1977 D

For the system 2 $SO_{2(g)}$ + $O_{2(g)} \leftrightarrow$ 2 $SO_{3(g)}$, ΔH is negative for the production of SO_3. 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 SO_3 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.

Answer:

(a) As volume decreases, pressure increases and the reaction shifts in the direction of fewer molecules (less volume; more SO_3) to relieve the stress. Value of K_{eq} does not change.
(b) Additional O_2 disturbs the equilibrium and SO_3 is formed to relieve the stress. Value of K_{eq} does not change.
(c) Increase in temperature shifts the reaction to the left to "use up" some of the added heat. Less SO_3 remains. Value of K_{eq} decreases due to the relative greater increase in the rate of the endothermic reaction (reaction to the left).

1980 D

$NH_4Cl_{(s)} \leftrightarrow NH_{3(g)} + HCl_{(g)}$ ΔH = +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 $NH_{3(g)}$ will have increased, decreased, or remained the same when equilibrium is reestablished after each of the following disturbances of the original system. Some solid NH_4Cl remains in the flask at all times. Justify each answer with a one-or-two sentence explanation.

(a) A small quantity of NH_4Cl 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 NH_3 is added.

Answer:

(a) P_{NH_3} does not change. Since $NH_4Cl_{(s)}$ has constant concentration (a = 1), equilibrium does not shift.
(b) P_{NH_3} increases. Since the reaction is endothermic, increasing the temperature shifts the equilibrium to the right and more NH_3 is present.
(c) P_{NH_3} does not change. As V increases, some solid NH_4Cl decomposes to produce more NH_3. But as the volume increases, P_{NH3} remains constant due to the additional decomposition.
(d) P_{NH_3} decreases. Some NH_3 reacts with the added HCl to relieve the stress from the HCl addition.
(e) P_{NH_3} increases. Some of the added NH_3 reacts with HCl to relieve the stress, but only a part of the added NH_3 reacts, so P_{NH3} increases.

1981 A

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

$$NH_4HS_{(s)} \leftrightarrow NH_{3(g)} + H_2S_{(g)}$$

(a) Some solid NH_4HS 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 NH_4HS remains in the vessel at equilibrium. For this decomposition, write the expression for K_P and calculate its numerical value at 25ºC.

(b) Some extra NH_3 gas is injected into the vessel containing the sample described in part (a). When equilibrium is reestablished at 25ºC, the partial pressure of NH_3 in the vessel is twice the partial pressure of H_2S. Calculate the numerical value of the partial pressure of NH_3 and the partial pressure of H_2S in the vessel after the NH_3 has been added and the equilibrium has been reestablished.

(c) In a different experiment, NH_3 gas and H_2S 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 NH_4HS that is present when equilibrium is established.

Answer:

(a) $K_P = (P_{NH3})(P_{H2S})$

$P_{NH3} = P_{H2S} = 0.659/2$ atm $= 0.330$ atm

$K_P = (0.330)^2 = 0.109$

(b) $P_{NH3} = 2\ P_{H2S}$

$(2x)(x) = 0.109$; $x = 0.233$ atm $= P_{H2S}$

$P_{NH3} = 0.466$ atm

(c) Equilibrium pressures of NH_3 and H_2S are each 0.330 atm. Amounts of each NH_3 and H_2S that have reacted correspond to (0.500 - 0.330) = 0.170 atm.

n = mol of each reactant = mol of solid product

$$n = \frac{PV}{RT} = \frac{(0.170\,\text{atm})(1.00\,\text{L})}{(0.08205\,\frac{\text{L\_atm}}{\text{mol\_K}})(298\text{K})} = 6.95 \infty 10^{-3}\,\text{mol}$$

1983 A

Sulfuryl chloride, SO_2Cl_2, is a highly reactive gaseous compound. When heated, it decomposes as follows: $SO_2Cl_{2(g)} \rightarrow SO_{2(g)} + Cl_{2(g)}$. This decomposition is endothermic. A sample of 3.509 grams of SO_2Cl_2 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 $SO_2Cl_{2(g)}$ 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 SO_2, Cl_2, and SO_2Cl_2 at equilibrium at 375K.

(c) Give the expression for the equilibrium constant (either K_p or K_c) for the decomposition of $SO_2Cl_{2(g)}$ 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.

Answer:

(a)

$$P = \frac{nRT}{V} = \frac{\left(3.509\ \text{g} \times \frac{1\ \text{mol}}{135.0\text{g}}\right)\left(0.0821\ \frac{\text{L\_atm}}{\text{mol\_K}}\right)(375\text{K})}{1.00\text{L}}$$

= 0.800 atm

(b) $P_{SO2Cl2} = (0.800 - y)$ atm

$P_{SO2} = P_{Cl2} = y$ atm

$P_T = P_{SO2Cl2} + P_{SO2} + P_{Cl2}$

1.43 atm = (0.800 - y + y + y) atm

y = 0.63 atm = $P_{SO2} = P_{Cl2}$

P_{SO2Cl2} = (0.800 - 0.63) atm = 0.17 atm

$$K_p = \frac{(P_{SO_2})(P_{Cl_2})}{P_{SO_2Cl_2}} = \frac{(0.63\,\text{atm})^2}{0.17\,\text{atm}} = 2.3\text{atm}$$

(c)

(d) Heat is absorbed during the dissociation and so $K_{500} > K_{375}$. A stress is place on the system and K increases, which reduces the stress associated with the higher temperature.

1985 A

At 25ºC the solubility product constant, K_{sp}, for strontium sulfate, $SrSO_4$, is 7.6×10^{-7}. The solubility product constant for strontium fluoride, SrF_2, is 7.9×10^{-10}.

(a) What is the molar solubility of $SrSO_4$ in pure water at 25ºC?

(b) What is the molar solubility of SrF_2 in pure water at 25ºC?

(c) An aqueous solution of $Sr(NO_3)_2$ is added slowly to 1.0 litre of a well-stirred solution containing 0.020 mole F^- and 0.10 mole SO_4^{2-} at 25ºC. (You may assume that the added $Sr(NO_3)_2$ solution does not materially affect the total volume of the system.)

1. Which salt precipitates first?
2. What is the concentration of strontium ion, Sr^{2+}, in the solution when the first precipitate begins to form?

(d) As more $Sr(NO_3)_2$ 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?

Answer:

(a) $SrSO_{4(s)} \leftrightarrow Sr^{2+}(aq) + SO_4^{2-}(aq)$

At equilibrium: $[Sr^{2+}] = X$ M $= [SO_4^{2-}]$

$X^2 = K_{sp} = 7.6\times10^{-7}$

$X = 8.7\times10^{-4}$ mol/L, solubility of $SrSO_4$

(b) $SrF_{2(s)} \leftrightarrow Sr^{2+}(aq) + 2\ F^-(aq)$

At equilibrium: $[Sr^{2+}] = X$ M $= [F^-] = 2X$ M

$K_{SP} = [Sr^{2+}][F^-]^2 = (X)(2X)^2 = 7.9\times10^{-10}$

$X = 5.8\times10^{-4}$ mol/L, solubility of SrF_2

(c) Solve for $[Sr^{2+}]$ required for precipitation of each salt.

$K_{sp} = [Sr^{2+}][F^-]^2 = 7.9\times10^{-10}$

$$= (x)\left(\frac{(0.020\,\text{mol})^2}{1.0\,\text{L}}\right) = 7.9\times10^{-10}\ ;\ x = 2.0\times10^{-6}\,\text{M}$$

$K_{sp} = [Sr^{2+}][SO_4^{2-}] = 7.6\times10^{-7}$

$$= (y)\left(\frac{0.10\,\text{mol})}{1.0\,\text{L}}\right) = 7.6\times10^{-7}\ ;\ y = 7.6\times10^{-6}\,\text{M}$$

Since 2.0×10^{-6} M < 7.6×10^{-6} M, SrF_2 must precipitate first.

When SrF_2 precipitates, $[Sr^{2+}] = 2.0\times10^{-6}$ M

(d) The second precipitate to form is $SrSO_4$, which appears when $[Sr^{2+}] = 7.6\times10^{-6}$ M (based on calculations in Part c.)

When $[Sr^{2+}] = 7.6\times10^{-6}$ M, $[F^-]$ is determined as follows:

$K_{sp} = [Sr^{2+}][F^-]^2 = 7.9\times10^{-10}$

$= (7.6\times10^{-6})(z)^2 = 7.9\times10^{-10}$; $z = 1.0\times10^{-2}$ M

$$\%\ F^-\ \text{still in solution} = \frac{1.0\times10^{-2}}{2.0\times10^{-2}}\times100 = 50.\%$$

1988 A

At elevated temperatures, $SbCl_5$ gas decomposes into $SbCl_3$ gas and Cl_2 gas as shown by the following equation: $SbCl_{5(g)} \leftrightarrow SbCl_{3(g)} + Cl_{2(g)}$

(a) An 89.7 gram sample of $SbCl_5$ (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 $SbCl_5$ in the container before any decomposition occurs?
2. What is the pressure in atmospheres of $SbCl_5$ in the container before any decomposition occurs?

(b) If the $SbCl_5$ is 29.2 percent decomposed when equilibrium is established at 182ºC, calculate the value for either equilibrium constant K_p or K_c, for this decomposition reaction. Indicated whether you are calculating K_p or K_c.

(c) In order to produce some $SbCl_5$, a 1.00 mole sample of $SbCl_3$ is first placed in an empty 2.00 litre container maintained at a temperature different from 182ºC. At this temperature, K_c, equals 0.117. How many moles of Cl_2 must be added to this container to reduce the number of moles of $SbCl_3$ to 0.700 mole at equilibrium?

Answer:

(a) (1)

$$89.7\,\text{g SbCl}_5\times\frac{1\ \text{mol}}{299.0\,\text{g}} = 0.300\ \text{mol SbCl}_5$$

$$[\text{SbCl}_5]_{init.} = \frac{0.300\ \text{mol}}{15.0\,\text{L}} = 0.200\,\text{M}$$

(2) T = 182ºC + 273 = 455K

$$P = \frac{nRT}{V} = \frac{(0.300\,\text{mol})\left(0.0821\,\frac{\text{L\_atm}}{\text{mol\_K}}\right)(455\text{K})}{15.00\,\text{L}}$$

= 0.747 atm

OR

$$\frac{15.0\text{L}}{0.300\text{mol}} \infty \frac{273\text{K}}{455\text{K}} = \frac{30.0\text{L}}{\text{mol}} \text{ at std. temp.}$$

1 atm × f(22.4 L/mol,30.0 L/mol) = 0.747 atm

(b) $[SbCl_3] = [Cl_2]$ = (0.0200 mol/L)(0.292)

= 5.84×10^{-3}M

$[SbCl_5]$ = (0.0200 mol/L)(0.708) = 1.42×10^{-2}M

$$K_c = \frac{[SbCl_3][Cl_2]}{[SbCl_5]} = \frac{(5.84 \infty 10^{-3})^2}{1.42 \infty 10^{-2}} = 2.41 \infty 10^{-3}$$

OR

$P_{SbCl3} = P_{Cl2}$ = (0.747 atm)(0.292) = 0.218 atm

P_{SbCl5}= (0.747 atm)(0.708) = 0.529 atm

$$K_P = \frac{(P_{SbCl_3})(P_{Cl_2})}{P_{SbCl_5}} = \frac{(0.218)^2}{0.529} = 8.98 \infty 10^{-2}$$

(c)

$$K = \frac{[SbCl_3][Cl_2]}{[SbCl_5]} = 0.117$$

$$[SbCl_5]_{equil.} = \frac{(1.00 - 0.70)\text{mol}}{2.00\text{L}} = 0.15\text{M}$$

$$[SbCl_3]_{equil.} = \frac{0.700\text{ mol}}{2.00\text{L}} = 0.350\text{M}$$

$[Cl_2]_{equil.}$ = X M

$$K_C = \frac{(0.350)(X)}{0.15} = 0.117;\ X = 0.050\text{M} = [Cl_2]$$

Moles $Cl_{2(equil.)}$ = (0.0050 mol/L)(2.00L)

= 0.10 mol Cl_2

Moles Cl_2 needed to make 0.300 mol $SbCl_3$ into $SbCl_5$ = 0.30 mol

Moles Cl_2 that must be added = 0.40 mol

1988 D

$NH_4HS(s) \leftrightarrow NH_3(g) + H_2S(g)$ ΔH° = +93 kilojoules

The equilibrium above is established by placing solid NH_4HS in an evacuated container at 25°C. At equilibrium, some solid NH_4HS remains in the container. Predict and explain each of the following.

(a) The effect on the equilibrium partial pressure of NH_3 gas when additional solid NH_4HS is introduced into the container

(b) The effect on the equilibrium partial pressure of NH_3 gas when additional solid H_2S is introduced into the container

(c) The effect on the mass of solid NH_4HS present when the volume of the container is decreased

(d) The effect on the mass of solid NH_4HS present when the temperature is increased.

Answer:

(a) The equilibrium pressure of NH_3 gas would be unaffected. $K_P = (P_{NH3})(P_{H2S})$. Thus the amount of solid NH_4HS present does not affect the equilibrium.

(b) The equilibrium pressure of NH_3 gas would decrease. In order for the pressure equilibrium constant, K_P, to remain constant, the equilibrium pressure of NH_3 must decrease when the pressure of H_2S is increased. $K_P = (P_{NH3})(P_{H2S})$. (A complete explanation based on LeChatelierÆs principle is also acceptable.)

(c) The mass of NH_4HS increases. A decrease in volume causes the pressure of each gas to increase. To maintain the value of the pressure equilibrium constant, K_p, the pressure of each of the gases must decrease. The decrease is realized by the formation of more solid NH_4HS. $K_p = (P_{NH3})(P_{H2S})$. (A complete explanation based on LeChatelier's principle is also acceptable.)

(d) The mass of NH_4HS decreases because the endothermic reaction absorbs heat and goes nearer to completion (to the right) as the temperature increases.

1992 A

$$2\ NaHCO_{3(s)} \leftrightarrow Na_2CO_{3(s)} + H_2O_{(g)} + CO_{2(g)}$$

Solid sodium hydrogen carbonate, $NaHCO_3$, decomposes on heating according to the equation above.

(a) A sample of 100. grams of solid $NaHCO_3$ 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 $H_2O_{(g)}$ 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, K_P, and calculate its value for the reaction under the conditions in (a).

(d) If 110. grams of solid $NaHCO_3$ had been placed in the 5.00-liter container and heated to 160ºC, what would the total pressure have been at equilibrium? Explain.

Answer:

$$n_{gas} = \frac{PV}{RT} = \frac{(7.76\,atm)(5.00\,L)}{(0.0821\,\frac{L\_atm}{mol\_K})(433K)} = 1.09mol$$

(a)

mol H_2O = $(^1/_2)$(1.09 mol) = 0.545 mol $H_2O_{(g)}$

(b)

$$0.545 mol\ H_2O \infty \frac{2mol\ NaHCO_3}{1mol\ H_2O} \infty \frac{84.0gNaHCO_3}{1\ mol} =$$

= 91.9 g $NaHCO_3$ decomposed

remaining = 100.g - 91.6g = 8.4g

OR

100 -

$$\left(0.545\,molH_2O \infty \frac{18.0g}{1\ mol} + 0.545\,molCO_2 \infty \frac{44.0g}{1\ mol}\right)$$

= 100g - 33.8g = 66g (or 66.2g) [includes Na_2CO_3 solid in this mass]

(c) $K_p = (P_{H2O})(P_{CO2}) = (3.88)(3.88)atm^2 = 15.1\ atm^2$

(d) Pressure would remain at 7.76 atm. Since some solid remained when 100.g was used (and there has been no temperature change), then using 110g will not affect the equilibrium.

1994 A

$$MgF_{2(s)} \leftrightarrow Mg^{2+}(aq) + 2\ F^{-}(aq)$$

In a saturated solution of MgF_2 at 18ºC, the concentration of Mg^{2+} is 1.21×10^{-3} molar. The equilibrium is represented by the equation above.

(a) Write the expression for the solubility-product constant, K_{sp}, and calculate its value at 18ºC.

(b) Calculate the equilibrium concentration of Mg^{2+} in 1.000 liter of saturated MgF_2 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 MgF_2 will form when 100.0 milliliters of a 3.00×10^{-3}-molar $Mg(NO_3)_2$ solution is mixed with 200.0 milliliters of a $2.00\times l0^{-3}$-molar NaF solution at 18ºC. Calculations to support your prediction must be shown.

(d) At 27ºC the concentration of Mg^{2+} in a saturated solution of MgF_2 is 1.17×10^{-3} molar. Is the dissolving of MgF_2 in water an endothermic or an exothermic process? Give an explanation to support your conclusion.

Answer:

(a) $K_{sp} = [Mg^{2+}][F^-]^2 = (1.21\times10^{-3})(2.42\times10^{-3})^2$

$= 7.09\times10^{-9}$

(b) X = concentration loss by Mg^{2+} ion

$2X$ = concentration loss by F^- ion

$[Mg^{2+}] = (1.21\times10^{-3} - X)$ M

$[F^-] = (0.100 + 2.42\times10^{-3} - 2X)$ M

since X is a small number then $(0.100 + 2.42\times10^{-3} - 2X) \approx 0.100$

$K_{sp} = 7.09\times10^{-9} = (1.21\times10^{-3} - X)(0.100)^2$

$X = 1.2092914\times10^{-3}$

$[Mg^{2+}]= 1.21\times10^{-3}-1.20929\times10^{-3} = 7.09\times10^{-7}M$

(c) $[Mg^{2+}] = 3.00\times10^{-3}M \times 100.0\ mL/300.0\ mL = 1.00\times10^{-3}M$

$[F^-] = 2.00\times10^{-3}\ M \times 200.0\ mL/300.0\ mL = 1.33\times10^{-3}\ M$

trial $K_{sp} = (1.00\times10^{-3})(1.33\times10^{-3})^2 = 1.78\times10^{-9}$

trial $K_{sp} < = 7.09\times10^{-9}$, $\therefore$ no ppt.

(d) @ 18ºC, 1.21×10^{-3} M MgF_2 dissolves

@ 27ºC, 1.17×10^{-3} M MgF_2 dissolves

$MgF_2 \leftrightarrow Mg^{2+} + 2\ F^- +$ heat

dissolving is exothermic; if heat is increased it forces the equilibrium to shift left (according to LeChatelier’s Principle) and less MgF_2 will dissolve.

1995 A

$$CO_{2}(g) + H_{2}(g) \leftrightarrow H_2O(g) + CO(g)$$

When $H_2(g)$ is mixed with $CO_2(g)$ at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

$[H_2]$ = 0.20 mol/L

$[CO_2]$ = 0.30 mol/L

$[H_2O] = [CO] = 0.55$ 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 K_c, the equilibrium constant for the reaction.

(c) Determine K_p in terms of K_c 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 $CO_2(g)$. Calculate the value of K_c at this lower temperature.

(e) In a different experiment, 0.50 mole of $H_2(g)$ is mixed with 0.50 mole of $CO_2(g)$ 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.

Answer:

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

(b) $K_c = ([H_2O][CO])/([H_2][CO_2]) = (0.55\times0.55)/(0.20\times0.30) = 5.04$

(c) since $\Delta n = 0$, $K_c = K_p$

(d) $[CO] = 0.55 - 30.0\% = 0.55 - 0.165 = 0.385$ M

$[H_2O] = 0.55 - 0.165 = 0.385$ M

$[H_2] = 0.20 + 0.165 = 0.365$ M

$[CO_2] = 0.30 + 0.165 = 0.465$ M

$K = (0.385)^2/(0.365\times0.465) = 0.87$

(e) let $X = \Delta[H_2]$ to reach equilibrium

$[H_2] = 0.50 \text{ mol}/3.0\text{L} - X = 0.167 - X$

$[CO_2] = 0.50 \text{ mol}/3.0\text{L} - X = 0.167 - X$

$[CO] = +X$; $[H_2O] = +X$

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

1998 D

$C(s) + H_2O(g) \leftrightarrow CO(g) + H_2(g) \qquad \Delta H^o = +131\text{kJ}$

A rigid container holds a mixture of graphite pellets ($C(s)$), $H_2O(g)$, $CO(g)$, and $H_2(g)$ 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 $H_2(g)$ 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.

Answer

(a) CO will decrease. An increase of hydrogen gas molecule will increase the rate of the reverse reaction which consumes CO. A LeChatelier Principle shift to the left.

(b) CO will increase. Since the forward reaction is endothermic (a $\Delta H > 0$) an increase in temperature will cause the forward reaction to increase its rate and produce more CO. A LeChatelier Principle shift to the right.

(c) CO will decrease. A decrease in volume will result in an increase in pressure, the equilibrium will shift to the side with fewer gas molecules to decrease the pressure, ∴, a shift to the left.

(d) CO will remain the same. Once at equilibrium, the size of the solid will affect neither the reaction rates nor the equilibrium nor the concentrations of reactants or products.

2000 A Required

1. $2\ H_2S_{(g)} \rightarrow 2\ H_{2(g)} + S_{2(g)}$

When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of $H_2S_{(g)}$ 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 $S_{2(g)}$ 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 $mol{\cdot}L^{-1}$, of the following gases in the container at 483 K.

(i) $H_{2(g)}$

(ii) $H_2S_{(g)}$

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

(d) Calculate the partial pressure of $S_{2(g)}$ in the container at equilibrium at 483 K.

(e) For the reaction $H_{2(g)} + S_{2(g)} \rightarrow H_2S_{(g)}$ at 483 K, calculate the value of the equilibrium constant, K_c.

Answer:

(a) $K_c =$

(b) (i) $\times \quad = 5.95\times10^{-2}\ M\ H_2$

(ii) $$\frac{\left(3.40\text{ g }H_2S\times\frac{1\text{ mol}}{34.0\text{ g}}\right)-\left(3.72\times10^{-2}\text{ mol }S_2\times\frac{2\text{ mol }H_2S}{1\text{ mol }S_2}\right)}{1.25\text{ L}} = 2.05\times10^{-2}\ M\ H_2S$$

(c) $$K_c = \frac{\left[5.952\times10^{-2}\right]^2\left[\frac{3.72\times10^{-2}}{1.25}\right]}{[0.02048]^2} = 0.251$$

(d) PV=nRT = 1.18

(e) $K'_c = \sqrt{1/K_c} = 2.00$