

CHM 1410-001
Spring 2005
Test 2 (100 pts)

Name (Please Print) _____

$$\text{Rate} = k[A]^x[B]^y; \quad \ln [A]_t/[A]_0 = -kt; \quad \ln [A]_t = -kt + \ln [A]_0; \quad t_{1/2} = 0.693/k;$$

$$1/[A]_t = kt + 1/[A]_0; \quad k = A \exp(-E_a/RT) \quad \ln k = (-E_a/R)(1/T) + \ln A;$$

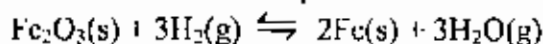
$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{T_1 - T_2}{T_1 T_2} \right) \quad R = 8.31 \text{ J/mol.K}; \quad R = 0.082 \text{ L.atm/mole.K}$$

Multiple Choice

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

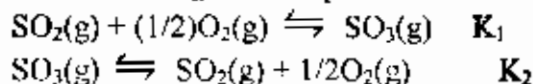
- A. $-k$ B. k C. E_a D. $-E_a/R$ E. A

2.(5) Which is the correct equilibrium constant expression for the following reaction?



- A. $K_c = [\text{Fe}_2\text{O}_3] [\text{H}_2]^3 / [\text{Fe}]^2 [\text{H}_2\text{O}]^3$
 B. $K_c = [\text{H}_2] / [\text{H}_2\text{O}]$
C. $K_c = [\text{H}_2\text{O}]^3 / [\text{H}_2]^3$
 D. $K_c = [\text{Fe}]^2 [\text{H}_2\text{O}]^3 / [\text{Fe}_2\text{O}_3] [\text{H}_2]^3$
 E. $K_c = [\text{Fe}] [\text{H}_2\text{O}] / [\text{Fe}_2\text{O}_3] [\text{H}_2]$

3.(5) Consider the two gaseous equilibria



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

- A. $K_2 = K_1^2$ B. $K_2^2 = K_1$ C. $K_2 = 1/K_1^2$ D. $K_2 = 1/K_1$
 E. none of these.

$$(S_2) = 0.2 \quad (H_2S) = 0.25 \\ (H_2) = 1$$

- 4.(5) On analysis, an equilibrium mixture for the reaction $2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$ was found to contain 1.0 mol H_2S , 4.0 mol H_2 , and 0.80 mol S_2 in a 4.0 L vessel. Calculate the equilibrium constant, K_c , for this reaction.

$$K = \frac{(1)^2(0.2)}{(1)^4}$$

- A. 1.6 ☒ B. 3.2 C. 12.8 D. 0.64 E. 0.8
- 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:



- 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 NO 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 NOBr 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) The rate constant for the first-order decomposition of C_4H_8 at $500^\circ C$ is $9.2 \times 10^{-3} s^{-1}$. How long will it take for 10.0% of a 0.100 M sample of C_4H_8 to decompose at $500^\circ C$?

$$10\% \text{ of } 0.100 M = 0.0100 M$$

$$C_4H_8 \text{ remaining} = 0.100 - 0.0100 = 0.090 M$$

$$\ln([A]_t/[A]_0) = -kt$$

$$- \frac{\ln([A]_t/[A]_0)}{k} = t ; - \frac{\ln([C_4H_8]_t/[C_4H_8]_0)}{k} = t$$

$$t = - \frac{\ln(0.09/0.10)}{9.2 \times 10^{-3} s^{-1}} = \frac{-(-0.105)}{9.2 \times 10^{-3} s^{-1}} = 11.4 \text{ sec.}$$

- 2.(14) At 700 K, the reaction $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ has the equilibrium constant $K_c = 4.3 \times 10^6$, and the following concentrations are present: $[\text{SO}_2] = 0.10 \text{ M}$; $[\text{SO}_3] = 10 \text{ M}$; $[\text{O}_2] = 0.10 \text{ M}$.

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).

$$Q_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{10^2}{(0.10)^2 (0.10)} = 1.0 \times 10^5$$

$$Q_c = 1.0 \times 10^5, K_c = 4.3 \times 10^6$$

$$Q_c < K_c$$

So Reaction proceeds to products
(left to right)

until $Q = K$

- 3.(14) 1.25 moles of NOCl were placed in a 2.50 L reaction chamber at 427°C. After equilibrium was reached, 1.10 moles of NOCl remained. Calculate the equilibrium constant, K_c , for the reaction $2\text{NOCl}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g})$.

$$[\text{NOCl}]_0 = 0.500 \text{ M}$$

	$2\text{NOCl} \rightleftharpoons 2\text{NO} + \text{Cl}_2$		
I	0.500	0	0
C	-0.06	+0.06	+0.03
E	0.44	0.06	0.03

$$K_c = \frac{[\text{NO}]^2 [\text{Cl}_2]}{[\text{NOCl}]^2} = \frac{(0.06)^2 (0.03)}{(0.44)^2}$$

$$K_c = 5.58 \times 10^{-4}$$

- 4.(20) For the reaction $\text{SO}_2(\text{g}) + \text{NO}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g}) + \text{NO}(\text{g})$, the equilibrium constant is 18.0 at $1,200^\circ\text{C}$. If 1.0 mole of SO_2 and 2.0 moles of NO_2 are placed in a 20. L container, what concentration of SO_3 will be present at equilibrium?

$$[\text{SO}_2]_0 = \frac{1 \text{ mole}}{20 \text{ L}} = 0.05 \text{ M}$$

$$[\text{NO}_2]_0 = \frac{2 \text{ moles}}{20 \text{ L}} = 0.10 \text{ M}$$

	SO_2	$+$	NO_2	\rightleftharpoons	SO_3	$+$	NO
I	0.05		0.10		0		0
C	-x		-x		+x		+x
E	0.05-x		0.10-x		x		x

$$K_c = 18 = \frac{[\text{SO}_3][\text{NO}]}{[\text{SO}_2][\text{NO}_2]} = \frac{(x)(x)}{(0.05-x)(0.10-x)}$$

$$18 = \frac{x^2}{5 \times 10^{-3} - 0.15x + x^2}$$

$$0.09 - 2.7x + 18x^2 = x^2$$

$$0.09 - 2.7x + 17x^2 = 0$$

$$a = 17, \quad b = -2.7, \quad c = 0.09$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{2.7 \pm \sqrt{(2.7)^2 - 4(17)(0.09)}}{2(17)}$$

$$x = \frac{2.7 \pm \sqrt{7.29 - 6.12}}{34} = \frac{2.7 \pm 1.08}{34}$$

$$x = 0.11 \quad \text{or} \quad x = 0.047$$

~~$x = 0.11$~~
gives
negative
 SO_2 conc.

$$x = 0.047 \text{ M} = [\text{SO}_3]$$