1. Consider the following reaction:

\[ \text{N}_2 (g) + 3 \text{H}_2 (g) \rightleftharpoons 2 \text{NH}_3 (g) \]

a. Write the equilibrium expression \((K_p)\).

b. What are the values of \(K_p\) and \(K\) at 25°C if the equilibrium pressures for \(\text{N}_2\), \(\text{H}_2\) and \(\text{NH}_3\) were 0.5 atm, 0.1 atm and 0.17 atm respectively.

c. Does the reaction favor reactants or products at 25°C?

2. Determine if \(K = K_p\), \(K > K_p\) or \(K < K_p\) for each of the following reactions at 300K.

a. \(\text{P}_4 (s) + 6 \text{Cl}_2 (g) \rightleftharpoons 4 \text{PCl}_3 (g)\)

b. \(\text{H}_2 (g) + \text{Cl}_2 (g) \rightleftharpoons 2 \text{HCl (g)}\)

3. Given the following information:

\[ \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI(g)} \quad K = 54.0 \]

\[ \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) \quad K = 96.0 \]

Determine the value of the equilibrium constant \((K)\) for the following reaction:

\[ 2 \text{NH}_3(g) + 3 \text{I}_2(g) \rightleftharpoons 6 \text{HI(g)} + \text{N}_2(g) \]

4. Consider the following reaction in a 3.5 L flask at constant temperature:

\[ \text{CO}_2 (g) + \text{H}_2 (g) \rightleftharpoons \text{CO (g)} + \text{H}_2\text{O (g)} \quad K_p = 11 \]

What will happen to the partial pressure of hydrogen if initially there is 1.4 atm of carbon dioxide, 2.5 atm of hydrogen, 0.75 atm of carbon monoxide and 5.3 atm of water?

5. Consider the following reaction:

\[ \text{N}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2 \text{NO (g)} \]

Initially the partial pressures of \(\text{N}_2\) and \(\text{O}_2\) are 1 atm and 3 atm respectively. At equilibrium the pressure of NO is measured to be 1.5 atm. What is the value of \(K_p\) for this reaction?

6. Pure \(\text{PCl}_5\) is introduced into an empty rigid 5-L flask at 90 °C with a pressure of 0.45 atm. The \(\text{PCl}_5\) decomposes into solid P and gaseous \(\text{Cl}_2\) according to the following reaction.

\[ 2 \text{PCl}_5(g) \rightleftharpoons 2 \text{P (s)} + 5 \text{Cl}_2(g) \]

At equilibrium the total pressure is measured to be 0.93 atm. Determine the equilibrium constant \((K_p)\) at 90°C.

7. Consider the following reaction:

\[ 2 \text{NOBr (g)} \rightleftharpoons 2 \text{NO (g)} + \text{Br}_2 (g) \quad K_p = 34 \]

An unknown pressure of NOBr is put into a rigid container at 40 °C, equilibrium is reached when 86 % of the original partial pressure of NOBr has decomposed. Determine the initial pressure of the NOBr.

8. At a particular temperature \(K_p = 5.7 \times 10^{-8}\) for the following reaction: \(2 \text{C (s)} + 3 \text{H}_2 (g) \rightleftharpoons \text{C}_2\text{H}_6 (g)\)

Determine the equilibrium pressure of \(\text{C}_2\text{H}_6\) if initially there are 5 g of C and 3 atm of \(\text{H}_2\)?

9. \(K_p\) is 10.5 at 250.°C for the following reaction:

\[ \text{H}_2 (g) + \text{Cl}_2 (g) \rightleftharpoons 2 \text{HCl(g)} \]

A 4.0 L reaction vessel was charged with 0.2 atm of hydrogen, 0.2 atm of chlorine and 0.8 atm of hydrogen chloride at 250.°C. What is the total pressure in the vessel at equilibrium?

10. Which of the following statements concerning equilibrium is not true?

a. The value of the equilibrium constant for a given reaction mixture is the same regardless of the direction from which equilibrium is attained.

b. The equilibrium constant changes with temperature.

c. A system that is disturbed from an equilibrium condition responds in such a way as to restore equilibrium.

d. The value of the equilibrium constant changes if one changes the concentration of the products or reactants.

e. Equilibrium in molecular systems is dynamic, with two opposing processes balancing one another.
11. Consider the following reaction at equilibrium:

\[ \text{CCl}_4 (g) \rightleftharpoons C (s) + 2 \text{Cl}_2 (g) \]

a. What happens the partial pressure of \( \text{Cl}_2 \) if the partial pressure of \( \text{CCl}_4 \) is increased?
b. What happens to the partial pressure of \( \text{CCl}_4 \) if you add more \( C \)?
c. What happens to the mass of \( C \) if you remove \( \text{Cl}_2 \)?
d. What happens the partial pressure of \( \text{Cl}_2 \) if you compress the system?
e. What happens to the partial pressure of \( \text{CCl}_4 \) if you add a catalyst?

12. The following data was collected for the reaction: \( \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) \):

\[
\begin{align*}
K_p &= 4.34 \times 10^{-3} \text{ at } 300 ^\circ \text{C} \\
K_p &= 4.51 \times 10^{-4} \text{ at } 450 ^\circ \text{C} \\
K_p &= 2.24 \times 10^{-6} \text{ at } 600 ^\circ \text{C}
\end{align*}
\]

a. Is the reaction endothermic or exothermic?
b. If the reaction is at equilibrium; which way will the reaction shift if the temperature is increased?

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**Equilibrium Expression**

\[ jA + kB \rightleftharpoons lC + mD \]

\[ K = \frac{[C]^l[D]^m}{[A]^j[B]^k} \]

\[ K_p = \frac{P_C^lP_D^m}{P_A^jP_B^k} \]

\[ K_p = K(RT)^{\Delta n} \] (\( \Delta n \) = moles of gas products – moles of gas reactant)

Pure solids and liquids are given the value of 1 in the expression

Where does the equilibrium lie?

\[ K > 1 \Rightarrow \text{Products are favored} \]

\[ K < 1 \Rightarrow \text{Reactants are favored} \]

Since reversible reactions can be written in either direction or balanced with any multiple of coefficients you must consider how this affects the equilibrium constant (\( K \) or \( K_p \))

1. If you flip a reaction take the reciprocal of the constant
   
   \( \text{ex: if } K=5 \text{ for } 2A \rightleftharpoons B \text{ then } K=\frac{1}{5} \text{ for } B \rightleftharpoons 2A \)

2. If you multiply the reaction coefficients raise the constant to the same multiple
   
   \( \text{ex: if } K=5 \text{ for } 2A \rightleftharpoons B \text{ then } K=5^{15} \text{ for } A \rightleftharpoons \frac{1}{2}B \)

3. If you add two or more reactions take the product of the constants
   
   \( \text{ex: if } K=5 \text{ for } 2A \rightleftharpoons B \) and if \( K=8 \text{ for } B \rightleftharpoons 3C \) then
   \[ K = (5)(8) = 40 \text{ for } 2A \rightleftharpoons 3C \]

To determine the direction a reaction must proceed to get to equilibrium

\[ K > Q \text{ shifts to the right (causes } Q \text{ to increase until } K=Q) \]

\[ K < Q \text{ shifts to the left (causes } Q \text{ to decrease until } K=Q) \]

\[ K = Q \text{ reaction is at equilibrium so no shift} \]