**Equilibrium Practice Problems**

**Multiple Choice**

*Identify the letter of the choice that best completes the statement or answers the question.*

\_\_\_\_ 1. Calculate the equilibrium constant K for the following equilibrium

3 F2(g) + Cl2(g)  2 ClF3(g)

with the following equilibrium concentrations:

 [F2]= 2.0 M; [Cl2] = 2.5 M; [ClF3] = 3.0 M

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 0.40 | c. | 1.2 |
| b. | 0.90 | d. | 0.45 |

\_\_\_\_ 2.

H2*(g)* + Br2*(g)*  2 HBr*(g) K =* 

At a certain temperature, the value of the equilibrium constant, *K*, for the reaction represented above is equal to . What is the value of *K* for the following reaction at the same temperature?

 6 HBr*(g)*   *3* H2*(g)* + 3Br2*(g) K = ?*

|  |  |  |  |
| --- | --- | --- | --- |
| a. |  | c. |  |
| b. | 1/ | d. | 1/ |

\_\_\_\_ 3.

HCO3ñ*(aq)* + OH–*(aq)*  H2O*(l)* + CO32ñ*(aq)* *H* = –41.4 kJ

When the reaction represented by the equation above is at equilibrium at 1 atm and 25°C, the ratio $\frac{[CO\_{3}^{2-}]}{[HCO\_{3}^{-}]}$ can be increased by doing which of the following?

|  |  |
| --- | --- |
| a. | Decreasing the temperature |
| b. | Addin Adding acid |
| c. | Adding a catalyst |
| d. | Diluting the solution with distilled water |

\_\_\_\_ 4.

2 NO*(g)* + O2*(g)*  2 NO2*(g)* Δ*H* < 0

Which of the following changes alone would cause a decrease in the value of *Keq* for the reaction represented above?

|  |  |
| --- | --- |
| a. | Decreasing the temperature |
| b. | Increasing the temperature |
| c. | Decreasing the volume of the reaction vessel |
| d. | Adding a catalyst |

\_\_\_\_ 5. PCl3*(g)* + Cl2*(g)*  PCl5*(g)* + energy

 Some PCl3 and Cl2 are mixed in a container at 200oC and the system reaches equilibrium according to the equation above. Which of the following causes an increase in the number of moles of PCl5 present at equilibrium?

|  |  |
| --- | --- |
| a. | Decreasing the volume of the container |
| b. | Raising the temperature |
| c. | Adding a mole of He gas at constant volume  |
| d. | Increasing the volume of the container |

\_\_\_\_ 6. 2 SO2*(g)* + O2*(g)*  2 SO3*(g)*

 When 0.40 mole of SO2 and 0.60 mole of O2 are placed in an evacuated 1.00–liter flask, the reaction represented above occurs. After the reactants and the product reach equilibrium and the initial temperature is restored, the flask is found to contain 0.30 mole of SO3. Based on these results, the equilibrium constant, *Kc,* for the reaction is

|  |  |  |  |
| --- | --- | --- | --- |
| a. | $$\frac{(0.30)^{2}}{(0.45)(0.10)^{2}}$$ | c. | $$\frac{(2 × 0.30)}{(0.45)(2 × 0.10)}$$ |
| b. | $$\frac{(0.30)^{2}}{(0.60)(0.40)^{2}}$$ | d. | $$\frac{(0.30)}{(0.45)(0.10)}$$ |

\_\_\_\_ 7. CuO*(s)* + H2*(g)*  Cu*(s)* + H2O*(g)* *H* = –2.0 kilojoules

When the substances in the equation above are at equilibrium at pressure P and temperature T, the equilibrium can be shifted to favor the products by

|  |  |
| --- | --- |
| a. | increasing the pressure by means of a moving piston at constant T |
| b. | increasing the pressure by adding an inert gas such as nitrogen |
| c. | decreasing the temperature |
| d. | allowing some gases to escape at constant P and T |

**Problem**

 8. Consider the following reaction. (2003)

 2 HI(g)  H2(g) + I2(g)

After a 1.0 mole sample of HI(g) is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of HI(g) as a function of time is shown below.



a) Write the expression for the equilibrium constant, Kc, for the reaction. (1 pt)

b) What is [HI] at equilibrium? (1 pt)

c) Determine the equilibrium concentrations of H2(g) and I2(g). (2 pts)

d) On the graph above, make a sketch that shows how the concentration of H2(g) changes as a function of time. (2 pts)

e) Calculate the value of the equilibrium constant Kc at 700. K. (1 pt)

f) At 1000 K, the value of Kc for the reaction is 2.6 x 10-2. In an experiment, 0.75 mole of HI(g), 0.10 mol of H2(g), and 0.50 mol of I2(g) are placed in a 1.0 L container and allowed to reach equilibrium at 1000 K. Determine whether the equilibrium concentration of HI(g) will be greater than, less than, or equal to the initial concentration of HI(g). Justify your answer. (2 pts)

 9. 2 H2S*(g)*  2 H2*(g)* + S2*(g)*

When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of H2S*(g)* is introduced into an evacuated rigid 1.25 L container. The sealed container is heated to 483 K, and 3.7210ñ2 mol of S2*(g)* is present at equilibrium. (2000)

(a) Write the expression for the equilibrium constant, *K*c, for the decomposition reaction represented above. (1 pt)

(b) Calculate the equilibrium concentration, in molL-1, of the following gases in the container at 483 K.

(i) H2*(g)* (1 pt)

(ii) H2S*(g)* (2 pts)

(c) Calculate the value of the equilibrium constant, *K*c, for the decomposition reaction at 483 K. (2 pts)

(d) Calculate the partial pressure of S2*(g)* in the container at equilibrium at 483 K. (2 pts)

 (e) For the reaction H2*(g)* + 1/2 S2*(g)*  H2S*(g)* at 483 K, calculate the value of the equilibrium constant, *K*c. (2 pts)

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 10. 1995 A

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

When H2*(g)* is mixed with CO2*(g)* at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

 [H2] = 0.20 mol/L

 [CO2] = 0.30 mol/L

 [H2O] = [CO] = 0.55 mol/L

(a) What is the mole fraction of CO*(g)* in the equilibrium mixture?

(b) Using the equilibrium concentrations given above, calculate the value of *Kc*, the equilibrium constant for the reaction.

(c) When the system is cooled from 2,000 K to a lower temperature, 30.0 percent of the CO*(g)* is converted back to CO2*(g)*. Calculate the value of *Kc* at this lower temperature.

 (e) In a different experiment, 0.50 mole of H2*(g)* is mixed with 0.50 mole of CO2*(g)* in a 3.0-liter reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of CO*(g)* at this temperature.

 11. The polyatomic ion C10H12N2O84- is commonly abbreviated as EDTA4-. The ion can form complexes with metal ions in aqueous solutions. A complex of EDTA4- with Ba2+ ions forms according to the equation below. (2016)

 Ba2+(aq) + EDTA4-(aq)  Ba(EDTA)2-(aq) K = 7.7 x 107

A 50.0 mLvolume of a solution that has an EDTA4-(aq) concentration of 0.30 M is mixed with 50.0 mL of 0.20 M Ba(NO3)2 to produce 100.0 mL of solution.

a) Considering the value of K for the reaction, determine the concentration of Ba(EDTA)2-(aq) in the 100.0 mL of solution. Justify your answer. (2 pts)

b) The solution is diluted with distilled water to a total volume of 1.00 L. After equilibrium has been reestablished, is the number of moles of Ba2+(aq) present in the solution greater than, less than, or equal to the number of moles of Ba2+(aq) present in the original solution before it was diluted? Justify your answer. (2 pts)

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**Equilibrium Practice Problems**

**Answer Section**

**MULTIPLE CHOICE**

 1. ANS: D PTS: 1

 2. ANS: D PTS: 1

 3. ANS: A PTS: 1

 4. ANS: B PTS: 1

 5. ANS: A PTS: 1

 6. ANS: A PTS: 1

 7. ANS: C PTS: 1

**PROBLEM**

 8. ANS:

a) Kc = [H2][I2]/[HI]2

b) from the graph, [HI]eq = 0.80 M

c) ICE problem, or use stoichiometry

 [I2] = [H2] = 0.10 M

d) [H2]eq equals 0.10 M; starts at 0M, goes to 0.10 M at same time HI reaches equillibrium

e) i) Kc = 0.016

 ii) Kc = Kp because #moles reactant = # moles product

f) calculate Q = 8.9 x 10-2

 find that Q >Kc too much product, not enough reactant so [HI] will increase

PTS: 1

 9. ANS:

(a) *K*c =

(b) (i)  = 5.9510–2 *M* H2

 (ii)  = 2.0510–2 *M* H2S

(c) *K*c = = 0.251

(d) PV=nRT = 1.18

(e) *K*’c =  = 2.00

PTS: 1

 10. ANS:

(a) CO = f(0.55 mol, 1.6 mol) = 0.34

(b) *Kc* = ([H2O][CO])/([H2][CO2]) = (0.550.55)/(0.200.30) = 5.04

(c) [CO] = 0.55 - 30.0% = 0.55 - 0.165 = 0.385 M

 [H2O] = 0.55 - 0.165 = 0.385 M

 [H2] = 0.20 + 0.165 = 0.365 M

 [CO2] = 0.30 + 0.165 = 0.465 M

 K = (0.385)2/(0.3650.465) = 0.87

(e) let *X* = [H2] to reach equilibrium

 [H2] = 0.50 mol/3.0L - X = 0.167 - *X*

 [CO2] = 0.50 mol/3.0L - X = 0.167 - *X*

 [CO] = +*X* ; [H2O] = +*X*

 K = *X*2/(0.167 - *X*)2 = 5.04 ; *X* = [CO] = 0.12 M

PTS: 1

 11. ANS:

DO NOT PUT THIS ON A UNIT EXAM--OK for a problem set, though

a) Based on the K value, the reaction goes essentially to completing. Barium ion is the LR.

The concentration of barium ion when the solutions are first mixed but before any reaction takes place is 0.10 M (the volume is doubled so the concentration is halved).

Thus the equilibrium concentration of the complex is 0.10 M.

 (1 pt for stating that the [complex] = [Ba2+ after mixing] , 1 pt for figuring out the concentration with an appropriate calculation)

b) The number of moles of barium ion INCREASES (1 pt)

because the percent dissociation of complex increases as the solution is diluted OR a Q vs. K calculation showing that Q > K and therefore the system will shift left in order to reestablish equilibrium (1 pt for an appropriate justification)

PTS: 1