

GAS PHASE EQUILIBRIUM PRACTICE PROBLEMS

- (1) Sulfuryl chloride, SO_2Cl_2 , is a highly reactive gaseous compound. When heated, it decomposes as follows:



This decomposition is endothermic. A sample of 3.509 grams of SO_2Cl_2 is placed in an evacuated 1.00 L bulb and the temperature is raised to 375 K.

- a. What would be the pressure in atmospheres in the bulb if no dissociation of the SO_2Cl_2 occurred?

 - b. When the bulb has come to equilibrium at 375 K, 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 375 K.

 - c. Give the expression for the equilibrium constant, K_c , for the decomposition of SO_2Cl_2 . Calculate the value of the equilibrium constant at 375 K.

 - d. If the temperature were raised to 500 K, what effect would this have on the equilibrium constant? Explain your answer.
- (2) At elevated temperatures, SbCl_5 gas decomposes into SbCl_3 gas and Cl_2 gas, as shown by the following equation:

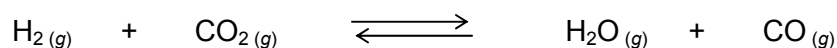


An 89.7 gram sample of SbCl_5 (molar mass 299.0) is placed in an evacuated 15.0 L container at 182 °C.

- a. What is the concentration in moles per liter of SbCl_5 in the container before any decomposition occurs?

- b. What is the pressure in atmospheres of SbCl_5 in the container before decomposition occurs?
- c. If the SbCl_5 is 29.2 percent decomposed when equilibrium is established at 182°C , calculate the value for both equilibrium constants, K_p and K_c , for the decomposition reaction.
- d. In order to produce some SbCl_5 , a 1.00 mole sample of SbCl_3 is first placed in an empty 2.00 liter container maintained at a temperature different for 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 moles at equilibrium?

(3) Consider the following equation:



when $\text{H}_2(g)$ is mixed with $\text{CO}_2(g)$ at 2000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

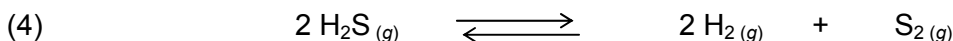
$$[\text{H}_2] = 0.20 \text{ mol/L}$$

$$[\text{CO}_2] = 0.30 \text{ mol/L}$$

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

- a. What is the mole fraction of $\text{CO}(g)$ in the equilibrium mixture?
- b. Using the information above, write the equilibrium expression and calculate the value of K_c , the equilibrium constant for the reaction.
- c. Calculate the value of K_p for this system.

- d. When the system is cooled from 2000 K to a lower temperature, 30.0 % of the $\text{CO}_{(g)}$ is converted back to $\text{CO}_{2(g)}$. Calculate the value of K_C at this lower temperature.
- e. In a different experiment, 0.50 moles of $\text{H}_2(g)$ is mixed with 0.50 moles of $\text{CO}_2(g)$ in a 3.0 L reaction vessel at 2000 K. Calculate the equilibrium concentration, in moles per liter, of $\text{CO}_{(g)}$ at this temperature.



When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 gram sample of $\text{H}_2\text{S}_{(g)}$ is introduced into an evacuated 1.25 L container. The sealed container is heated to 483 K, and 3.72×10^{-2} moles of $\text{S}_{2(g)}$ are present at equilibrium.

- a. Write the expression for the equilibrium constant, K_C , for the decomposition reaction represented above.
- b. Calculate the equilibrium concentrations, in mol/L, of $\text{H}_2(g)$ and $\text{H}_2\text{S}_{(g)}$.
- 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 483 K. Express your answer in atmospheres.
- e. For the reaction:
$$\text{H}_{2(g)} + \frac{1}{2} \text{S}_{2(g)} \rightleftharpoons \text{H}_2\text{S}_{(g)}$$
 at 483 K, calculate the value of the equilibrium constant, K_C .

- (5) Solid sodium hydrogen carbonate (sodium bicarbonate), NaHCO_3 , decomposes on heating according to the equation below:



- a. A sample of 100. grams of solid NaHCO_3 was placed in a previously evacuated rigid 5.00 L 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 part (a).
- c. Write the equilibrium expression for the equilibrium constant, K_p , and calculate its value for the reaction under the conditions in part (a).
- d. If 110. grams of solid NaHCO_3 had been placed in the 5.00 L container and heated to 160°C , what would the total pressure have been at equilibrium?

KEY: GAS PHASE EQUILIBRIUM PRACTICE PROBLEMS

- (1)
- 0.800 atmospheres
 - SO_2Cl_2 : 0.170 atmospheres
 SO_2 : 0.630 atmospheres
 Cl_2 : 0.630 atmospheres
 - $K_C = \frac{[\text{SO}_2][\text{Cl}_2]}{[\text{SO}_2\text{Cl}_2]}$, $K_C = 7.58 \times 10^{-2}$
 - Since the reaction is endothermic, raising the temperature to 500 K will cause the equilibrium to shift toward the product side. This will increase the concentration of the products and decrease the concentration of the reactant, thus increasing the equilibrium constant.
- (2)
- 0.0200 M
 - 0.747 atmospheres
 - $K_P = 8.98 \times 10^{-2}$
 $K_C = 2.40 \times 10^{-3}$
 - 0.100 moles Cl_2
- (3)
- $X_{\text{CO}} = 0.34$
 - $K_C = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]}$, $K_C = 5.04$
 - $K_P = 5.04$
 - $K_C = 0.873$
 - $[\text{CO}]_{\text{eq}} = 0.115$
- (4)
- $K_C = \frac{[\text{H}_2]^2 [\text{S}_2]}{[\text{H}_2\text{S}]^2}$
 - $[\text{H}_2] = 5.95 \times 10^{-2}$
 $[\text{H}_2\text{S}] = 2.05 \times 10^{-2}$
 - $K_C = 0.251$
 - $P = 1.18$ atmospheres
 - $K_C = 0.631$
- (5)
- 0.546 moles
 - 8.27 grams NaHCO_3 remaining
 - $K_P = (P_{\text{H}_2\text{O}}) (P_{\text{CO}_2})$
 $K_P = 15.1$
 - The total pressure would remain at 7.76 atm. Since some solid remained when 100. g was used (and there had been no temperature change), then using 110 g would not affect the equilibrium.