightarrow HOOBr$ slow
$HOOBr + HBr \rightarrow 2HOBr$ fast
$HBr + HOBr \rightarrow H_2O + Br_2$ fast

III. $2HBr + O_2 \rightarrow 2HOBr$ slow
$HBr + HOBr \rightarrow H_2O + Br_2$ fast

h. For the reaction $3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$, the expression for the equilibrium constant K_c would be

$\frac{[Fe_3O_4][H_2]^4}{[Fe]^3[H_2O]^4}$ $\frac{[Fe]^3[H_2O]^4}{[Fe_3O_4][H_2]^4}$ $\frac{[H_2]^4}{[H_2O]^4}$ $\frac{[H_2O]^4}{[H_2]^4}$ $[H_2]$

i. Of the following equilibria, which one will *shift to the left* in response to an increase in pressure?

$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $4Fe(s) + 3O_2(g) \rightleftharpoons 2Fe_2O_3(s)$

$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ $CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s)$

j. Consider the following reaction at equilibrium

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g) \quad \Delta H° = -514 \text{ kJ}$$

Increasing the temperature of this reaction will make the equilibrium constant

stay the same increase decrease change unpredictably

k. For the equilibrium $2NO_2(g) \rightleftharpoons N_2O_4(g)$, $K_c = 217 \text{ L·mol}^{-1}$ at 25 °C. What is the value of K_p?

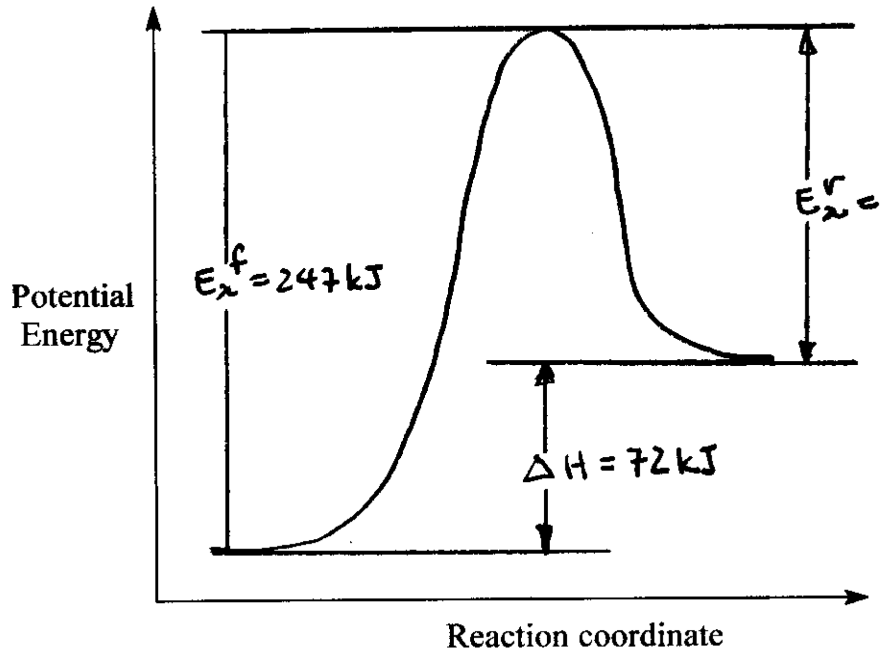
0.0876 atm^{-1} 8.87 atm^{-1} 106 atm^{-1} 445 atm^{-1} 5310 atm^{-1}

Name Key

2. (16 points) The reaction

$$C_2H_5Cl(g) \rightleftharpoons C_2H_4(g) + HCl(g)$$

proceeds by a one-step mechanism. The activation energy for the forward reaction (E_a^f) is 247 kJ and the enthalpy of the reaction (ΔH^o) is +72 kJ. In the space below, sketch *roughly to scale* the reaction energy profile (Arrhenius plot) for this reaction. On your sketch, clearly indicate (a) the relative energies of the reactants and products, (b) the activation energy of the forward and reverse reactions, and (c) the enthalpy of the reaction.



The numerical value of the activation energy of the reverse reaction (E_a^r) is 175 kJ.

Name Key

3. (20 points) For the equilibrium

$$Br_2(g) + Cl_2(g) \rightleftharpoons 2BrCl(g)$$

at 400 K, K_c is 7.0 If 0.076 mol $Br_2(g)$ and 0.076 mol $Cl_2(g)$ are introduced into a one-liter vessel at 400 K, what will be the concentrations of all species when equilibrium is established? Summarize your answers on the lines provided below, but be sure to show work that leads to your answers.

	$Br_2(g)$	+ $Cl_2(g)$	$\rightleftharpoons$ $2BrCl(g)$
Add	0.076	0.076	0
Get	0.076−x	0.076−x	2x

$$K_c = \frac{[BrCl]^2}{[Br_2][Cl_2]} = 7.0 = \frac{(2x)^2}{(0.076-x)^2}$$

Taking square roots of both sides:

$$2.6_{46} = \frac{2x}{0.076-x}$$

$$2x = 0.20_{11} - 2.6_{46}x$$

$$4.6_{46}x = 0.20_{11}$$

$$x = 0.043$$

$$\Rightarrow [Br_2] = [Cl_2] = 0.076 - 0.043 = 0.033 \text{ mol/L}$$

$$[BrCl] = (2)(0.043) = 0.086 \text{ mol/L}$$

Check:

$$Q = \frac{(0.086)^2}{(0.033)^2} = 6.79 \approx 7.0 = K_c$$

$[Br_2] = [Cl_2] =$ 0.033 mol/L $[BrCl] =$ 0.086 mol/L

Name Key

4. (20 points) Consider the hypothetical reaction

$$A_2 + B + C \rightarrow AB + AC$$

a. (15 points) Determine the differential rate law expression for the reaction from the following experimental data.

$4 \times B \Rightarrow 4 \times Rate$ $\therefore [B]^1$

Exp.	$[A_2]$, M	[B], M	[C], M	Rate, $M \cdot s^{-1}$
#1	0.125	0.111	0.702	1.07×10^{-3}
#2	0.500	0.111	0.702	2.14×10^{-3}
#3	0.125	0.444	0.702	4.28×10^{-3}
#4	0.125	0.444	0.351	4.28×10^{-3}

$4 \times [A_2] \Rightarrow 2 \times Rate$ $\therefore [A_2]^{1/2}$

$2 \times [C] \Rightarrow$ same Rate $\therefore [C]^0$

$Rate = k\,[A_2]^{1/2}[B]$

b. (5 points) Using data from Exp. #1, calculate the value of the rate constant, k, giving your answer with the proper units.

$$k = \frac{Rate}{[A_2]^{1/2}[B]} = \frac{1.07 \times 10^{-3}\,\cancel{M} \cdot s^{-1}}{(0.125M)^{1/2}(0.111\,\cancel{M})} = 0.0273\ M^{-1/2} \cdot s^{-1}$$

BONUS (5 points) The following mechanism has been proposed for this reaction.

$A_2 \rightleftharpoons 2A$	fast equilibrium	
$A + B \rightarrow AB$	slow	← Rate determining step
$A + C \rightarrow AC$	fast	

By deriving the rate law expression for this mechanism in terms of starting materials (not reaction intermediates), show that it is consistent with the experimentally determined rate law, found in part a. (Use the back of this page if you need more room.)

$Rate = k_2[A][B]$ But A is a reaction intermediate.

$K_1 = \frac{[A]^2}{[A_2]} \Rightarrow [A]^2 = K_1[A_2] \Rightarrow [A] = K_1^{1/2}[A_2]^{1/2}$

Substituting into $Rate = k_2[A][B]$

$Rate = k_2 K_1^{1/2}[A_2]^{1/2}[B] \equiv k[A_2]^{1/2}[B]$ q.e.d.