Chemistry 1142
Fall 2008
Exam 2

Name: KEY

Take a deep breath, and relax! First, answer the questions you know how to do and then work on the more difficult problems. Don't forget to show all your work, so I can give you as much credit as possible.

Good Luck!

Andy

"This is a lovely old song that tells of a young woman who leaves her cottage, and goes off to work. She arrives at her destination, and places some solid NH_4HS in a flask containing 0.50 atm of ammonia, and attempts to determine the pressures of ammonia and hydrogen sulfide when equilibrium is reached."

Multiple Choice. (3 pts. each.)

Q1. Which is the correct equilibrium constant expression for the following reaction?

$$Fe_2O_3(s) + 3H_2(g) \rightleftharpoons 2Fe(s) + 3H_2O(g)$$

Remember: Solids/liqs have an effective [ ] = 1

A. $K_c = [Fe_2O_3] [H_2]^3 / [Fe]^2[H_2O]^3$
B. $K_c = [H_2]/ [H_2O]$
C. $K_c = [H_2O]^3 / [H_2]^3$
D. $K_c = [Fe]^2[H_2O]^3 / [Fe_2O_3] [H_2]^3$
E. $K_c = [Fe] [H_2O] / [Fe_2O_3] [H_2]$

Q2. Consider the two gaseous equilibria

$SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g)$ $\quad K_1$

$2SO_3(g) \rightleftharpoons SO_2(g) + O_2(g)$ $\quad K_2$

doubled (square K) and reversed (invert K)

The values of the equilibrium constants K_1 and K_2 are related by

A. $K_2 = K_1^2$
B. $K_2 = \sqrt{K_1}$
C. $K_2 = -\frac{1}{2} K_1$
D. $K_2 = 1/K_1$
E. $K_2 = 1/(K_1^2)$

Q3. Which of the following is a true statement about chemical equilibria in general?

A. At equilibrium the total concentration of products equals the total concentration of reactants, that is, [products] = [reactants].
B. Equilibrium is the result of the cessation of all chemical change.
C. There is only one set of equilibrium concentrations that equals the K_c value.
D. At equilibrium, the rate constant of the forward reaction is equal to the rate constant for the reverse reaction.
E. At equilibrium, the rate of the forward reaction is equal to ~~as~~ the rate of the reverse reaction.

$$K_p = \underbrace{K_c}_{0.083}(\underbrace{R}_{0.08206\frac{atm\cdot L}{mol\cdot K}}\underbrace{T}_{1173K})^{\underbrace{\Delta n_g}_{+1}}$$

Q4. Phosgene, $COCl_2$, a poisonous gas, decomposes according to the equation

$COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g)$. Calculate K_p for this reaction if $K_c = 0.083$ at 900°C.

A. 0.125
B. 8.0
C. 6.1
D. 0.16
E. 0.083

Q5. For the following reaction at equilibrium, which one of the changes below would cause the equilibrium to shift to the *left*?

endothermic

"heat" + $2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)$, $\Delta H^\circ_{rxn} = +30$ kJ/mol

A. Increase the container volume.
B. Remove some NO.
C. Remove some Br_2.
D. Add more NOBr.
E. Decrease the temperature.

Q6. Consider the following reaction at equilibrium: (15 pts.)

$$A(g) \rightleftharpoons 2B(g)$$

From the following data, calculate the equilibrium constant, K_c at each temperature. (Show work!)

Is the reaction endothermic or exothermic? (No credit without a correct explaination.)

Temperature (°C)	[A] / M	[B] / M	K_c
200	0.0125	0.843	56.9
300	0.171	0.764	3.41
400	0.250	0.724	2.10

$K_c = \frac{[B]^2}{[A]}$

$K_c \downarrow$ as $T \uparrow$ $\Rightarrow$ eq^{m} must be exothermic!

$A \rightleftharpoons 2B$ + "heat"

as $T\uparrow$, "heat" $\uparrow$ $\Rightarrow$ Causes a shift to LHS.

$K_c \cong \frac{\text{Prods} \downarrow}{\text{Reacts} \uparrow} \Rightarrow K_c \downarrow$

Q7. 4.2 mol of oxygen and 4.0 mol of NO are introduced to an evacuated 0.50 L reaction vessel. At a specific temperature, the equilibrium $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ is reached when [NO] = 1.6 M. Calculate K_c for the reaction at this temperature? (15 pts.)

$[O_2]_0 = \frac{4.2\,mol}{0.50\,L} = 8.4\,M$, $[NO]_0 = \frac{4.0\,mol}{0.50\,L} = 8.0\,M$

	$2NO$	$+ O_2$	$\rightleftharpoons 2NO_2$
I	8.0	8.4	0
C	$-2x$	$-x$	$+2x$
E	$(8.0-2x)$	$(8.4-x)$	$(2x)$

@ eq^{m}, $[NO] = 1.6\,M = 8.0 - 2x$

$\Rightarrow 8.0 - 1.6 = 2x$

$\Rightarrow 6.4 = 2x$

$\Rightarrow x = 3.2$

$\Rightarrow [O_2]_{eq} = 8.4 - 3.2 = 5.2\,M$

$[NO_2]_{eq} = 2x = 6.4\,M$

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]} = \frac{6.4^2}{1.6^2 \times 5.2} = 3.1$$

Q8. A first-order chemical reaction has a rate-constant of 0.0542 s^{-1} at a temperature of −12 °C, and a rate-constant of 1.06 s^{-1} at a temperature of 21 °C. Assuming that the pre-exponential factor is the same at both temperatures, calculate the activation energy of this reaction. (15 pts.)

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_A}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\Rightarrow E_A = \frac{\ln(k_2/k_1) \times R}{\left(\frac{1}{T_1} - \frac{1}{T_2}\right)}$$

$$\left.\begin{array}{l} K_1 = 0.0542 s^{-1} \\ T_1 = -12 + 273 = 261 K \\ K_2 = 1.06 s^{-1} \\ T_2 = 21 + 273 = 294 K \\ R = 8.314 \text{ J/mol·K} \end{array}\right\} \quad E_A = \frac{\ln\left(\frac{1.06 s^{-1}}{0.0542 s^{-1}}\right) \times 8.314 \text{ J/mol·}\cancel{K}}{\left(\frac{1}{261 \cancel{K}} - \frac{1}{294 \cancel{K}}\right)}$$

$$\Rightarrow E_A = 57{,}500 \text{ J/mol}$$

$$= 57.5 \text{ kJ/mol}$$

Q9. Kinetic data for the following reaction was determined experimentally (20 Pts.)

$$2A_2(g) + B_2(g) \longrightarrow 2A_2B(g)$$

Experiment Number	Initial Conc $[A_2]$, M	Initial Conc $[B_2]$, M	Initial Rate of Reaction, $M \cdot s^{-1}$
1	0.10	0.60	3.10×10^{-4}
2	0.10	0.30	7.75×10^{-5}
3	0.20	0.60	6.20×10^{-4}

a) What is the rate law for the reaction? $rate = K[A_2]^x[B_2]^y$

$$\frac{rate(1)}{rate(2)} = \frac{3.10\times10^{-4} M\cdot s^{-1}}{7.75\times10^{-5} M\cdot s^{-1}} = \frac{K\cdot[0.10]^x\times[0.60M]^y}{K\cdot[0.10M]^x\times[0.30M]^y} \Rightarrow 4.00 = 2.0^y \Rightarrow y = 2$$

$$\frac{rate(3)}{rate(1)} = \frac{6.20\times10^{-4} M s^{-1}}{3.10\times10^{-4} M\cdot s^{-1}} = \frac{K[0.20M]^x\times[0.60M]^2}{K[0.10M]^x\times[0.60M]^2} \Rightarrow 2.00 = 2.0^x \Rightarrow x = 1$$

$$\Rightarrow rate = K[A_2][B_2]^2$$

b) Calculate the rate constant for the reaction. Be sure to include units.

$$K = \frac{rate}{[A_2][B_2]^2} \overset{xpt\#1}{=} \frac{3.10\times10^{-4} M\cdot s^{-1}}{[0.10M][0.60M]^2} = 8.61\times10^{-3} \frac{\cancel{M}\cdot s^{-1}}{\cancel{M}\cdot M^2}$$

$$= 8.61\times10^{-3} M^{-2}\cdot s^{-1}$$

Q10. The chemical reaction, $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ has an equilibrium constant of 7120 at 250 °C. If a flask containing only HCl at a concentration of 0.541 M at a temperature of 250 °C is allowed to reach equilibrium, then calculate the concentrations of H_2, Cl_2, and HCl. (20 pts.)

$$H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2HCl_{(g)}$$

	H_2	Cl_2	HCl
I	0	0	0.541
C	$+x$	$+x$	$-2x$
E	(x)	(x)	$(0.541-2x)$

$$K_c = \frac{[HCl]^2}{[H_2][Cl_2]}_{eq}$$

$$\Rightarrow 7120 = \frac{(0.541-2x)^2}{(x)(x)}$$

since this is a PERFECT SQUARE...

$$\sqrt{7120} = \frac{0.541-2x}{x}$$

$$\Rightarrow 84.38 \times x = 0.541 - 2x$$

$$\Rightarrow 86.38x = 0.541$$

$$\Rightarrow x = \frac{0.541}{86.38} = 0.00641$$

$$\Rightarrow [H_2]_{eq} = [Cl_2]_{eq} = x = 0.00641 M$$

$$\Rightarrow [HCl] = 0.541 - 2x = 0.528 M.$$

BONUS Questions:

1. Write formulas for the following compounds:

a) sodium phosphate Na_3PO_4

b) copper(I) carbonate Cu_2CO_3

c) heptasulfur nonafluoride S_7F_9

d) sulfuric acid H_2SO_4

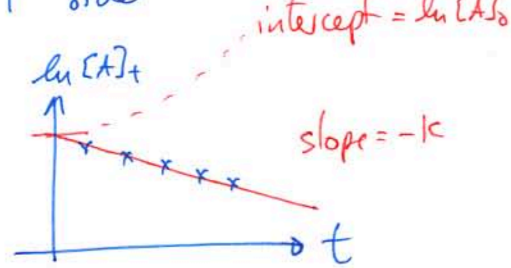
Na^{+} PO_4^{3-}

Cu^{+} CO_3^{2-}

hepta = 7
nona = 9

2. What graphs could you plot to determine whether a reaction has a rate law that is either first order or second order in a reactant, A.

$\ln[A]$ vs. t is linear if 1^{st} order:



WHY!

$$\ln[A]_t = -kt + \ln[A]_0$$
$$y = mx + b$$

$\frac{1}{[A]}$ vs. t is linear if 2^{nd} order:

WHY?

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$
$$y = mx + b$$