CHM 1410-002
Spring 2005
Test 2 (100 pts)

Name (Please Print) ______________ __________

Rate = $k[A]^x[B]^y$; $\ln [A]_t/[A]_o = -kt$; $\ln[A]_t = -kt + \ln[A]_o$; $t_{1/2} = 0.693/k$;

$1/[A]_t = kt + 1/[A]_o$; $k = A\exp(-E_a/RT)$ $\ln k = (-E_a/R)(1/T) + \ln A$;

$$\ln\frac{k1}{k2} = \frac{Ea}{R}\left(\frac{T1-T2}{T1T2}\right)$$

R = 8.31 J/mol.K; R = 0.082 L.atm/mole.K

Multiple Choice

1.(5) The Arrhenius equation is $k = A\,e^{-(Ea/RT)}$. The slope of a plot of ln k vs. 1/T is equal to

A. $-k$ B. k C. E_a D. $-E_a/R$ E. A

2.(5) Which is the correct equilibrium constant expression for the following reaction?
$Ni(s) + 4CO(g) \leftrightharpoons Ni(CO)_4(g)$

A. $K_c = [Ni(CO)_4]/[Ni][CO]^4$
B. $K_c = [Ni(CO)_4]/[Ni][CO]$
C. $K_c = [Ni(CO)_4]/[CO]^4$
D. $K_c = [Ni][CO]/[Ni(CO)_4]$
E. $K_c = [CO]^4/[Ni(CO)_4]$

3.(5) The equilibrium constant for the reaction $Ni(s) + 4CO(g) \leftrightharpoons Ni(CO)_4(g)$ is 5.0×10^4 at 25°C. What is the equilibrium constant for the reaction $Ni(CO)_4(g) \leftrightharpoons Ni(s) + 4CO(g)$?

A. 2.0×10^{-5} B. 2.5×10^9 C. 5.0×10^4 D. 5.0×10^{-4} E. 2.0×10^{-3}

4.(5) Calculate K_c for the reaction $2HI(g) \leftrightharpoons H_2(g) - I_2(g)$ given that upon analysis the following number of moles of each substance in a 2 liter vessel is found.
$HI = 1.70$ moles, $[I_2] = 1.20$ moles , $[H_2] = 0.54$ moles.

A. 5.25 B. 0.22 C. 4.5 D. 0.19 E. 1.6×10^2

5.(5) For the reaction $A + B \rightarrow C + D$, the activation energy of the uncatalyzed reaction is 45 kJ/mol. If a catalyst is added to this reaction, what is a feasible activation energy for the catalyzed reaction?

A. 50 kJ/mol
B. 45 kJ/mol
C. 40 kJ/mol
D. 0 kJ/mol
E. Less than 0 kJ/mol

For the following questions consider the reaction below at equilibrium:

$$2SO_2(g) + O_2(g) \leftrightharpoons 2SO_3(g), \Delta H^\circ_{rxn} = -198 \text{ kJ/mol.}$$

6.(3) Predict the direction of reaction if the container volume is increased.

A. To *products* or *to the right*
B. To *reactants* or *to the left*
C. No effect

7.(3) Predict the direction of reaction if some O_2 is removed.

A. To *products* or *to the right*
B. To *reactants* or *to the left*
C. No effect

8.(3) Predict the direction of reaction if some SO_2 is added.

A. To *products* or *to the right*
B. To *reactants* or *to the left*
C. No effect

9.(3) Predict the direction of reaction if the temperature is decreased.

A. To *products* or *to the right*
B. To *reactants* or *to the left*
C. No effect

10.(3) Predict the direction of reaction if a catalyst is added to the system.

A. To *products* or *to the right*
B. To *reactants* or *to the left*
C. No effect

Problems

1.(12) Sucrose, $C_{12}H_{22}O_{11}$, reacts slowly with water in the presence of an acid to form two other sugars, glucose and fructose, both of which have the same molecular formulas, but different structures.

$$C_{12}H_{22}O_{11} + H_2O \rightarrow C_6H_{12}O_6 \text{ (glucose)} + C_6H_{12}O_6 \text{ (fructose)}$$

The reaction is first order and has a rate constant of 6.2×10^{-5}/s at 35°C when the H^+ concentration is 0.10 M. Suppose that the initial concentration of sucrose in the solution is 0.40 M. How long will it take for 25% of the sucrose to decompose at 35°C?

25% of sucrose to decompose = 75% of sucrose remaining

$(0.75)0.40 = 0.30$ M sucrose remaining

$$\ln\left([\text{sucrose}]_t / [\text{sucrose}]_0\right) = -kt$$

$$-\frac{\ln\left([\text{sucrose}]_t / [\text{sucrose}]_0\right)}{k} = t$$

$$-\frac{\ln(0.30/0.40)}{6.2 \times 10^{-5}\ s^{-1}} = t$$

$$-\frac{-0.28}{6.2 \times 10^{-5}\ s^{-1}} = t$$

4,640 sec = t

2.(14) For the reaction $H_2(g) + I_2(g) \leftrightharpoons 2HI(g)$, $K_c = 50.2$ at 445°C, if $[H_2] = [I_2] = 1.75 \times 10^{-3}$ M, and $[HI] = 1 \times 10^{-2}$ M at 445°C, is the mixture at equilibrium? If not at equilibrium, in which direction (as the equation is written), *left to right* or *right to left*, will the reaction proceed to reach equilibrium? (Show work for credit).

$$Q_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1\times10^{-2})^2}{(1.75\times10^{-3})(1.75\times10^{-3})} = \frac{1\times10^{-4}}{3.06\times10^{-6}} = 32.7$$

$Q_c = 32.7$, $K_c = 50.2$

$Q < K$

so reaction must proceed to products (left to right) until $Q = K$

3.(14) The reaction $A(g) + 2B(g) \leftrightharpoons C(g)$ was allowed to come to equilibrium. The initial amounts of reactants placed into a 5.00 L vessel were 1.0 mol A and 1.8 mol B. After the reaction reached equilibrium, 1.0 mol of B was found. Calculate K_c for this reaction.

$[A]_0 = 1.00 \text{ mol}/5.00\text{ L} = 0.200$ M, $[B]_0 = 1.8 \text{ mols}/5.00\text{L} = 0.36$ M

$[B]_{eq} = 0.20$ M

	A	+ 2B	⇌ C
I	0.200	0.36	0
C	-0.08	-0.16	+0.08
E	0.12	0.20	0.08

$$K_c = \frac{[C]}{[A][B]^2} = \frac{0.08}{(0.12)(0.2)^2} = 16.67$$

4.(20) For the reaction $H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$ at 700°C, $K_c = 0.534$. Calculate the concentration of CO present at equilibrium if a mixture of 0.300 moles of H_2 and 0.300 moles of CO_2 is heated to 700°C in a 10.0 liter container.

$[H_2]_o = \dfrac{0.300 \text{ moles}}{10 L} = 0.0300\text{ M}$, $[CO_2]_o = \dfrac{0.300 \text{ moles}}{10 L} = 0.0300\text{ M}$

	H_2 +	CO_2	$\rightleftharpoons H_2O$ +	CO
i	0.03	0.03	0	0
c	$-x$	$-x$	$+x$	$+x$
e	$0.03-x$	$0.03-x$	x	x

$$K_c = 0.534 = \frac{[H_2O][CO]}{[H_2][CO_2]} = \frac{(x)(x)}{(0.03-x)(0.03-x)} = \frac{x^2}{(0.03-x)^2}$$

$$0.534 = \frac{x^2}{(0.03-x)^2}$$

could expand to quadratic, but this can be avoided in this case by taking square root of both sides

$$\sqrt{0.534} = \sqrt{\frac{x^2}{(0.03-x)^2}}$$

$$0.731 = \frac{x}{0.03-x}$$

$$0.0219 - 0.731x = x$$

$$0.0219 = 1.731x$$

$$\frac{0.0219}{1.731} = x$$

$$0.013 = x = [CO]$$