

1977 D

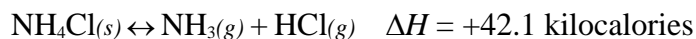
For the system $2 \text{SO}_2(g) + \text{O}_2(g) \leftrightarrow 2 \text{SO}_3(g)$, ΔH is negative for the production of SO_3 . Assume that one has an equilibrium mixture of these substances. Predict the effect of each of the following changes on the value of the equilibrium constant and on the number 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



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 $\text{NH}_3(g)$ 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 $\text{NH}_4\text{Cl}(s)$ 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_{NH_3} 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_{NH_3} increases.

1981 A

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



- (a) Some solid NH_4HS is placed in an evacuated vessel at 25°C . After equilibrium is attained, the total pressure inside the vessel is found to be 0.659 atmospheres. 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 25°C .
- (b) Some extra NH_3 gas is injected into the vessel containing the sample described in part (a). When equilibrium is reestablished at 25°C , 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 25°C . 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) $K_P = (P_{\text{NH}_3})(P_{\text{H}_2\text{S}})$
 $P_{\text{NH}_3} = P_{\text{H}_2\text{S}} = 0.659/2 \text{ atm} = 0.330 \text{ atm}$
 $K_P = (0.330)^2 = 0.109$
- (b) $P_{\text{NH}_3} = 2 P_{\text{H}_2\text{S}}$
 $(2x)(x) = 0.109 ; x = 0.233 \text{ atm} = P_{\text{H}_2\text{S}}$
 $P_{\text{NH}_3} = 0.466 \text{ atm}$
- (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 $(0.500 - 0.330) = 0.170 \text{ atm}$.

$n = \text{mol of each reactant} = \text{mol of solid product}$

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

1983 A

Sulfuryl chloride, SO_2Cl_2 , is a highly reactive gaseous compound. When heated, it decomposes as follows: $\text{SO}_2\text{Cl}_{2(g)} \rightarrow \text{SO}_{2(g)} + \text{Cl}_{2(g)}$. 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 $\text{SO}_2\text{Cl}_{2(g)}$ 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 $\text{SO}_2\text{Cl}_{2(g)}$ 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)

$$P = \frac{nRT}{V} = \frac{\frac{33.509 \text{ g}}{135.0 \text{ g}} \cdot \frac{1 \text{ mol}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}} (375 \text{ K})}{1.00 \text{ L}}$$

$$= 0.800 \text{ atm}$$

$$(b) P_{\text{SO}_2\text{Cl}_2} = (0.800 - y) \text{ atm}$$

$$P_{\text{SO}_2} = P_{\text{Cl}_2} = y \text{ atm}$$

$$P_T = P_{\text{SO}_2\text{Cl}_2} + P_{\text{SO}_2} + P_{\text{Cl}_2}$$

$$1.43 \text{ atm} = (0.800 - y + y + y) \text{ atm}$$

$$y = 0.63 \text{ atm} = P_{\text{SO}_2} = P_{\text{Cl}_2}$$

$$P_{\text{SO}_2\text{Cl}_2} = (0.800 - 0.63) \text{ atm} = 0.17 \text{ atm}$$

$$K_p = \frac{(P_{\text{SO}_2})(P_{\text{Cl}_2})}{P_{\text{SO}_2\text{Cl}_2}} = \frac{(0.63 \text{ atm})^2}{0.17 \text{ atm}} = 2.3 \text{ atm}$$

(c)

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

1985 A

At 25°C 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 25°C?

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

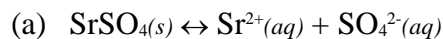
(c) An aqueous solution of $\text{Sr}(\text{NO}_3)_2$ is added slowly to 1.0 litre of a well-stirred solution containing 0.020 mole F^- and 0.10 mole SO_4^{2-} at 25°C. (You may assume that the added $\text{Sr}(\text{NO}_3)_2$ solution does not materially affect the total volume of the system.)

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 $\text{Sr}(\text{NO}_3)_2$ 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:



$$\text{At equilibrium: } [\text{Sr}^{2+}] = X \text{ M} = [\text{SO}_4^{2-}]$$

$$X^2 = K_{sp} = 7.6 \times 10^{-7}$$

$$X = 8.7 \times 10^{-4} \text{ mol/L, solubility of } \text{SrSO}_4$$



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

$$K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 = (X)(2X)^2 = 7.9 \times 10^{-10}$$

$$X = 5.8 \times 10^{-4} \text{ mol/L, solubility of } \text{SrF}_2$$

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

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

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

$$K_{sp} = [\text{Sr}^{2+}][\text{SO}_4^{2-}] = 7.6 \times 10^{-7}$$

$$= (y) \frac{0.10 \text{ mol}}{1.0 \text{ L}} = 7.6 \times 10^{-7} ; y = 7.6 \times 10^{-6} \text{ M}$$

Since $2.0 \times 10^{-6} \text{ M} < 7.6 \times 10^{-6} \text{ M}$, SrF_2 must precipitate first.

When SrF_2 precipitates, $[\text{Sr}^{2+}] = 2.0 \times 10^{-6} \text{ M}$

- (d) The second precipitate to form is SrSO_4 , which appears when $[\text{Sr}^{2+}] = 7.6 \times 10^{-6} \text{ M}$ (based on calculations in Part c.)

When $[\text{Sr}^{2+}] = 7.6 \times 10^{-6} \text{ M}$, $[\text{F}^-]$ is determined as follows:

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

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

$$\% \text{ F}^- \text{ still in solution} = \frac{1.0 \times 10^{-2}}{2.0 \times 10^{-2}} \times 100 = 50. \%$$

1988 A

At elevated temperatures, SbCl_5 gas decomposes into SbCl_3 gas and Cl_2 gas as shown by the following equation: $\text{SbCl}_5(g) \leftrightarrow \text{SbCl}_3(g) + \text{Cl}_2(g)$

- (a) An 89.7 gram sample of SbCl_5 (molecular weight 299.0) is placed in an evacuated 15.0 litre container at 182°C .
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 182°C , 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 182°C . 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:

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

- (a) (1)

$$[\text{SbCl}_5]_{\text{init}} = \frac{0.300 \text{ mol}}{15.0 \text{ L}} = 0.020 \text{ M}$$

$$(2) T = 182^\circ\text{C} + 273 = 455 \text{ K}$$

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

$$= 0.747 \text{ atm}$$

OR

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

$$1\text{ atm} \times f(22.4\text{ L/mol}, 30.0\text{ L/mol}) = 0.747\text{ atm}$$

$$(b) [\text{SbCl}_3] = [\text{Cl}_2] = (0.0200\text{ mol/L})(0.292)$$

$$= 5.84 \times 10^{-3}\text{M}$$

$$[\text{SbCl}_5] = (0.0200\text{ mol/L})(0.708) = 1.42 \times 10^{-2}\text{M}$$

$$K_c = \frac{[\text{SbCl}_3][\text{Cl}_2]}{[\text{SbCl}_5]} = \frac{(5.84 \times 10^{-3})^2}{1.42 \times 10^{-2}} = 2.41 \times 10^{-3}$$

OR

$$P_{\text{SbCl}_3} = P_{\text{Cl}_2} = (0.747\text{ atm})(0.292) = 0.218\text{ atm}$$

$$P_{\text{SbCl}_5} = (0.747\text{ atm})(0.708) = 0.529\text{ atm}$$

$$K_P = \frac{(P_{\text{SbCl}_3})(P_{\text{Cl}_2})}{P_{\text{SbCl}_5}} = \frac{(0.218)^2}{0.529} = 8.98 \times 10^{-2}$$

$$K = \frac{[\text{SbCl}_3][\text{Cl}_2]}{[\text{SbCl}_5]} = 0.117$$

(c)

$$[\text{SbCl}_5]_{\text{equil.}} = \frac{(1.00 - 0.70)\text{mol}}{2.00\text{L}} = 0.15\text{M}$$

$$[\text{SbCl}_3]_{\text{equil.}} = \frac{0.700\text{ mol}}{2.00\text{L}} = 0.350\text{M}$$

$$[\text{Cl}_2]_{\text{equil.}} = X\text{M}$$

$$K_C = \frac{(0.350)(X)}{0.15} = 0.117; X = 0.050\text{M} = [\text{Cl}_2]$$

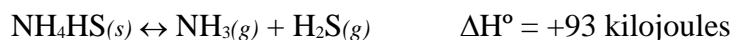
$$\text{Moles Cl}_{2(\text{equil.})} = (0.0050\text{ mol/L})(2.00\text{L})$$

$$= 0.10\text{ mol Cl}_2$$

$$\text{Moles Cl}_2 \text{ needed to make } 0.300\text{ mol SbCl}_3 \text{ into SbCl}_5 = 0.30\text{ mol}$$

$$\text{Moles Cl}_2 \text{ that must be added} = 0.40\text{ mol}$$

1988 D



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.

- The effect on the equilibrium partial pressure of NH_3 gas when additional solid NH_4HS is introduced into the container
- The effect on the equilibrium partial pressure of NH_3 gas when additional solid H_2S is introduced into the container
- The effect on the mass of solid NH_4HS present when the volume of the container is decreased
- The effect on the mass of solid NH_4HS present when the temperature is increased.

Answer:

- The equilibrium pressure of NH_3 gas would be unaffected. $K_P = (P_{\text{NH}_3})(P_{\text{H}_2\text{S}})$. 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. $K_p = (P_{\text{NH}_3})(P_{\text{H}_2\text{S}})$. (A complete explanation based on LeChatelier's 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 . $K_p = (P_{\text{NH}_3})(P_{\text{H}_2\text{S}})$. (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



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 160°C . 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 $\text{H}_2\text{O}(g)$ 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 160°C , what would the total pressure have been at equilibrium? Explain.

Answer:

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

(a)

$$\text{mol H}_2\text{O} = (1/2)(1.09 \text{ mol}) = 0.545 \text{ mol H}_2\text{O}(g)$$

(b)

$$0.545 \text{ mol H}_2\text{O} \propto \frac{2 \text{ mol NaHCO}_3}{1 \text{ mol H}_2\text{O}} \propto \frac{84.0 \text{ g NaHCO}_3}{1 \text{ mol}} =$$

$$= 91.9 \text{ g NaHCO}_3 \text{ decomposed}$$

$$\text{remaining} = 100. \text{ g} - 91.9 \text{ g} = 8.1 \text{ g}$$

OR

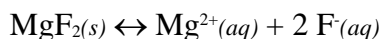
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$$\begin{array}{l} \nearrow \\ 0.545 \text{ mol H}_2\text{O} \propto \frac{18.0 \text{ g}}{1 \text{ mol}} + 0.545 \text{ mol CO}_2 \propto \frac{44.0 \text{ g}}{1 \text{ mol}} \searrow \end{array}$$

$$= 100 \text{ g} - 33.8 \text{ g} = 66.2 \text{ g} \text{ [includes Na}_2\text{CO}_3 \text{ solid in this mass]}$$

- (c) $K_p = (P_{\text{H}_2\text{O}})(P_{\text{CO}_2}) = (3.88)(3.88) \text{ atm}^2 = 15.1 \text{ atm}^2$
- (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 110 g will not affect the equilibrium.

1994 A



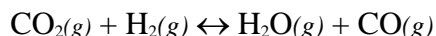
In a saturated solution of MgF_2 at 18°C , the concentration of Mg^{2+} is 1.21×10^{-3} molar. The equilibrium is represented by the equation above.

- Write the expression for the solubility-product constant, K_{sp} , and calculate its value at 18°C .
- Calculate the equilibrium concentration of Mg^{2+} in 1.000 liter of saturated MgF_2 solution at 18°C to which 0.100 mole of solid KF has been added. The KF dissolves completely. Assume the volume change is negligible.
- Predict whether a precipitate of MgF_2 will form when 100.0 milliliters of a 3.00×10^{-3} -molar $\text{Mg}(\text{NO}_3)_2$ solution is mixed with 200.0 milliliters of a 2.00×10^{-3} -molar NaF solution at 18°C . Calculations to support your prediction must be shown.
- At 27°C 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:

- $K_{sp} = [\text{Mg}^{2+}][\text{F}^{-}]^2 = (1.21 \times 10^{-3})(2.42 \times 10^{-3})^2$
 $= 7.09 \times 10^{-9}$
- X = concentration loss by Mg^{2+} ion
 $2X$ = concentration loss by F^{-} ion
 $[\text{Mg}^{2+}] = (1.21 \times 10^{-3} - X) \text{ M}$
 $[\text{F}^{-}] = (0.100 + 2.42 \times 10^{-3} - 2X) \text{ M}$
since X is a small number then $(0.100 + 2.42 \times 10^{-3} - 2X) \approx 0.100$
 $K_{sp} = 7.09 \times 10^{-9} = (1.21 \times 10^{-3} - X)(0.100)^2$
 $X = 1.2092914 \times 10^{-3}$
 $[\text{Mg}^{2+}] = 1.21 \times 10^{-3} - 1.20929 \times 10^{-3} = 7.09 \times 10^{-7} \text{ M}$
- $[\text{Mg}^{2+}] = 3.00 \times 10^{-3} \text{ M} \times 100.0 \text{ mL} / 300.0 \text{ mL} = 1.00 \times 10^{-3} \text{ M}$
 $[\text{F}^{-}] = 2.00 \times 10^{-3} \text{ M} \times 200.0 \text{ mL} / 300.0 \text{ mL} = 1.33 \times 10^{-3} \text{ M}$
trial $K_{sp} = (1.00 \times 10^{-3})(1.33 \times 10^{-3})^2 = 1.78 \times 10^{-9}$
trial $K_{sp} < 7.09 \times 10^{-9}$, \therefore no ppt.
- @ 18°C , $1.21 \times 10^{-3} \text{ M}$ MgF_2 dissolves
@ 27°C , $1.17 \times 10^{-3} \text{ M}$ MgF_2 dissolves
 $\text{MgF}_2 \leftrightarrow \text{Mg}^{2+} + 2 \text{F}^{-} + \text{heat}$
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



When $\text{H}_{2(g)}$ is mixed with $\text{CO}_{2(g)}$ at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

$$\begin{aligned} [\text{H}_2] &= 0.20 \text{ mol/L} \\ [\text{CO}_2] &= 0.30 \text{ mol/L} \end{aligned}$$

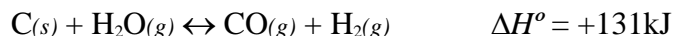
$$[\text{H}_2\text{O}] = [\text{CO}] = 0.55 \text{ mol/L}$$

- What is the mole fraction of $\text{CO}(g)$ in the equilibrium mixture?
- Using the equilibrium concentrations given above, calculate the value of K_c , the equilibrium constant for the reaction.
- Determine K_p in terms of K_c for this system.
- When the system is cooled from 2,000 K to a lower temperature, 30.0 percent of the $\text{CO}(g)$ is converted back to $\text{CO}_2(g)$. Calculate the value of K_c at this lower temperature.
- In a different experiment, 0.50 mole of $\text{H}_2(g)$ is mixed with 0.50 mole of $\text{CO}_2(g)$ in a 3.0-liter reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of $\text{CO}(g)$ at this temperature.

Answer:

- $\text{CO} = f(0.55 \text{ mol}, 1.6 \text{ mol}) = 0.34$
- $K_c = ([\text{H}_2\text{O}][\text{CO}])/([\text{H}_2][\text{CO}_2]) = (0.55 \times 0.55)/(0.20 \times 0.30) = 5.04$
- since $\Delta n = 0$, $K_c = K_p$
- $[\text{CO}] = 0.55 - 30.0\% = 0.55 - 0.165 = 0.385 \text{ M}$
 $[\text{H}_2\text{O}] = 0.55 - 0.165 = 0.385 \text{ M}$
 $[\text{H}_2] = 0.20 + 0.165 = 0.365 \text{ M}$
 $[\text{CO}_2] = 0.30 + 0.165 = 0.465 \text{ M}$
 $K = (0.385)^2/(0.365 \times 0.465) = 0.87$
- let $X = \Delta[\text{H}_2]$ to reach equilibrium
 $[\text{H}_2] = 0.50 \text{ mol}/3.0\text{L} - X = 0.167 - X$
 $[\text{CO}_2] = 0.50 \text{ mol}/3.0\text{L} - X = 0.167 - X$
 $[\text{CO}] = +X$; $[\text{H}_2\text{O}] = +X$
 $K = X^2/(0.167 - X)^2 = 5.04$; $X = [\text{CO}] = 0.12 \text{ M}$

1998 D



A rigid container holds a mixture of graphite pellets ($\text{C}(s)$), $\text{H}_2\text{O}(g)$, $\text{CO}(g)$, and $\text{H}_2(g)$ at equilibrium. State whether the number of moles of $\text{CO}(g)$ 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.

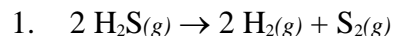
- Additional $\text{H}_2(g)$ is added to the equilibrium mixture at constant volume.
- The temperature of the equilibrium mixture is increased at constant volume.
- The volume of the container is decreased at constant temperature.
- The graphite pellets are pulverized.

Answer

- 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.
- 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, \therefore , 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



When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of $\text{H}_2\text{S}(g)$ 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 $\text{S}_2(g)$ 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 $\text{mol} \cdot \text{L}^{-1}$, of the following gases in the container at 483 K.
- $\text{H}_2(g)$
 - $\text{H}_2\text{S}(g)$
- (c) Calculate the value of the equilibrium constant, K_c , for the decomposition reaction at 483 K.
- (d) Calculate the partial pressure of $\text{S}_2(g)$ in the container at equilibrium at 483 K.
- (e) For the reaction $\text{H}_2(g) + \text{S}_2(g) \rightarrow \text{H}_2\text{S}(g)$ at 483 K, calculate the value of the equilibrium constant, K_c .

Answer :

(a) $K_c =$

(b) (i) $\times = 5.95 \times 10^{-2} M \text{H}_2$

$$(ii) \frac{\left(3.40 \text{ g H}_2\text{S} \times \frac{1 \text{ mol}}{34.0 \text{ g}} \right) - \left(3.72 \times 10^{-2} \text{ mol S}_2 \times \frac{2 \text{ mol H}_2\text{S}}{1 \text{ mol S}_2} \right)}{1.25 \text{ L}} = 2.05 \times 10^{-2} M \text{H}_2\text{S}$$

$$\frac{[5.952 \times 10^{-2}]^2 \left[\frac{3.72 \times 10^{-2}}{1.25} \right]}{[0.02048]^2}$$

(c) $K_c = \frac{[0.02048]^2}{[0.02048]^2} = 0.251$

(d) $PV=nRT = 1.18$

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