

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

Name (Please Print) _____

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

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

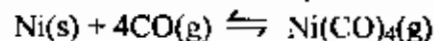
$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{T_1 - T_2}{T_1 T_2} \right) \quad R = 8.31 \text{ J/mol.K}; \quad R = 0.082 \text{ L.atm/mole.K}$$

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. $K_c = [\text{Ni(CO)}_4]/[\text{Ni}][\text{CO}]^4$
 B. $K_c = [\text{Ni(CO)}_4]/[\text{Ni}][\text{CO}]$
☒ C. $K_c = [\text{Ni(CO)}_4]/[\text{CO}]^4$
 D. $K_c = [\text{Ni}][\text{CO}]/[\text{Ni(CO)}_4]$
 E. $K_c = [\text{CO}]^4/[\text{Ni(CO)}_4]$

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

☒ 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 $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ given that upon analysis the following number of moles of each substance in a 2 liter vessel is found.

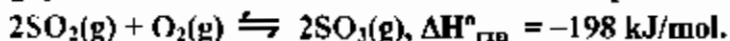
$\text{HI} = 1.70$ moles, $[\text{I}_2] = 1.20$ moles, $[\text{H}_2] = 0.54$ moles.

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

- 5.(5) For the reaction $\text{A} + \text{B} \rightarrow \text{C} + \text{D}$, 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 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.

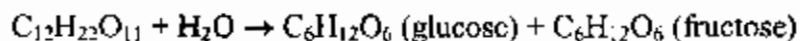
- ☒ A. To products or to the right
B. To reactants or to the left
C. No effect

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.



The reaction is first order and has a rate constant of $6.2 \times 10^{-5}/s$ at $35^\circ C$ 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 $35^\circ C$?

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

$(0.75) 0.40 = 0.30 \text{ M sucrose remaining}$

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

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

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

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

$$4,640 \text{ sec} = t$$

- 2.(14) For the reaction $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$, $K_c = 50.2$ at 445°C , if $[\text{H}_2] = [\text{I}_2] = 1.75 \times 10^{-3} \text{ M}$, and $[\text{HI}] = 1 \times 10^{-2} \text{ M}$ 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).

$$Q_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(1 \times 10^{-2})^2}{(1.75 \times 10^{-3})(1.75 \times 10^{-3})} = \frac{1 \times 10^{-4}}{3.06 \times 10^{-6}} = 32.7$$

$$Q_c = 32.7, K_c = 50.2$$

$$Q < K$$

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

- 3.(14) The reaction $\text{A}(\text{g}) + 2\text{B}(\text{g}) \rightleftharpoons \text{C}(\text{g})$ 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.

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

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

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

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

4.(20) For the reaction $\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g})$ 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.

$$[\text{H}_2]_0 = \frac{0.300 \text{ moles}}{10\text{L}} = 0.0300 \text{ M}, \quad [\text{CO}_2]_0 = \frac{0.300 \text{ moles}}{10\text{L}} = 0.0300 \text{ M}$$

	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

$$K_c = 0.534 = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} = \frac{(x)(x)}{(0.03-x)(0.03-x)} = \frac{x^2}{(0.03-x)^2}$$

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

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

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

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

$$0.0219 - 0.731x = x$$

$$0.0219 = 1.731x$$

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

$$0.013 = x = [\text{CO}]$$