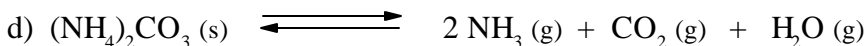
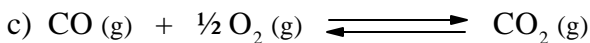
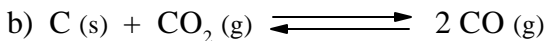
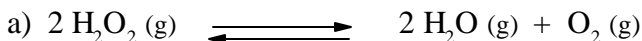


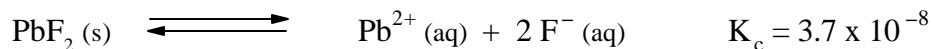
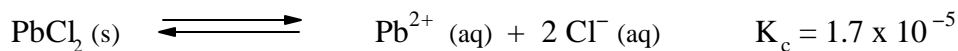
1. Write equilibrium constant expressions for the following reactions. For gases, use either pressures or concentrations.



2. Three moles of pure  $\text{SO}_3$  are placed in a 800 L flask at 1150 K. At equilibrium, 0.58 mol of  $\text{O}_2$  has formed. Calculate  $K_c$  for the reaction at 1150 K.

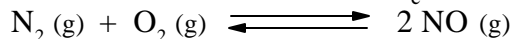


3. Neither  $\text{PbCl}_2$  nor  $\text{PbF}_2$  is appreciably soluble in water. If solid  $\text{PbCl}_2$  and solid  $\text{PbF}_2$  are placed in equal amounts of water in separate beakers, in which beaker is the concentration of  $\text{Pb}^{2+}$  greater? Equilibrium constants for these solids dissolving in water are



4. How does the reaction quotient for a reaction differ from the equilibrium constant for that reaction?

5. At 2300 K the equilibrium constant for the formation of NO (g) is  $K_c = 1.7 \times 10^{-3}$ .

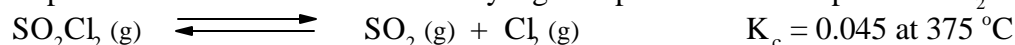


a) If analysis show that the concentration of  $\text{N}_2$  and  $\text{O}_2$  are 0.25 M, and that of NO is 0.0042 M, is the system at equilibrium?

b) If the system is not at equilibrium, in which direction does the reaction proceed?

c) When the system is at equilibrium, what are the equilibrium concentrations?

6. Sulfuryl chloride,  $\text{SO}_2\text{Cl}_2$ , is a compound with very irritating vapors; it is used as a reagent in the synthesis of organic compounds. When heated to a sufficiently high temperature it decomposes to  $\text{SO}_2$  and  $\text{Cl}_2$ .

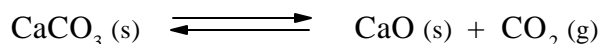


a) In a 1.00 L flask you place 6.70 g of  $\text{SO}_2\text{Cl}_2$  and heat it to  $375^\circ\text{C}$ . What is the concentration of each of the compounds in the system when equilibrium is achieved?

b) What are the concentrations of  $\text{SO}_2\text{Cl}_2$ ,  $\text{SO}_2$ , and  $\text{Cl}_2$  at equilibrium in the 1.00 L flask at  $375^\circ\text{C}$  if you begin with a mixture of  $\text{SO}_2\text{Cl}_2$  (6.70 g) and  $\text{Cl}_2$  (1 atm)?

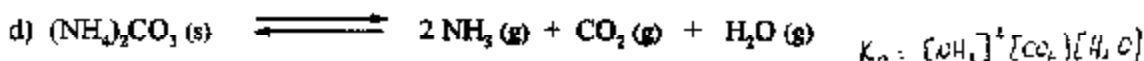
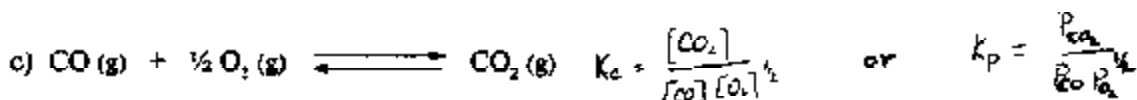
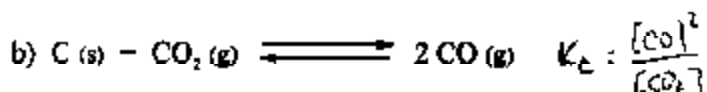
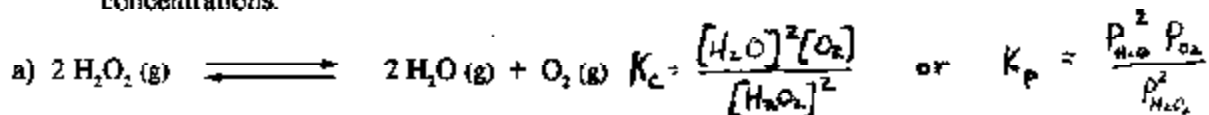
c) Do the fractions of  $\text{SO}_2\text{Cl}_2$  in parts (a) and (b) agree with your expectation based on Le Chatelier's principle?

7. The decomposition of calcium carbonate

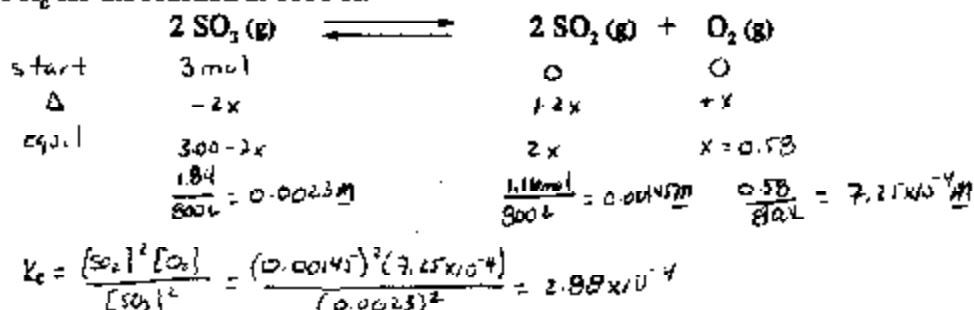


is an endothermic process. Using Le Chatelier's principle, explain how increasing the temperature affects the equilibrium. If more  $\text{CaCO}_3$  is added to the flask in which this equilibrium exists, how is the equilibrium affected? What if some additional  $\text{CO}_2$  is placed in the flask?

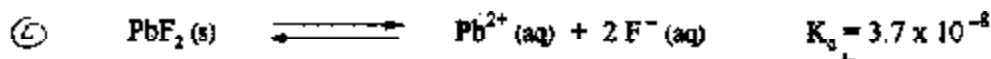
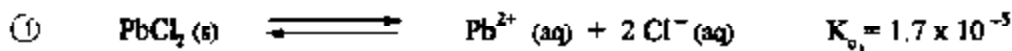
1. Write equilibrium constant expressions for the following reactions. For gases, use either pressures or concentrations.



2. Three moles of pure  $\text{SO}_3$  are placed in a 800 L flask at 1150 K. At equilibrium, 0.58 mol of  $\text{O}_2$  has formed. Calculate  $K_c$  for the reaction at 1150 K.



3. Neither  $\text{PbCl}_2$  nor  $\text{PbF}_2$  is appreciably soluble in water. If solid  $\text{PbCl}_2$  and solid  $\text{PbF}_2$  are placed in equal amounts of water in separate beakers, in which beaker is the concentration of  $\text{Pb}^{2+}$  greater? Equilibrium constants for these solids dissolving in water are



$K_{s1}$  for rxn (1) is larger than  $K_{s2}$  for rxn (2), which means rxn (1) proceeded more towards the products than rxn (2). Since more products is produced in rxn (1), the conc of  $\text{Pb}^{2+}$  is greater in rxn (1) than rxn (2).

4. How does the reaction quotient for a reaction differ from the equilibrium constant for that reaction?

The rxn quotient is determined for a rxn which may not be at equilibrium, whereas  $K_c$  is determined for a rxn at equilibrium.

5. At 2300 K the equilibrium constant for the formation of NO (g) is  $K_c = 1.7 \times 10^{-3}$ .



- a) If analysis show that the concentration of  $\text{N}_2$  and  $\text{O}_2$  are 0.25 M, and that of NO is 0.0042 M, is the system at equilibrium?

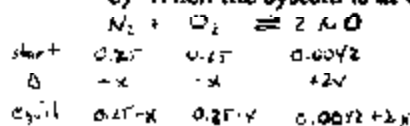
$$Q_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} = \frac{(0.0042)^2}{(0.25)^2} = 2.8 \times 10^{-4} < 1.7 \times 10^{-3}$$

system is not at equilibrium

- b) If the system is not at equilibrium, in which direction does the reaction proceed?

To obtain equilibrium, the rxn will proceed towards the products

- c) When the system is at equilibrium, what are the equilibrium concentrations?



$$K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$$

$$1.7 \times 10^{-3} = \frac{(0.0042+2x)^2}{(0.25-x)^2}$$

$$0.0412 = \frac{0.01716x}{0.25+x}$$

$$0.0103 - 0.0112x = 0.0042 + 2x$$

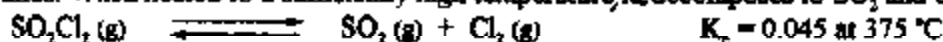
$$0.00610 = 2.0112x$$

$$0.003 = x$$

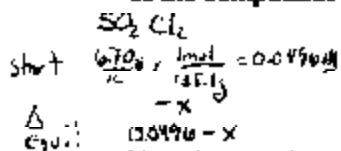
$$[\text{N}_2] = [\text{O}_2] = 0.2471$$

$$[\text{NO}] = 0.0102 \text{ M}$$

6. Sulfuryl chloride,  $\text{SO}_2\text{Cl}_2$ , is a compound with very irritating vapors; it is used as a reagent in the synthesis of organic compounds. When heated to a sufficiently high temperature, it decomposes to  $\text{SO}_2$  and  $\text{Cl}_2$ .



- a) In a 1.00 L flask you place 6.70 g of  $\text{SO}_2\text{Cl}_2$  and heat it to  $375^\circ\text{C}$ . What is the concentration of each of the compounds in the system when equilibrium is achieved?



$$K_c = \frac{[\text{SO}_2][\text{Cl}_2]}{[\text{SO}_2\text{Cl}_2]}$$

$$0.045 = \frac{x^2}{0.0496-x}$$

$$x^2 + 0.045x - 0.00223 = 0$$

$$x = \frac{-0.045 \pm \sqrt{(0.045)^2 - 4(1)(-0.00223)}}{2}$$

$$x = \frac{-0.045 \pm 0.0523}{2}$$

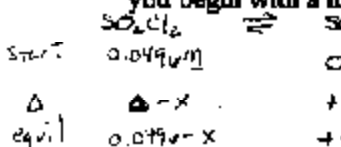
$$x = 0.0037 \text{ or } x = -0.0487$$

$$[\text{SO}_2] = 0.0298 \text{ M}$$

$$[\text{Cl}_2] = 0.0298 \text{ M}$$

$$[\text{SO}_2\text{Cl}_2] = 0.0198 \text{ M}$$

- b) What are the concentrations of  $\text{SO}_2\text{Cl}_2$ ,  $\text{SO}_2$ , and  $\text{Cl}_2$  at equilibrium in the 1.00 L flask at  $375^\circ\text{C}$  if you begin with a mixture of  $\text{SO}_2\text{Cl}_2$  (6.70 g) and  $\text{Cl}_2$  (1 atm)?



$$K_c = \frac{[\text{SO}_2][\text{Cl}_2]}{[\text{SO}_2\text{Cl}_2]}$$

$$0.045 = \frac{(x)(0.0188+x)}{0.0496-x}$$

$$0.00223 - 0.045x = 0.0188x + x^2$$

$$0 = x^2 + 0.0638x - 0.00223$$

$$x = \frac{-0.0638 \pm \sqrt{(0.0638)^2 - 4(1)(-0.00223)}}{2}$$

$$x = 0.025 \text{ M}$$

$$[\text{SO}_2] = 0.045 \text{ M}$$

$$[\text{Cl}_2] = 0.0437 \text{ M}$$

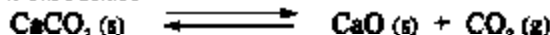
$$[\text{SO}_2\text{Cl}_2] = 0.0245 \text{ M}$$

- c) Do the fractions of  $\text{SO}_2\text{Cl}_2$  in parts (a) and (b) agree with your expectation based on Le Chatelier's principle? Adding the product,  $\text{Cl}_2$  to the rxn, will shift the rxn towards the reactants.

$\therefore$  expect  $[\text{SO}_2\text{Cl}_2] \uparrow$  and  $[\text{SO}_2] \downarrow$

Comparing (a) & b)  $[\text{SO}_2\text{Cl}_2]$  increases from 0.0198 M to 0.0245 M  
 $\rightarrow$   $[\text{SO}_2]$  decreases from 0.0298 M to 0.025 M

7. The decomposition of calcium carbonate



is an endothermic process. Using Le Chatelier's principle, explain how increasing the temperature affects the equilibrium. If more  $\text{CaCO}_3$  is added to the flask in which this equilibrium exists, how is the equilibrium affected? What if some additional  $\text{CO}_2$  is placed in the flask?

- ①  $\uparrow$  temp shifts the rxn towards the product
- ② Adding  $\text{CaCO}_3$  will shift the rxn towards the product
- ③ Adding  $\text{CO}_2$  will shift the rxn towards the reactants