**1977 D**

For the system 2 SO2*(g)* + O2*(g)* ↔ 2 SO3*(g)* , Δ*H* is negative for the production of SO3. Assume that one has an equilibrium mixture of these substances. Predict the effect of each of the following changes on the value of the equilibrium constant and on the number of moles of SO3 present in the mixture at equilibrium. Briefly account for each of your predictions. (Assume that in each case all other factors remain constant.)

(a) Decreasing the volume of the system.

(b) Adding oxygen to the equilibrium mixture.

(c) Raising the temperature of the system.

**1980 D**

NH4Cl*(s)* ↔ NH3*(g)* + HCl*(g)* Δ*H* = +42.1 kilocalories

Suppose the substances in the reaction above are at equilibrium at 600K in volume V and at pressure P. State whether the partial pressure of NH3*(g)* will have increased, decreased, or remained the same when equilibrium is reestablished after each of the following disturbances of the original system. Some solid NH4Cl remains in the flask at all times. Justify each answer with a one-or-two sentence explanation.

(a) A small quantity of NH4Cl is added.

(b) The temperature of the system is increased.

(c) The volume of the system is increased.

(d) A quantity of gaseous HCl is added.

(e) A quantity of gaseous NH3 is added.

**1981 A**

Ammonium hydrogen sulfide is a crystalline solid that decomposes as follows:

NH4HS*(s)* ↔ NH3*(g)* + H2S*(g)*

(a) Some solid NH4HS is placed in an evacuated vessel at 25ºC. After equilibrium is attained, the total pressure inside the vessel is found to be 0.659 atmosphere. Some solid NH4HS remains in the vessel at equilibrium. For this decomposition, write the expression for KP and calculate its numerical value at 25ºC.

(b) Some extra NH3 gas is injected into the vessel containing the sample described in part (a). When equilibrium is reestablished at 25ºC, the partial pressure of NH3 in the vessel is twice the partial pressure of H2S. Calculate the numerical value of the partial pressure of NH3 and the partial pressure of H2S in the vessel after the NH3 has been added and the equilibrium has been reestablished.

(c) In a different experiment, NH3 gas and H2S gas are introduced into an empty 1.00 liter vessel at 25ºC. The initial partial pressure of each gas is 0.500 atmospheres. Calculate the number of moles of solid NH4HS that is present when equilibrium is established.

**1983 A**

Sulfuryl chloride, SO2Cl2, is a highly reactive gaseous compound. When heated, it decomposes as follows: SO2Cl2*(g)* → SO2*(g)* + Cl2*(g)*. This decomposition is endothermic. A sample of 3.509 grams of SO2Cl2 is placed in an evacuated 1.00 litre bulb and the temperature is raised to 375K.

(a) What would be the pressure in atmospheres in the bulb if no dissociation of the SO2Cl2*(g)* occurred?

(b) When the system has come to equilibrium at 375K, the total pressure in the bulb is found to be 1.43 atmospheres. Calculate the partial pressures of SO2, Cl2, and SO2Cl2 at equilibrium at 375K.

(c) Give the expression for the equilibrium constant (either Kp or Kc) for the decomposition of SO2Cl2*(g)* at 375K. Calculate the value of the equilibrium constant you have given, and specify its units.

(d) If the temperature were raised to 500K, what effect would this have on the equilibrium constant? Explain briefly.

**1985 A**

At 25ºC the solubility product constant, Ksp, for strontium sulfate, SrSO4, is 7.6×10-7. The solubility product constant for strontium fluoride, SrF2, is 7.910-10.

(a) What is the molar solubility of SrSO4 in pure water at 25ºC?

(b) What is the molar solubility of SrF2 in pure water at 25ºC?

(c) An aqueous solution of Sr(NO3)2 is added slowly to 1.0 litre of a well-stirred solution containing 0.020 mole F- and 0.10 mole SO42- at 25ºC. (You may assume that the added Sr(NO3)2 solution does not materially affect the total volume of the system.)

 1. Which salt precipitates first?

2. What is the concentration of strontium ion, Sr2+, in the solution when the first precipitate begins to form?

(d) As more Sr(NO3)2 is added to the mixture in (c) a second precipitate begins to form. At that stage, what percent of the anion of the first precipitate remains in solution?

**1988 A**

At elevated temperatures, SbCl5 gas decomposes into SbCl3 gas and Cl2 gas as shown by the following equation:

SbCl5*(g)* ↔ SbCl3*(g)* + Cl2*(g)*

(a) An 89.7 gram sample of SbCl5 (molecular weight 299.0) is placed in an evacuated 15.0 litre container at 182ºC.

1. What is the concentration in moles per litre of SbCl5 in the container before any decomposition occurs?

2. What is the pressure in atmospheres of SbCl5 in the container before any decomposition occurs?

(b) If the SbCl5 is 29.2 percent decomposed when equilibrium is established at 182ºC, calculate the value for either equilibrium constant Kp or Kc, for this decomposition reaction. Indicated whether you are calculating Kp or Kc.

(c) In order to produce some SbCl5, a 1.00 mole sample of SbCl3 is first placed in an empty 2.00 litre container maintained at a temperature different from 182ºC. At this temperature, Kc, equals 0.117. How many moles of Cl2 must be added to this container to reduce the number of moles of SbCl3 to 0.700 mole at equilibrium?

**1988 D**

NH4HS*(s)* ↔ NH3*(g)* + H2S*(g)* ΔHº = +93 kilojoules

The equilibrium above is established by placing solid NH4HS in an evacuated container at 25ºC. At equilibrium, some solid NH4HS remains in the container. Predict and explain each of the following.

(a) The effect on the equilibrium partial pressure of NH3 gas when additional solid NH4HS is introduced into the container

(b) The effect on the equilibrium partial pressure of NH3 gas when additional solid H2S is introduced into the container

(c) The effect on the mass of solid NH4HS present when the volume of the container is decreased

(d) The effect on the mass of solid NH4HS present when the temperature is increased.

**1992 A**

2 NaHCO3*(s)* ↔ Na2CO3*(s)* + H2O*(g)* + CO2*(g)*

Solid sodium hydrogen carbonate, NaHCO3, decomposes on heating according to the equation above.

(a) A sample of 100. grams of solid NaHCO3 was placed in a previously evacuated rigid 5.00-liter container and heated to 160ºC. Some of the original solid remained and the total pressure in the container was 7.76 atmospheres when equilibrium was reached. Calculate the number of moles of H2O*(g)* present at equilibrium.

(b) How many grams of the original solid remain in the container under the conditions described in (a)?

(c) Write the equilibrium expression for the equilibrium constant, KP, and calculate its value for the reaction under the conditions in (a).

(d) If 110. grams of solid NaHCO3 had been placed in the 5.00-liter container and heated to 160ºC, what would the total pressure have been at equilibrium? Explain.

**1994 A**

MgF2*(s)* ↔ Mg2+*(aq)* + 2 F-*(aq)*

In a saturated solution of MgF2 at 18ºC, the concentration of Mg2+ is 1.2110-3 molar. The equilibrium is represented by the equation above.

(a) Write the expression for the solubility-product constant, *Ksp,* and calculate its value at 18ºC.

(b) Calculate the equilibrium concentration of Mg2+ in 1.000 liter of saturated MgF2 solution at 18ºC to which 0.100 mole of solid KF has been added. The KF dissolves completely. Assume the volume change is negligible.

(c) Predict whether a precipitate of MgF2 will form when 100.0 milliliters of a 3.0010-3-molar Mg(NO3)2 solution is mixed with 200.0 milliliters of a 2.00l0-3-molar NaF solution at 18ºC. Calculations to support your prediction must be shown.

(d) At 27ºC the concentration of Mg2+ in a saturated solution of MgF2 is 1.1710-3 molar. Is the dissolving of MgF2 in water an endothermic or an exothermic process? Give an explanation to support your conclusion.

**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) Determine *Kp* in terms of *Kc* for this system.

(d) 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.

**1998 D**

C*(s)* + H2O*(g)* ↔ CO*(g)* + H2*(g)* Δ*Hº* = +131kJ

A rigid container holds a mixture of graphite pellets (C*(s)*), H2O*(g)*, CO*(g)*, and H2*(g)* at equilibrium. State whether the number of moles of CO*(g)* in the container will increase, decrease, or remain the same after each of the following disturbances is applied to the original mixture. For each case, assume that all other variables remain constant except for the given disturbance. Explain each answer with a short statement.

(a) Additional H2*(g)* is added to the equilibrium mixture at constant volume.

(b) The temperature of the equilibrium mixture is increased at constant volume.

(c) The volume of the container is decreased at constant temperature.

(d) The graphite pellets are pulverized.

**2000 A Required**

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10–2 mol of S2*(g)* is present at equilibrium.

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

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

(i) H2*(g)*

(ii) H2S*(g)*

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

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

(e) For the reaction H2*(g)* + S2*(g)* → H2S*(g)* at 483 K, calculate the value of the equilibrium constant, *K*c.

**2007 *part A*, form B, question #1**

A sample of solid U308 is placed in a rigid 1.500 L flask. Chlorine gas, Cl2*(g)*, is added, and the flask is heated to 862˚C. The equation for the reaction that takes place and the equilibrium-constant expression for the reaction are given below.

U308*(s)* + 3 Cl2*(g)*, ↔ 3 UO2Cl2*(g)* + O2*(g)* 

When the system is at equilibrium, the partial pressure of Cl2*(g)* is 1.007 atm and the partial pressure of UO2Cl2*(g)* is 9.734×10-4 atm

(a) Calculate the partial pressure of O2*(g)* at equilibrium at 862˚C.

(b) Calculate the value of the equilibrium constant, *KP*, for the system at 862˚C.

(c) Calculate the Gibbs free-energy change, ∆*G*˚, for the reaction at 862˚C.

(d) State whether the entropy change, ∆*S*˚ for the reaction at 862˚C is positive, negative, or zero. Justify your answer.

(e) State whether the enthalpy change, ∆*H*˚, for the reaction at 862˚C is positive, negative, or zero. Justify your answer.

(f) After a certain period of time, 1.000 mol of O2*(g)* is added to the mixture in the flask. Does the mass of U308*(s)* in the flask increase, decrease, or remain the same? Justify your answer.