Practice Test Chapter 15

Equilibrium

***Part I:*** *Circle the letter of the best annswer(s). No calculators are allowed on the multiple choice portion of the practice test!*

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| Target 1: I can write the equilibrium expression for a balanced chemical equation. |

1. At equilibrium . . .

a. all chemical processes have ceased.

b. the rate of the forward reaction equals that of the reverse reaction.

c. the rate constant of the forward reaction equals the rate constant of the

reverse reaction.

d. both the rate of the forward reaction equals that of the reverse reaction and

the rate constant for the forward reaction equals the rate constant of the

reverse reaction.

e. None of the above.

2. Considering the equation below, which expression is correct for Keq?

2 SO2 (g) + O2 (g) **⮀** 2 SO3 (g)



d.



3. What reaction has the equilibrium expression . . . [D]2 / [A] [B] [C]?

* 1. 2 D **⮀** A + B + C
  2. 2 D **⮀** ABC
  3. A + B + C **⮀** 2 D
  4. ABC **⮀** 2 D

4. Considering the equation below, which expression is correct for *Keq*?

CuS(s) + O2(g) **⮀** Cu(s) + SO2(g)

a. *Keq* = b. *Keq* =



c. *Keq* = d. *Keq* =



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| Target 2: I can numerically evaluate Kc and Kp from a knowledge of the equilibrium concentrations (or pressures) of reactants or products, or from the initial concentrations and the equilibrium concentration of at least one substance. |

5. The relationship between the rate constants for the forward and reverse

reactions and the equilibrium constant for a process is . . .

a. Keq = Kf • Kr

b. Keq = Kf - Kr

c. Keq = Kf + Kr

d. Keq = Kf / Kr

e. Keq = 1/(Kf • Kr)

6. The equilibrium constant for reaction (1) below is 0.02. The value of the

equilibrium constant for reaction (2) is \_\_\_\_\_\_.

(1) 3A + 2B **⮀** 2D + E

(2) 2D + E **⮀** 3A + 2B

a. 0.02

b. 0.04

c. -0.02

d. 50

e. 3.0 X 103

7. Which **one** of the following will change the value of an equilibrium constant?

a. Changing the temperature of the reaction.

b. Adding other substances that do not react with any of the species

involved in the equilibrium.

c. Varying initial concentration of reactants.

d. Varying initial concentration of products.

e. All of these will affect the value of the equilibrium constant.

8. Consider the following equilibrium: 2 SO2(g) + O2(g) **⮀** 2 SO3(g)

From which of the following starting conditions would it be impossible for this

equilibrium to be achieved?

a. 1.0 mol SO3(g) in a 1.0 liter container

b. 0.25 mol SO2(g), 0.25 mol O2(g) in a 1.0 liter container

c. 0.25 mol SO2(g), 0.50 mol O2(g), and 0.10 mol SO3(g) in a 1.0 liter

container

d. 0.50 mol O2(g) and 0.50 mol SO3(g) in a 1.0 liter container

e. Equilibrium can be achieved from any of these starting conditions

9. The reaction: X + Y **⮀** 2 M has a Kc = 0.89 at 672 K. At

equilibrium\_\_\_\_\_\_.

a. products predominate substantially

b. reactants predominate substantially

c. roughly equal molar amounts of products and reactants are present

d. only products exist

e. only reactants exist

10. Consider the following reaction: CO(g) + 2 H2 (g) **⮀** CH3OH (g)

In an experiment, 0.200 mol of CO and 0.300 mol of H2 were placed in a

1.00-L reaction vessel. At equilibrium, there were 0.100 mol of CO

remaining. Calculate the Keq of the equilibrium.

a. 0.0100

b. 1.00

c. 100.

d. 1.00 X 104

e. 1.00 X 107

11. In the coal-gasification process, carbon monoxide is converted to carbon dioxide via the following reaction:

CO (g) + H2O (g) **⮀** CO2 (g) + H2(g)

In an experiment, 3.00 mol of CO and 5.00 mol of H2O were placed in a

2.00-L reaction vessel. At equilibrium, there were 2.00 mol of CO remaining.

Keq at the temperature of the experiment is \_\_\_.

a. 0.125

b. 0.250

c. 0.500

d. 1.00

e. 2.00

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| Target 3: I can interconvert Kc and Kp. |

12. For the reaction below, Kp = 2.1 x 104 at 400 K.  
 2 A(g) + B(s) **⮀** C(g) + D(g)

What is the value of Kc for the above reaction at this temperature?

a. 4.2 x 108 b. 3.5 x 101 c. 1.0 x 109 d. 0.50 x 104 e. 2.1 x 104

13. For the following reaction at 25oC, Kc = 3.0 x 105. What is Kp? (You can use a

calculator for this one!)

2 H2S(g) + 3 O2(g) **⮀** 2 H2O(g) + 2 SO2(g)

a. 1.2 x 104

b. 8.2 x 10-5

c. 3.3 x 10-6

d. 3.0 x 105

e. 7.3 x 106

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| Target 4: I can use the equilibrium constant to calculate equilibrium concentrations. |

14. Consider the following equilibrium at 27°C:

NH4HS(s) **⮀** NH3(g) + H2S(g); Kp = 0.16

A sample of solid NH4HS is placed in a closed vessel and allowed to

equilibrate. Calculate the equilibrium partial pressure (atm) of ammonia,

assuming that some solid NH4HS remains.

a. 0.40

b. 0.16

c. 0.32

d. 0.64

e. 0.0032

15. Consider the following equilibrium: 2 AB(g) **⮀** A2(g) + B2(g) ; Kc = 16

Assume that the initial concentration of AB is 2.0 M and that there is no A2 or B2

originally present. Calculate the equilibrium concentration of B2.

a. 0.0034

b. 0.00049

c. 1.5

d. -0.104

e. 0.89

16. Consider the following reaction:

CuS(s) + O2(g) **⮀** Cu(s) + SO2(g)

A reaction mixture initially contains 3.0 M O2. Determine the equilibrium concentration of O2 if Kc for the reaction at this temperature is 2.0.

a. 0.10 M

b. 0.20 M

c. 0.50 M

d. 1.0 M

e. 2.0 M

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| Target 5: I can calculate the reaction quotient, Q, and by comparison with the value of Kc or Kp determine whether a reaction is at equilibrium. If it is not at equilibrium, I can predict in which direction it will shift to reach equilibrium. |

17. Consider the following reaction and its equilibrium constant:

4 CuO(s) + CH4(g) **⮀** CO2(g) + 4 Cu(s) + 2 H2O(g); Kc = 0.100

A reaction mixture contains 2.0 M CH4, 0.50 M CO2 and 0.80 M H2O. Which of the following statements is TRUE concerning this system?

a. The reaction will shift in the direction of reactants.

b. The equilibrium constant will increase.

c. The reaction quotient will increase.

d. The reaction will shift in the direction of products.

e. The system is at equilibrium.

18. Consider the following reaction:

I2(g) + Br2(g) **⮀** 2 IBr(g) Kc = 1.1 x 102 ΔH = 40.88 kJ

A 3.0 liter vessel was found to contain 3.0 moles of I2, 6.0 moles of Br2 and 3.0 moles of IBr. Is the system at equilibrium? If not, which reaction will be favored in order for the reaction to achieve equilibrium?

a. Yes, the reaction is at equilibrium.

b. No, the endothermic reaction will be favored in order to reach equilibrium.

c. No, the exothermic reaction will be favored in order to reach equilibrium.

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| Target 6: I can explain how the relative equilibrium quantities of reactants and products are shifted by changes in temperature, pressure or the concentrations of substances in the equilibrium reaction. |

19. Consider the following endothermic reaction:

CaCO3(s) **⮀** CaO(s) + CO2(g)

Which of the following would shift the equilibrium position to form more

CO2 gas?

a. Increasing the system temperature.

b. Decreasing the system temperature.

c. Increasing the system pressure.

d. Decreasing the volume of the system.

e. Adding neon gas to the system.

20. In which of the following reactions would increasing pressure at constant

temperature not change the concentration of reactants and products.?

a. N2(g) + 3 H2(g) **⮀** 2 NH3(g)

b. N2H4(g) **⮀** 2 NO2(g)

c. N2(g) + 2 O2(g) **⮀** 2 NO2(g)

d. 2 N2(g) + O2(g) **⮀** 2 N2O(g)

e. N2(g) + O2(g) **⮀** 2 NO(g)

21. Consider the following reaction at equilibrium:

2 NH3(g) **⮀** N2(g) + 3 H2(g) ; ∆Ho = +92.4 kJ

This reaction will shift to the right with \_\_\_\_\_\_\_\_\_\_\_\_.

a. increasing both temperature and pressure

b. increasing temperature and decreasing pressure

c. decreasing both temperature and pressure

d. decreasing temperature and increasing pressure

e. the addition of extra N2 to the container

22. Consider the following reaction at equilibrium:

2 SO2(g) + O2(g) **⮀** 2 SO3(g) ΔH° = -99 kJ

Le Châtelier's Principle predicts that an increase in temperature will result

in \_\_\_\_\_\_\_\_\_\_.

a. a decrease in the partial pressure of SO2

b. a decrease in the partial pressure of SO3

c. an increase in *Keq*

d. the partial pressure of O2 will decrease

e. no changes in equilibrium partial pressures

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| Target 7: I can describe the effect of a catalyst on a system as it approaches equilibrium. |

23. The effect of a catalyst on an equilibrium is to \_\_\_\_\_\_\_\_\_\_.

a. increase the rate of the forward reaction only

b. increase the equilibrium constant so that products are favored

c. slow the reverse reaction only

d. increase the rate at which equilibrium is achieved without changing the composition of the equilibrium mixture

e. shift the equilibrium to the right

**Part II:** Solve each of the following problems on separate sheets of paper. Show all of your work and label your answer with appropriate units. Please clearly indicate your final answer by circling it!

1. The compound butane, C4H10 , occurs in two isomeric forms, *n*-butane and

isobutane (2-methyl propane). Both compounds exist as gases at 25°C and

1.0 atm.

(a) Draw the structural formula of each of the isomers (include all atoms). Clearly

label each structure.

(b) On the basis of molecular structure, identify the isomer that has the higher

boiling point. Justify your answer.

The two isomers exist in equilibrium as represented by the equation below.

*n*-butane(*g*) **⮀** isobutane(*g*) *Kc* = 2.5 at 25°C

Suppose that a 0.010 mol sample of pure *n*-butane is placed in an evacuated 1.0 L

rigid container at 25°C.

(c) Write the expression for the equilibrium constant, *Kc* , for the reaction.

(d) Calculate the initial pressure in the container when the *n*-butane is first

introduced (before the reaction starts).

(e) The *n*-butane reacts until equilibrium has been established at 25°C.

(i) Calculate the total pressure in the container at equilibrium. Justify your

answer.

(ii) Calculate the molar concentration of each species at equilibrium.

(iii) If the volume of the system is reduced to half of its original volume, what

will be the new concentration of *n*-butane after equilibrium has been

reestablished at 25°C? Justify your answer.

Suppose that in another experiment a 0.010 mol sample of pure isobutane is placed in an evacuated 1.0 L rigid container and allowed to come to equilibrium at 25°C.

(f) Calculate the molar concentration of each species after equilibrium has been

established.

2. Consider the following reaction:

C(*s*) + CO2(*g*) **⮀** 2 CO(*g*)

Solid carbon and carbon dioxide gas at 1,160 K were placed in a rigid 2.00 L container, and the reaction represented above occurred. As the reaction proceeded, the total pressure in the container was monitored. When equilibrium was reached, there was still some C(*s*) remaining in the container. Results are recorded in the table below.

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| Time  (hours) | Total pressure of the gases in the container at 1,160 K |
| 0.0 | 5.00 |
| 2.0 | 6.26 |
| 4.0 | 7.09 |
| 6.0 | 7.75 |
| 8.0 | 8.37 |
| 10.0 | 8.37 |

(a) Write the expression for the equilibrium constant, *Kp* , for the reaction.

(b) Calculate the number of moles of CO2(*g*) initially placed in the container. (Assume

that the volume of the solid carbon is negligible.)

(c) For the reaction mixture at equilibrium at 1,160 K, the partial pressure of the

CO2(*g*) is 1.63 atm. Calculate:

(i) the partial pressure of CO(*g*) , and

(ii) the value of the equilibrium constant, *Kp* .

(d) If a suitable solid catalyst were placed in the reaction vessel, would the final total pressure of the gases at equilibrium be greater than, less than, or equal to the final total pressure of the gases at equilibrium without the catalyst? Justify your answer. (Assume that the volume of the solid catalyst is negligible.)

In another experiment involving the same reaction, a rigid 2.00 L container initially contains 10.0 g of C(*s*), plus CO(*g*) and CO2(*g*), each at a partial pressure of 2.00 atm at 1,160 K.

(e) Predict whether the partial pressure of CO2(*g*) will increase, decrease, or remain the same as this system approaches equilibrium. Justify your prediction with a calculation.

3. Answer the following questions regarding the decomposition of arsenic pentafluoride,

AsF5(*g*) .

(a) A 55.8 g sample of AsF5(*g*) is introduced into an evacuated 10.5 L

container at 105°C.

(i) What is the initial molar concentration of AsF5(*g*) in the container?

(ii) What is the initial pressure, in atmospheres, of the AsF5(*g*) in the

container? At 105°C, AsF5(*g*) decomposes into AsF3(*g*) and F2(*g*)

according to the following chemical equation.

AsF5(*g*) **⮀** AsF3(*g*) + F2(*g*)

(b) In terms of molar concentrations, write the equilibrium-constant

expression for the decomposition of AsF5(*g*).

(c) When equilibrium is established, 27.7 percent of the original number of

moles of AsF5(*g*) has decomposed.

(i) Calculate the molar concentration of AsF5(*g*) at equilibrium.

(ii) Using molar concentrations, calculate the value of the equilibrium

constant, *Keq*, at 105°C.

(d) Calculate the mole fraction of F2(*g*) in the container at equilibrium.