## Chapter 16 Chemical Equilibrium

## Solutions to Practice Problems

**1.**
**Problem**
Write the equilibrium expression for the reaction at 200 °C between ethanol and ethanoic acid to form ethyl ethanoate and water:
$CH_3CH_2OH(g) + CH_3COOH(g) \leftrightarrow CH_3COOCH_2CH_3 + H_2O(g)$

**What is Required?**
You must write the equilibrium expression for the reaction between ethanol and ethanoic acid to form ethyl ethanoate and water.

**What is Given?**
The chemical equation is given.

**Plan Your Strategy**
The expression for $K_c$ is a fraction. The concentrations of the products are in the numerator and the concentrations of the reactants are in the denominator. Each concentration term must be raised to the power of its coefficient in the balanced equation.

**Act on Your Strategy**
$$K_c = \frac{[CH_3COOCH_2CH_3][H_2O]}{[CH_3CH_2OH][CH_3COOH]}$$

**Check Your Solution**
The expression is a fraction. The concentrations of the products are in the numerator and the concentrations of the reactants are in the denominator. There are no coefficients.

**2.**
**Problem**
Write the equilibrium expression for the reaction between nitrogen gas and oxygen gas at high temperatures:
$N_2(g) + O_2(g) \leftrightarrow 2NO(g)$

**What is Required?**
You must write the equilibrium expression for the reaction between nitrogen gas and oxygen gas.

**What is Given?**
The chemical equation is given.

**Plan Your Strategy**
The expression for $K_c$ is a fraction. The concentrations of the products are in the numerator and the concentrations of the reactants are in the denominator. Each concentration term must be raised to the power of its coefficient in the balanced equation.

**Act on Your Strategy**

$$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

**Check Your Solution**

The expression is a fraction. The concentration of the product is in the numerator and the concentrations of the reactants are in the denominator. The concentration term for the product is raised to the power of its coefficient in the balanced equation.

**3.**

**Problem**

Write the equilibrium expression for the reaction between hydrogen gas and oxygen gas to form water vapour:

$2H_2(g) + O_2(g) \leftrightarrow 2H_2O(g)$

**What is Required?**

You must write the equilibrium expression for the reaction between hydrogen gas and oxygen gas.

**What is Given?**

The chemical equation is given.

**Plan Your Strategy**

The expression for $K_c$ is a fraction. The concentration of the product is in the numerator and the concentrations of the reactants are in the denominator. Each concentration term must be raised to the power of its coefficient in the balanced equation.

**Act on Your Strategy**

$$K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]}$$

**Check Your Solution**

The expression is a fraction. The concentration of the product is in the numerator and the concentrations of the reactants are in the denominator. Each concentration term is raised to the power of its coefficient in the balanced equation.

**4.**

**Problem**

Write the reduction-oxidation equilibrium expression for ferric and iodine ions in aqueous solution:

$2Fe^{3+}(aq) + 2I^{-}(aq) \leftrightarrow 2Fe^{2+}(aq) + I_2(aq)$

**What is Required?**

You must write the equilibrium expression for the reaction for ferric and iodine ions in aqueous solution.

**What is Given?**
The chemical equation is given.

**Plan Your Strategy**
The expression for $K_c$ is a fraction. The concentrations of the products are in the numerator and the concentrations of the reactants are in the denominator. Each concentration term must be raised to the power of its coefficient in the balanced equation.

**Act on Your Strategy**

$$K_c = \frac{[Fe^{2+}]^2[I_2]}{[Fe^{3+}]^2[I^-]^2}$$

**Check Your Solution**
The expression is a fraction. The concentration of the product is in the numerator and the concentrations of the reactants are in the denominator. Each concentration term is raised to the power of its coefficient in the balanced equation.

**5.**
**Problem**
Write the equilibrium expression for the oxidation of ammonia (one of the reactions in the manufacture of nitric acid):
$4NH_3(g) + 5O_2(g) \leftrightarrow 4NO(g) + 6H_2O(g)$

**What is Required?**
You must write the equilibrium expression for the oxidation of ammonia.

**What is Given?**
The chemical equation is given.

**Plan Your Strategy**
The expression for $K_c$ is a fraction. The concentrations of the products are in the numerator and the concentrations of the reactants are in the denominator. Each concentration term must be raised to the power of its coefficient in the balanced equation.

**Act on Your Strategy**

$$K_c = \frac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$$

**Check Your Solution**
The expression is a fraction. The concentrations of the products are in the numerator and the concentrations of the reactants are in the denominator. Each concentration term is raised to the power of its coefficient in the balanced equation.

**6.**
**Problem**
The following reaction took place in a sealed flask at 250 °C:

$$PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$$

At equilibrium, the gases in the flask had the following concentrations:
$[PCl_5(g)] = 1.2 \times 10^{-2}$ mol/L,
$[PCl_3(g)] = 1.5 \times 10^{-2}$ mol/L, and
$[Cl_2(g)] = 1.5 \times 10^{-2}$ mol/L.
Calculate the value of the equilibrium constant at 250 °C.

**What is Required?**
You must calculate the value of $K_c$.

**What is Given?**
The equation for the equilibrium system is: $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$
Equilibrium concentrations:
$[PCl_5(g)] = 1.2 \times 10^{-2}$ mol/L
$[PCl_3(g)] = 1.5 \times 10^{-2}$ mol/L
$[Cl_2(g)] = 1.5 \times 10^{-2}$ mol/L.

**Plan Your Strategy**
Write the equilibrium expression and substitute the equilibrium molar concentrations into the expression. Calculate the value of $K_c$.

**Act on Your Strategy**
The equilibrium expression is $Kc = \dfrac{[PCl_3(g)][Cl_2(g)]}{[PCl_5(g)]} = \dfrac{(1.50 \times 10^{-2})^2}{(1.2 \times 10^{-2})} = 1.9 \times 10^{-2}$

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number of significant digits (2).

**7.**
**Problem**
Iodine and bromine react to form iodine monobromide, IBr(g):

$$I_2(g) + Br_2(g) \leftrightarrow 2IBr(g)$$

At 150 °C, an equilibrium mixture in a 2.0 L flask was found to contain 0.024 mol iodine, 0.050 mol bromine, and 0.38 mol iodine monobromide. What is the magnitude of $K_c$ for the reaction at 150 °C?

**What is Required?**
You must calculate the value of $K_c$.

**What is Given?**

The equation for the equilibrium system is: $I_2(g) + Br_2(g) \leftrightarrow 2IBr(g)$
Volume of the container = 2.0 L
Amount of iodine, $I_2(g)$, = 0.024 mol
Amount of bromine, $Br_2(g)$ = 0.050 mol
Amount of iodine monobromide, IBr(g) = 0.38 mol

**Plan Your Strategy**
Calculate the molar concentration of each compound at equilibrium.
Write the equilibrium expression and substitute the equilibrium molar concentrations into this expression. Calculate the value of $K_c$.

**Act on Your Strategy**
The molar concentrations at equilibrium are:

$$[I_2(g)] = \frac{0.024\,\text{mol}}{2.0\,\text{L}} = 0.012\ \text{mol/L}$$

$$[Br_2(g)] = \frac{0.050\,\text{mol}}{2.0\,\text{L}} = 0.025\ \text{mol/L}$$

$$[IBr(g)] = \frac{0.38\,\text{mol}}{2.0\,\text{L}} = 0.19\ \text{mol/L}$$

The equilibrium expression for this reaction is: $K_c = \dfrac{[IBr(g)]^2}{[I_2(g)][Br_2(g)]}$

Substitute the equilibrium molar concentrations into this expression.

$$K_c = \frac{[IBr(g)]^2}{[I_2(g)][Br_2(g)]} = \frac{(0.19)^2}{(0.012)(0.025)} = 1.2 \times 10^2$$

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number of significant digits (2).

**8.**
**Problem**
At high temperatures, carbon dioxide gas decomposes into carbon monoxide and oxygen gas. The concentration of gases at equilibrium were measured and found to be $[CO_2(g)] = 1.2$ mol/L, $[CO(g)] = 0.35$ mol/L, and $[O_2(g)] = 0.15$ mol/L. Determine $K_c$ at the temperature of the reaction.

**What is Required?**
You must calculate the value of $K_c$.

**What is Given?**
Equilibrium concentrations:
$[CO_2(g)] = 1.2$ mol/L
$[CO(g)] = 0.35$ mol/L
$[O_2(g)] = 0.15$ mol/L

**Plan Your Strategy**
Write the chemical equation for the equilibrium system.
Write the equilibrium expression and substitute the equilibrium molar concentrations into the expression for $K_c$. Calculate the value of $K_c$.

**Act on Your Strategy**
The chemical equation for this equilibrium is: $2CO_2(g) \leftrightarrow 2CO(g) + O_2(g)$

The equilibrium expression for this system is $K_c = \dfrac{[CO(g)]^2[O_2(g)]}{[CO_2(g)]^2}$

$$K_c = \frac{[CO(g)]^2[O_2(g)]}{[CO_2(g)]^2} = \frac{(0.35)^2(0.15)}{(1.2)^2} = 0.013$$

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number of significant digits (2).

**9.**
**Problem**
Hydrogen sulfide is a pungent, poisonous gas. At 1400 K, an equilibrium mixture was found to contain 0.013 mol/L hydrogen, 0.18 mol/L hydrogen sulfide, and an undetermined amount of sulfur in the form of $S_2(g)$. The reaction is

$$2H_2S(g) \leftrightarrow 2H_2(g) + S_2(g)$$

If the value of $K_c$ at 1400 K is $2.4 \times 10^{-4}$, what concentration of $S_2(g)$ is present at equilibrium?

**What is Required?**
You must determine the equilibrium concentration of sulfur, $S_2(g)$.

**What is Given?**
The equation for the equilibrium system is: $2H_2S(g) \leftrightarrow 2H_2(g) + S_2(g)$
The equilibrium constant , $K_c = 2.4 \times 10^{-4}$
Equilibrium concentration $[H_2(g)] = 0.013$ mol/L
Equilibrium concentration $[H_2S(g)] = 0.18$ mol/L

**Plan Your Strategy**
Write the equilibrium expression, substitute the known values into this expression and calculate the equilibrium concentration for $S_2(g)$.

**Act on Your Strategy**

The expression for the equilibrium constant is: $K_c = \dfrac{[H_2(g)^2[S_2(g)]}{[H_2S(g)]^2}$

Substituting the known values: $2.4 \times 10^{-4} = \dfrac{(0.013)^2[S_2(g)]}{(0.18)^2}$

$[S_2(g)] = \dfrac{(2.4\times10^{-4})(0.18)^2}{(0.013)^2} = 0.046$ mol/L

**Check Your Solution**

The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number units (mol/L) and the correct number of significant digits (2).

**10.**

**Problem**

Methane, ethyne, and hydrogen form the following equilibrium mixture:

$$2CH_4(g) \leftrightarrow C_2H_2(g) + 3H_2(g)$$

While studying this reaction, a chemist analyzed a 4.0 L sealed flask containing an equilibrium mixture of the gases at 1700 °C and found 0.64 mol of $C_2H_2(g)$ and 0.92 mol of $H_2(g)$. Given that $K_c = 0.15$ for the reaction at 1700 °C, what concentration should the chemist expect for $CH_4(g)$ at equilibrium?

**What is Required?**

You must calculate the equilibrium concentration of $CH_4(g)$.

**What is Given?**

The equation for this equilibrium system is: $2CH_4(g) \leftrightarrow C_2H_2(g) + 3H_2(g)$
The equilibrium constant, $K_c = 0.15$
The volume of the container = 4.0 L
Amount of $C_2H_2(g)$ at equilibrium = 0.64 mol
Amount of $H_2(g)$ = 0.92 mol

**Plan Your Strategy**

Calculate the molar concentration of each compound at equilibrium.
Write the equilibrium expression and substitute these equilibrium molar concentrations of $C_2H_2(g)$ and $H_2(g)$ and the given value of $K_c$ into this expression. Calculate the molar equilibrium concentration of $CH_4(g)$.

**Act on Your Strategy**

The molar equilibrium concentrations are:

$[C_2H_2(g)] = \dfrac{0.64\,\text{mol}}{4.0\,\text{L}} = 0.16$ mol/L

$[H_2(g)] = \dfrac{0.92\,\text{mol}}{4.0\,\text{L}} = 0.23$ mol/L

The equilibrium expression is $K_c = \dfrac{[C_2H_2(g)][H_2(g)]^3}{[CH_4(g)]^2}$

Substitute the known values into this expression: $0.15 = \dfrac{(0.16)(0.23)^3}{[CH_4(g)]^2}$

$[CH_4(g)] = \sqrt{\dfrac{(0.16)(0.23)^3}{0.15}} = 0.11$ mol/L

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct units (mol/L) and the correct number of significant digits (2).

**11.**
**Problem**
For the reaction $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$, the value of $K_c$ is 25.0 at 1100 K and $8.0 \times 10^2$ at room temperature, 300 K. Which temperature favours the dissociation of hydrogen iodide into its component gases?

**What is Required?**
You must choose the temperature that favours the greater concentration of reactants, $H_2(g)$ and $I_2(g)$.

**What is Given?**
$K_c = 25.0$ at 1100 K
$K_c = 8.0 \times 10^2$ at 300 K

**Plan Your Strategy**
Since HI(g) is a product in this chemical equilibrium, a smaller value of $K_c$ corresponds to a greater amount of reactant.

**Act on Your Strategy**
The value of $K_c$ is smaller at 1100 K. The position of equilibrium lies farther to the left and favours the formation of $H_2(g)$ and $I_2(g)$ at the higher temperature.

**Check Your Solution**

Since the expression for $K_c$ has the reactant terms in the denominator, the smaller value of $K_c$ must correspond to the larger concentration of reactants.

**12.**
**Problem**
Three reactions and their equilibrium constants are given below:
I. $N_2(g) + O_2(g) \leftrightarrow 2NO(g)$ $K_c = 4.7 \times 10^{-31}$
II. $2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$ $K_c = 1.8 \times 10^{-6}$
III. $N_2O_4(g) \leftrightarrow 2NO_2(g)$ $K_c = 0.025$
Arrange these reactions in order of their tendency to form products.

**What is Required?**
You must arrange the equations for the three reactions in increasing order that favours the formation of products.

**What is Given?**
Three equilibrium systems are given with the corresponding values of $K_c$. All three systems are at the same temperature.

**Plan Your Strategy**
A larger value of $K_c$ corresponds to a greater amount of product. The order favouring the formation of products is the same as the order for the value of $K_c$.

**Act on Your Strategy**
$K_c$ (III) > $K_c$ (II) > $K_c$ (I), so the order favouring the formation of products is reaction III > II > I.

**Check Your Solution**
Since the expression for $K_c$ has the product term in the numerator, the larger value of $K_c$ must correspond to the larger concentration of product.

**13.**
**Problem**
Identify each of the following reactions as essentially proceeding to completion or not taking place:
**a)** $N_2(g) + 3Cl_2(g) \leftrightarrow 2NCl_3(g)$ $K_c = 3.0 \times 10^{11}$
**b)** $2CH_4(g) \leftrightarrow C_2H_6(g) + H_2(g)$ $K_c = 9.5 \times 10^{-13}$
**c)** $2NO(g) + 2CO(g) \leftrightarrow N_2(g) + 2CO_2(g)$ $K_c = 2.2 \times 10^{59}$

**What is Given?**
Three equations representing systems at equilibrium are given with the corresponding values of $K_c$.

**What is Required?**
You must determine which reactions go to completion and which essentially do not take place.

**Plan Your Strategy**
When $K > 1$, products are favoured. The equilibrium lies to the right. Reactions in which $K$ is greater than $10^{10}$ are usually regarded as proceeding to completion.
When $K \approx 1$, there are approximately equal concentrations of reactants and products at equilibrium.
When $K < 1$, reactants are favoured. The equilibrium lies to the left. Reactions in which $K$ is smaller than $10^{-10}$ are usually regarded as not taking place at all.

**Act on Your Strategy**
Using the above criteria, reactions (a) and (c) go to completion and reaction (b) essentially does not proceed to the right.

**Check Your Solution**
The answer is consistent with the criteria followed for predicting the position of equilibrium.

**14.**
**Problem**
Consider the following reaction:
$H_2(g) + Cl_2(g) \leftrightarrow 2HCl(g)$ $K_c = 2.4 \times 10^{33}$ at 25 °C
If a quantity of HCl(g) is placed into a reaction vessel, to what extent do you expect the equilibrium mixture will dissociate into hydrogen and chlorine?

**What is Required?**
You must determine the extent to which the reactants are favoured in the given equilibrium system.

**What is Given?**
An equilibrium system is given with the corresponding value of $K_c$.

**Plan Your Strategy**
When $K > 1$, products are favoured. The equilibrium lies to the right. Reactions in which $K$ is greater than $10^{10}$ are usually regarded as proceeding to completion.
When $K < 1$, reactants are favoured. The equilibrium lies to the left. Reactions in which $K$ is smaller than $10^{-10}$ are usually regarded as not taking place at all.

**Act on Your Strategy**
Since $K >> 1$, the equilibrium position is far to the right, favouring the product, HCl(g). Therefore, there is only the smallest amount of reactant present at equilibrium. There is essentially no dissociation of HCl(g) into $H_2(g)$ and $Cl_2(g)$.

**Check Your Solution**
The answer is consistent with the criteria followed for predicting the position of equilibrium.

**15.**
**Problem**

Most metal ions combine with other ions present in solution. For example, in aqueous ammonia, silver (Ag) ions are at equilibrium with different complex ions.

$$[Ag(H_2O)_2]^+(aq) + 2NH_3(aq) \leftrightarrow [Ag(NH_3)_2]^+(aq) + 2H_2O(l)$$

At room temperature, $K_c$ for this reaction is $1\times10^7$. Which of the two silver complex ions is the more stable? Explain your reasoning.

**What is Given?**
An equation representing the silver ion at equilibrium with two different complexes is given with the $K_c$ for the system.

**What is Required?**
You must determine which complex is favoured, the reactant, $[Ag(H_2O)_2]^+(aq)$, or product, $[Ag(NH_3)_2]^+(aq)$.

**Plan Your Strategy**
When $K > 1$, products are favoured. The equilibrium lies to the right.

**Act on Your Strategy**
Since $K_c = 1\times10^7$, the product complex, $[Ag(NH_3)_2]^+(aq)$, is more stable.

**Check Your Solution**
The answer is consistent with the criteria followed for predicting the position of equilibrium.

**16.**
**Problem**
Consider the following reaction:

$$2HI(g) \leftrightarrow H_2(g) + I_2(g) \quad \Delta H = -52 \text{ kJ}$$

In which direction does the reaction shift if the temperature increases?

**What is Required?**
You must determine whether increasing the temperature causes the equilibrium to shift to the left or to the right.

**What is Given?**
The balanced equation is given with the $\Delta H$ term for the reaction at equilibrium.

**Plan Your Strategy**
Use the $\Delta H$ term to identify the reaction as exothermic. Apply Le Châtelier's Principle to determine the shift in equilibrium that will minimize an increase in temperature.

**Act on Your Strategy**
The temperature is increased. Therefore, the equilibrium must shift in the direction in which heat is used up, i.e. the direction in which the reaction is endothermic. The reaction is endothermic from right to left. The reaction will shift to the left when the temperature is increased.

**Check Your Solution**

An increase in temperature must result in a shift in the reaction that minimizes the temperature increase. This answer is in agreement with Le Châtelier's Principle.

**17.**
**Problem**
In the gaseous equilibrium systems below, the volume of the container is increased, causing a decrease in pressure. In which direction does each reaction shift? Explain why for each case.
**a)** $CO_2(g) + H_2(g) \leftrightarrow CO(g) + H_2O(g)$
**b)** $2NO_2(g) \leftrightarrow N_2O_4(g)$
**c)** $2CO_2(g) \leftrightarrow 2CO(g) + O_2(g)$
**d)** $CH_4(g) + 2H_2S(g) \leftrightarrow CS_2(g) + 4H_2(g)$

**What is Required?**
You must determine whether decreasing the pressure causes the equilibrium to shift to the left, to the right, or if there is no effect.

**What is Given?**
The balanced equations are given for each equilibrium system.

**Plan Your Strategy**
Apply Le Châtelier's Principle to determine the shift in equilibrium that will minimize a decrease in pressure. The shift must be towards the side with the largest number of gas molecules.

**Act on Your Strategy**
**a)** As the reaction proceeds there is no change in the number of gas molecules on each side of the equation. Therefore, increasing the volume of the container has no effect on the position of equilibrium.
**b)** There are more gas molecules on the left side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the left.
**c)** There are more gas molecules on the right side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the right.
**d)** There are more gas molecules on the right side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the right.

**Check Your Solution**
The reaction must shift in the direction that minimizes the pressure decrease. The shift must be in the direction in which more gas molecules are formed. This answer is in agreement with Le Châtelier's Principle.

**18.**
**Problem**
The following reaction is exothermic from left to right:

$$2NO(g) + 2H_2(g) \leftrightarrow N_2(g) + 2H_2O(g)$$

In which direction does the reaction shift as a result of each of the following changes?
**a)** removing hydrogen gas
**b)** increasing the pressure of gases in the reaction vessel by decreasing the volume

**c)** increasing the pressure of gases in the reaction vessel by pumping in argon gas while keeping the volume of the vessel constant
**d)** increasing the temperature
**e)** using a catalyst

**What is Required?**
You need to determine whether each change causes the equilibrium to shift to the left, to the right, or if there is no effect.

**What is Given?**
The balanced equations are given for each equilibrium system and the reactions are exothermic from left to right.

**Plan Your Strategy**
Apply Le Châtelier's Principle to determine the shift in equilibrium that will minimize each identified change to the system.

**Act on Your Strategy**
**a)** [$H_2$(g)] is reduced. The equilibrium must shift to increase [$H_2$(g)]. The reaction shifts to the left.
**b)** The pressure is increased by decreasing the volume of the reaction vessel. The equilibrium must shift to decrease the pressure. Because there are fewer gas molecules on the right of the equation, the reaction shifts to the right.
**c)** Argon does not react with any of the gases in the mixture. Since the volume of the system does not change, the addition of argon gas does not represent a stress on the system. The position of equilibrium does not change.
**d)** The temperature increases. The equilibrium must shift in the direction in which heat is used i.e. the direction in which the reaction is endothermic. The reaction is endothermic from right to left. The reaction shifts to the left if the temperature is increased.
**e)** A catalyst has no effect on the position of equilibrium since the activation energy of both the forward and reverse reactions are affected to the same extent.

**Check Your Solution**
Each answer is in agreement with Le Châtelier's Principle.

**19.**
**Problem**
In question 18, which changes affect the value of $K_c$? Which changes do not affect the value of $K_c$?

**What is Required?**
You must determine which changes affect the value of $K_c$.

**What is Given?**
The balanced equations are given for each equilibrium system and the reactions are exothermic from left to right.

**Plan Your Strategy**
For a given reaction at equilibrium, $K_c$ depends on temperature.

**Act on Your Strategy**
Change (d) is the only one that affects the temperature of the reaction. Therefore, change (d) is the only one that affects the value of $K_c$. Changes (a), (b), (c), and (e) have no effect on the value of $K_c$.

**Check Your Solution**
The value of *K*c for a particular equilibrium system depends on temperature.

**20.**
**Problem**
Toluene, $C_7H_8(g)$, is an important organic solvent. It is made industrially from methyl cyclohexane:

$$C_7H_{14}(g) \leftrightarrow C_7H_8(g) + 3H_2(g)$$

The forward reaction is endothermic. State three different changes to an equilibrium mixture of these reacting gases that would shift the reaction toward greater production of toluene.

**What is Required?**
You must identify three different changes that would shift the equilibrium toward greater production of toluene.

**What is Given?**
The equation for the equilibrium system is: $C_7H_{14}(g) \leftrightarrow C_7H_8(g) + 3H_2(g)$

The reaction is endothermic from left to right.

**Plan Your Strategy**
Each change must shift the reaction to the right. According to Le Châtelier's Principle, the equilibrium will shift to minimize any change applied to the system. Consider changes in the volume of the system (pressure), temperature and adding or removing one of the reactants or products that will shift the equilibrium to the right as the change is minimized.

**Act on Your Strategy**
If the reactant, methylcyclohexane, $C_7H_{14}(g)$, is added, the equilibrium will shift to the right because this will use up this reactant and minimize the change. Therefore, increasing $[C_7H_{14}]$ increases the production of toluene.
If either product, toluene ($C_7H_8(g)$) or hydrogen ($H_2(g)$), is removed, the equilibrium will shift to the right because this will produce more of these products and minimize the change. Therefore, decreasing either $[C_7H_8]$ or $[H_2]$, or both, increases the production of toluene.
Since the number of gas molecules increases as the reaction proceeds from left to right, an increase in the volume of the system in the reaction vessel (decrease in pressure) will result in a shift to the right and increase the production of toluene.

Since the reaction is endothermic from left to right, an increase in temperature will cause a shift to the right to minimize this change. Therefore, an increase in temperature will increase the production of toluene.

**Check Your Solution**
This answer is in agreement with Le Châtelier's Principle.

**21.**
**Problem**
At 1100 K, hydrogen and iodine combine to form hydrogen iodide:

$$H_2(g) + I_2(g) \leftrightarrow 2HI(g)$$

At equilibrium in a 1.0 L reaction vessel, the mixture of gases contained 0.30 mol of $H_2(g)$, 1.3 mol of $I_2(g)$, and 3.4 mol of HI(g). What is the value of $K_c$ at 1100 K?

**What is Required?**
You must calculate the value of $K_c$.

**What is Given?**
The equation for the system at equilibrium is: $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$
The equilibrium concentrations of each gas are:
$H_2(g)$ = 0.30 mol/L
$I_2(g)$ = 1.3 mol/L
HI(g) = 3.4 mol/L

**Plan Your Strategy**
Write the equilibrium expression and substitute the equilibrium molar concentrations into the expression for $K_c$. Calculate the value of $K_c$.

**Act on Your Strategy**
The chemical equation for this equilibrium is: $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$

The equilibrium expression for this system is $K_c = \dfrac{[HI(g)]^2}{[H_2(g)][I_2(g)]}$

$$K_c = \frac{(3.4)^2}{(0.30)(1.3)} = 30$$

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number of significant digits (2).

**22.**
**Problem**
At 25 °C, the following reaction takes place:

$$I_2(g) + Cl_2(g) \leftrightarrow 2ICl(g)$$

$K_c$ for the reaction is 82. If 0.83 mol of iodine gas and 0.83 mol of chlorine gas are placed in a 10 L container at 25 °C, what will the concentrations of the various gases be at equilibrium?

**What is Required?**
You must calculate the equilibrium concentrations of each gas.

**What is Given?**
The equilibrium system is $I_2(g) + Cl_2(g) \leftrightarrow 2ICl(g)$
The equilibrium constant, $K_c$, = 82
The volume of the system = 10 L
Initial amount $I_2(g)$ = 0.83 mol
Initial amount of $Cl_2(g)$ = 0.83 mol

**Plan Your Strategy**
**Step 1**: Calculate the initial molar concentrations of $I_2(g)$ and $Cl_2(g)$.
**Step 2:** Set up an ICE table. Use the initial concentrations of $I_2(g)$ and $Cl_2(g)$ calculated from **Step 1**. Let the change in molar concentration of $I_2(g)$ and $Cl_2(g)$ be x. Use the stoichiometry of the chemical equation to derive the expression for the equilibrium concentrations.
**Step 3:** Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression and solve for x.
**Step 4:** Substitute x into the equilibrium line of the ICE table to find the equilibrium concentration of $I_2(g)$, $Cl_2(g)$ and $ICl(g)$.

**Act on Your Strategy**
**Step 1:** Initial concentrations:

$$[I_2(g)] = [Cl_2(g)] = \frac{0.83\,\text{mol}}{10\,\text{L}} = 0.0.083\ \text{mol/L}$$

**Step 2:** ICE table.

| Concentration (mol/L) | $I_2(g)$ + | $Cl_2(g)$ $\leftrightarrow$ | $2ICl(g)$ |
|---|---|---|---|
| initial | 0.083 | 0.083 | 0.00 |
| change | – x | – x | + 2x |
| equilibrium | 0.083 – x | 0.083 – x | 2x |

**Step 3**

$$K_c = \frac{[ICl(g)]^2}{[I_2(g)][Cl_2(g)]} = 82$$

$$\frac{(2x)^2}{(0.083 - x)^2} = 82$$

Take the square root of each side.

$$\sqrt{\frac{4x^2}{(0.083\text{-}x)^2}} = \sqrt{82}$$

$$\frac{2x}{0.083 - x} = \pm 9.06$$

x = +0.068 mol/L or x = −0.11 mol/L

The value x = − 0.11 mol/L is physically impossible.

Therefore, at equilibrium
$[I_2(g)] = [Cl_2(g)] = 0.083 - x = 0.083 - 0.068 = 0.015$ mol/L
$[ICl(g)] = 2x = 2(0.068) = 0.14$ mol/L

**Check Your Solution**
Substituting the calculated concentrations of each gas at equilibrium into the equilibrium expression:

$$K_c = \frac{[ICl(g)]^2}{[I_2(g)][Cl_2(g)]} = \frac{(0.14)^2}{(0.015)^2} = 87 \approx 82.$$

The calculated concentrations have the correct unit (mol/L) and the correct number of significant digits (2)

**23.**
**Problem**
A chemist was studying the following reaction at a certain temperature:

$$SO_2(g) + NO_2(g) \leftrightarrow NO(g) + SO_3(g)$$

In a 1.0 L container, the chemist added $1.7 \times 10^{-1}$ mol of sulfur dioxide to $1.1 \times 10^{-1}$ mol of nitrogen dioxide. At equilibrium, the concentration of $SO_3(g)$ was found to be 0.089 mol/L. What is the value of $K_c$ for the reaction at this temperature?

**What is Required?**
You must calculate the value of $K_c$ for the given equilibrium.

**What is Given?**
The equation for the equilibrium system is $SO_2(g) + NO_2(g) \leftrightarrow NO(g) + SO_3(g)$
Volume of the container = 1.0 L
Initial amount of $SO_2(g) = 1.7 \times 10^{-1}$ mol
Initial amount of $NO_2(g) = 1.1 \times 10^{-1}$ mol
Equilibrium concentration $[SO_3(g)] = 0.089$ mol/L.

**Plan Your Strategy**
**Step 1**: Calculate the initial molar concentrations of $SO_2(g)$ and $NO_2(g)$.

**Step 2:** Set up an ICE table. Using the initial concentrations of $SO_2(g)$ and $NO_2(g)$, the given equilibrium concentration of $SO_3(g)$,and the stoichiometry of the chemical equation, the change in concentration that occurred to reach equilibrium can be calculated. Since there was no $SO_3(g)$ or NO(g) initially present, the equilibrium concentration of $SO_3(g)$ represents the change in concentration that occurred to reach equilibrium. Calculate the equilibrium concentrations of $SO_2(g)$, $NO_2(g)$, and NO(g).
**Step 3:** Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression and solve for $K_c$.

**Act on Your Strategy**
**Step 1:** Initial concentrations:
$[SO_2(g)] = 1.7 \times 10^{-1}$ mol/L
$[NO_2(g)] = 1.1 \times 10^{-1}$ mol/L

**Step 2:** ICE table.

| Concentration (mol/L) | $SO_2(g)$ + | $NO_2(g)$ ↔ | NO(g) + | $SO_3(g)$ |
|---|---|---|---|---|
| initial | $1.7 \times 10^{-1}$ | $1.1 \times 10^{-1}$ | 0.00 | 0.00 |
| change | – 0.089 | – 0.089 | + 0.089 | + 0.089 |
| equilibrium | 0.17 – 0.089 = 0.081 | 0.11 – 0.089 = 0.021 | 0.089 | 0.089 |

**Step 3:**
$$K_c = \frac{[NO(g)]\,[SO_3(g)]}{[SO_2(g)][NO_2(g)]} = \frac{(0.089)^2}{(0.081)(0.021)} = 4.7$$

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number of significant digits (2).

**24.**
**Problem**
Hydrogen bromide decomposes at 700 K:
$$2HBr(g) \leftrightarrow H_2(g) + Br_2(g) \quad K_c = 4.2 \times 10^{-9}$$
0.090 mol of HBr is placed into a 2.0 L reaction vessel and heated to 700 K. What is the equilibrium concentration of each gas?

**What is Required?**
You must calculate the equilibrium concentrations of each gas.

**What is Given?**
The equilibrium system is $2HBr(g) \leftrightarrow H_2(g) + Br_2(g)$
The equilibrium constant, $K_c = 4.2 \times 10^{-9}$
The volume of the system = 2.0 L

Initial amount HBr(g) = 0.090 mol

**Plan Your Strategy**
**Step 1:** Calculate the initial molar concentrations of HBr(g).
**Step 2:** Set up an ICE table. The initial concentration of HBr(g) is used from S**tep 1**. Let the change in molar concentration of HBr(g) be 2x. Use the stoichiometry of the chemical equation to derive the expression for the equilibrium concentrations.
**Step 3:** Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression. Solve for x.
**Step 4:** Substitute x into the equilibrium line of the ICE table to find the equilibrium concentration of HBr(g), $H_2$(g), and $Br_2$(g).

**Act on Your Strategy**
**Step 1:** Initial concentration:

$$[\text{HBr(g)}] = \frac{0.090\,\text{mol}}{2.0\,\text{L}} = 0.045\ \text{mol/L}$$

**Step 2:** ICE Table.

| Concentration (mol/L) | 2HBr(g) ↔ | $H_2$(g) + | $Br_2$(g) |
|---|---|---|---|
| initial | 0.045 | 0.00 | 0.00 |
| change | – 2x | + x | + x |
| equilibrium | 0.045 – 2x | x | x |

**Step 3:**

$$K_c = \frac{[\text{H}_2\text{(g)}][\text{Br}_2\text{(g)}]}{[\text{HBr(g)}]^2} = 4.2 \times 10^{-9}$$

$$\frac{(\text{x})^2}{(0.045 - 2\text{x})^2} = 4.2 \times 10^{-9}$$

Take the square root of each side.

$$\sqrt{\frac{\text{x}^2}{(0.045\text{-x})^2}} = \sqrt{4.2 \times 10^{-9}}$$

$$\frac{\text{x}}{0.045 - \text{x}} = \pm\ 6.48 \times 10^{-5}$$

x = $\pm 2.9 \times 10^{-6}$ mol/L

The value x = $-2.9 \times 10^{-6}$ is physically impossible because it would result in negative concentration for both $H_2$(g) and $Br_2$(g).

Therefore, at equilibrium

$[HBr(g)] = 0.045 - 2(2.9 \times 10^{-6}) \approx 0.045$ mol/L
$[H_2(g)] = [Br_2(g)] = x = 2.9 \times 10^{-6}$ mol/L

**Check Your Solution**
Substituting the calculated concentrations of each gas at equilibrium into the equilibrium expression:

$K_c = \dfrac{[H_2(g)][Br_2(g)]}{[HBr(g)]^2} = \dfrac{(2.9\times10^{-6})^2}{(0.045)^2} = 4.2 \times 10^{-9}$. This is the same $K_c$ as was given.

The calculated concentrations have the correct unit (mol/L) and the correct number of significant digits (2).

**25.**
**Problem**
The following equation represents the equilibrium reaction for the dissociation of phosgene gas:

$$COCl_2(g) \leftrightarrow CO(g) + Cl_2(g)$$

At 100 °C, the value of $K_c$ for this reaction is $2.2 \times 10^{-8}$. The initial concentration of $COCl_2(g)$ in a closed container at 100 °C is 1.5 mol/L. What are the equilibrium concentrations of $CO(g)$ and $Cl_2(g)$?

**What is Required?**
You must calculate the equilibrium concentrations of $CO(g)$ and $Cl_2(g)$.

**What is Given?**
The equation for the equilibrium system is: $COCl_2(g) \leftrightarrow CO(g) + Cl_2(g)$
The equilibrium constant, $K_c = 2.2 \times 10^{-8}$
Initial concentration of $COCl_2(g) = 1.5$ mol/L

**Plan Your Strategy**
**Step 1:**
Divide the initial concentration by the value of $K_c$. If the answer is greater than 1000, the change in that concentration can be ignored.
**Step 2:**
Set up an ICE Table. Let x represent the change in concentration of $COCl_2(g)$. Use the stoichiometry of the chemical equation to derive the expression for the equilibrium concentrations.
**Step 3:**
Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression. Solve for x.
**Step 4** Substitute x into the equilibrium line of the ICE table to find the equilibrium concentrations of $CO(g)$ and $Cl_2(g)$.

**Act on Your Strategy**
**Step 1:**

$$\frac{\text{initial concentration } COCl_2}{K_c} = \frac{1.5}{2.2 \times 10^{-8}} >> 1000$$

The change in the initial concentration can be ignored.

**Step 2:** ICE Table

| concentration (mol/L) | $COCl_2(g)$ $\rightleftharpoons$ | $CO(g)$ + | $Cl_2(g)$ |
|---|---|---|---|
| initial | 1.5 | 0.00 | 0.00 |
| change | – x | + x | + x |
| equilibrium | 1.5 – x ≈ 1.5 | x | x |

**Step 3:**

$$K_c = \frac{[CO(g)][Cl_2(g)]}{[COCl_2(g)]} = \frac{x^2}{1.5} = 2.2 \times 10^{-8}$$

$x^2 = 3.3 \times 10^{-8}$
$x = \pm\ 1.8 \times 10^{-4}$

Therefore at equilibrium
$[CO(g)] = [Cl_2(g)] = 1.8 \times 10^{-4}$ mol/L

**Check Your Solution**
Since $1.5 - 1.8 \times 10^{-4}$ is essentially 1.5, the first assumption was valid.
Checking the equilibrium values,

$$K_c = \frac{(1.8 \times 10^{-4})^2}{1.5} = 2.2 \times 10^{-8}$$ which is the original value of $K_c$.

The answer has the correct unit (mol/L) and the correct number of significant digits (2).

**26.**
**Problem**
Hydrogen sulfide is a poisonous gas with a characteristic offensive odour. At 1400 °C, the gas dissociates, with $K_c$ equal to $2.4 \times 10^{-4}$:

$$2H_2S(g) \leftrightarrow 2H_2(g) + S_2(g)$$

4.0 mol of $H_2S$ is placed in a 3.0 L container. What is the equilibrium concentration of hydrogen gas at 1400 °C?

**What is Required?**
You must calculate the equilibrium concentration of $H_2(g)$.

**What is Given?**

The equation for the equilibrium system is: $2H_2S(g) \leftrightarrow 2H_2(g) + S_2(g)$

The equilibrium constant, $K_c = 2.4 \times 10^{-4}$

Volume of container = 3.0 L

Initial amount of $H_2S(g)$ = 4.0 mol

**Plan Your Strategy**

**Step 1:**

Calculate the initial molar concentration of $H_2S(g)$. Divide the initial concentration by the value of $K_c$. If the answer is greater than 1000, the change in that concentration can be ignored.

**Step 2:**

Set up an ICE Table. Let 2x represent the change in concentration of $H_2S(g)$. Use the stoichiometry of the chemical equation to derive the expression for the equilibrium concentrations.

**Step 3:**

Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression. Solve for x.

**Step 4:**

Substitute x into the equilibrium line of the ICE table to find the equilibrium concentration of $H_2(g)$.

**Act on Your Strategy**

**Step 1:** initial concentration of $H_2S(g) = \dfrac{4.0\,\text{mol}}{3.0\text{L}} = 1.33$ mol/L

$$\frac{\text{initial concentration } H_2S(g)}{K_c} = \frac{1.33}{2.4\times10^{-4}} > 1000$$

The change in the initial concentration can be ignored.

**Step 2:** ICE Table

| Concentration (mol/L) | $2H_2S(g)$ $\leftrightarrow$ | $2H_2(g)$ + | $S_2(g)$ |
|---|---|---|---|
| initial | 1.33 | 0.00 | 0.00 |
| change | – 2x | + 2x | + x |
| equilibrium | 1.33 – x ≈ 1.33 | 2x | x |

**Step 3:**

$$K_c = \frac{[H_2(g)]^2[S_2(g)]}{[H_2S(g)]^2} = \frac{(2x)^2(x)}{1.33} = 2.4 \times 10^{-4}$$

$4x^3 = 4.25\times 10^{-4}$

$x^3 = 1.06 \times 10^{-4}$

$x = \sqrt[3]{1.06 \times 10^{-4}} = 4.7 \times 10^{-2}$

Therefore at equilibrium, $[H_2(g)] = 2x = 2(4.7 \times 10^{-2}) = 9.4 \times 10^{-2}$ mol/L

**Check Your Solution**

Since $1.3 - 8.6 \times 10^{-2}$ is essentially 1.3, the first assumption was valid.
Checking the equilibrium values,

$K_c = \dfrac{(9.4 \times 10^{-2})^2 (4.3 \times 10^{-2})}{(1.3)^2} = 2.2 \times 10^{-4}$ which is within rounding error equal to the given value of $K_c$. The answer has the correct unit (mol/L) and the correct number of significant digits (2).

**27.**

**Problem**

At a particular temperature, $K_c$ for the decomposition of carbon dioxide gas is $2.0 \times 10^{-6}$:

$$2CO_2(g) \leftrightarrow 2CO(g) + O_2(g)$$

3.0 mol of $CO_2$ is put in a 5.0 L container. Calculate the equilibrium concentration of each gas.

**What is Required?**

You must calculate the equilibrium concentration of each gas in the equilibrium system.

**What is Given?**

The equation for the equilibrium system is: $2CO_2(g) \leftrightarrow 2CO(g) + O_2(g)$
The equilibrium constant, $K_c = 2.0 \times 10^{-6}$
Volume of container = 5.0 L
Initial amount of $CO_2(g)$ = 3.0 mol

**Plan Your Strategy**

**Step 1:**
Calculate the initial molar concentration of $CO_2(g)$. Divide the initial concentration by the value of $K_c$. If the answer is greater than 1000, the change in that concentration can be ignored.
**Step 2:**
Set up an ICE Table. Let 2x represent the change in concentration of $CO_2(g)$. Use the stoichiometry of the chemical equation to derive the expression for the equilibrium concentrations of each gas.
**Step 3:**
Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression. Solve for x.
**Step 4:**
Substitute x into the equilibrium line of the ICE table to find the equilibrium concentration of each gas.

**Act on Your Strategy**

**Step 1:** initial concentration of $CO_2(g) = \dfrac{3.0\text{ mol}}{5.0\text{L}} = 0.60$ mol/L

$$\frac{\text{initial concentration } CO_2(g)}{K_c} = \frac{0.60}{2.0\times10^{-6}} >> 1000$$

The change in the initial concentration can be ignored.

**Step 2:** ICE Table

| concentration (mol/L) | $2CO_2(g)$ ↔ | $2CO(g)$ + | $O_2(g)$ |
|---|---|---|---|
| initial | 0.60 | 0.00 | 0.00 |
| change | − 2x | +2 x | + x |
| equilibrium | 0.60 − 2x ≈ 0.60 | 2x | x |

**Step 3:**

$$K_c = \frac{[CO(g)]^2[O_2(g)]}{[CO_2(g)]^2} = \frac{(2x)^2(x)}{(0.60)^2} = 2.0 \times 10^{-6}$$

$4x^3 = 7.20 \times 10^{-7}$

$x^3 = 1.80 \times 10^{-7}$

$x = \sqrt[3]{1.80\times10^{-7}} = 5.646 \times 10^{-3}$

Therefore at equilibrium,
$[CO_2(g)] = 0.60$ mol/L
$[CO(g)] = 2x = 2(5.646 \times 10^{-3}) = 1.1 \times 10^{-2}$ mol/L
$[O_2(g)] = x = 5.6 \times 10^{-3}$ mol/L

**Check Your Solution**
Since $0.60 - 5.646 \times 10^{-3}$ is essentially 0.60 the first assumption was valid.
Checking the equilibrium values,

$$K_c = \frac{(1.1\times10^{-2})^2(5.6\times10^{-3})}{(0.60)^2} = 1.9 \times 10^{-6} \approx \text{original value of } K_c = 2.0 \times 10^{-6}$$

The answer has the correct units (mol/L) and the correct number of significant digits(2).

**28.**
**Problem**
At a certain temperature, the value of $K_c$ for the following reaction is $3.3 \times 10^{-12}$:

$$2NCl_3(g) \leftrightarrow N_2(g) + 3Cl_2(g)$$

A certain amount of nitrogen trichloride is put in a 1.0 L reaction vessel at this temperature. At equilibrium, $4.6 \times 10^{-6}$ mol of nitrogen gas is present. What amount of $NCl_3(g)$ was put in the reaction vessel?

**What is Required?**
You must calculate the amount of $NCl_3(g)$ initially placed in the reaction vessel.

**What is Given?**
The equation fro the equilibrium system is: $2NCl_3(g) \leftrightarrow N_2(g) + 3Cl_2(g)$
The equilibrium constant, $K_c = 3.3 \times 10^{-12}$
Volume of the container = 1.0 L
Amount of $N_2(g)$ at equilibrium = $4.6 \times 10^{-6}$ mol

**Plan Your Strategy**
**Step 1:**
Set up an ICE Table. Let x represent the original concentration of $NCl_3(g)$. The change in concentration of $N_2(g)$ will be its equilibrium concentration. Use the stoichiometry of the chemical equation to derive the equilibrium concentrations. The equilibrium concentration of $Cl_2(g)$ will be 3 times that of $N_2(g)$ and the change in concentration of $NCl_3(g)$ will be 2 times the equilibrium concentration that of $N_2(g)$.
**Step 2:**
Write the expression for $K_c$, calculate the equilibrium concentration of $NCl_3(g)$, and solve for x. This is also the initial amount since the volume of the container is 1.0 L.

**Act on Your Strategy**
**Step 1:** ICE Table

| concentration (mol/L) | $2NCl_3(g)$ $\leftrightarrow$ | $N_2(g)$ + | $3Cl_2(g)$ |
|---|---|---|---|
| initial | x | 0.00 | 0.00 |
| change | $-2(4.6 \times 10^{-6}) = -9.2 \times 10^{-6}$ | $+4.6 \times 10^{-6}$ | $+3(4.6 \times 10^{-6}) = 1.38 \times 10^{-5}$ |
| equilibrium | $x - 9.2 \times 10^{-6}$ | $4.6 \times 10^{-6}$ | $1.38 \times 10^{-5}$ |

**Step 2:**

$$K_c = \frac{[N_2(g)][Cl_2(g)]^3}{[NCl_3(g)]^2} = \frac{(4.6\times10^{-6})(1.38\times10^{-5})^3}{(x - 9.2\times10^{-6})^2}$$

$$(x - 9.2 \times 10^{-6})^2 = 1.21 \times 10^{-20}$$

$$(x - 9.2 \times 10^{-6}) = \sqrt{1.21\times10^{-20}} = 1.0995 \times 10^{-10}$$

$$x = 1.0995 \times 10^{-10} + 9.2 \times 10^{-6} = 9.2 \times 10^{-6} \text{ mol}$$

Therefore, the initial amount of $NCl_3(g)$ in the reaction vessel is $9.2 \times 10^{-6}$ mol in 1.0 L.

**Check Your Solution**
The equilibrium expression has the product terms in the numerator and the reactant terms in the denominator. The exponents in the equilibrium expression match the corresponding coefficients in the chemical equation. The answer has the correct number of significant digits (2).

**29.**
**Problem**
At a certain temperature, the value of $K_c$ for the following reaction is $4.2 \times 10^{-8}$:

$$N_2(g) + O_2(g) \leftrightarrow 2NO(g)$$

0.45 mol of nitrogen gas and 0.26 mol of oxygen gas are put in a 6.0 L reaction vessel. What is the equilibrium concentration of NO(g) at this temperature?

**What is Required?**
You must calculate the equilibrium concentration of NO(g).

**What is Given?**
The equation for the equilibrium system is: $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
The equilibrium constant, $K_c = 4.2 \times 10^{-8}$
Volume of container = 6.0 L
Initial amount of $N_2(g)$ = 0.45 mol
Initial amount of $O_2(g)$ = 0.26 mol

**Plan Your Strategy**
**Step 1:**
Calculate the initial molar concentrations of $N_2(g)$ and $O_2(g)$. Divide the smallest initial concentration by the value of $K_c$. If the answer is greater than 1000, the change in that concentration can be ignored.
**Step 2:**
Set up an ICE Table. Let x represent the change in concentration of $N_2(g)$. Use the stoichiometry of the chemical equation to derive the expression for the equilibrium concentrations of each gas.
**Step 3:**
Write the equilibrium expression. Substitute the values for the equilibrium concentrations into this expression. Solve for x.
**Step 4:**
Substitute x into the equilibrium line of the ICE table to find the equilibrium concentration of $H_2(g)$.

**Act on Your Strategy**
**Step 1.**

initial concentration of $N_2(g) = \dfrac{0.45\text{ mol}}{6.0\text{L}} = 0.0750$ mol/L

initial concentration of $O_2(g) = \dfrac{0.26\text{ mol}}{6.0\text{ L}} = 0.0433$ mol/L

$$\frac{\text{initial concentration } O_2(g)}{K_c} = \frac{0.0433}{4.2 \times 10^{-8}} >> 1000$$

The change in the initial concentration can be ignored.

**Step 2:** ICE Table

| concentration (mol/L) | $N_2(g)$ + | $O_2(g)$ $\rightleftharpoons$ | $2NO(g)$ |
|---|---|---|---|
| initial | 0.0750 | 0.0433 | 0.00 |
| change | − x | − x | + 2 x |
| equilibrium | 0.075 − x ≈ 0.075 | 0.0433 − x ≈0.0433 | 2x |

**Step 3:**

$$K_c = \frac{[NO(g)]^2}{[N_2(g)][O_2(g)]} = \frac{(2x)^2}{(0.075)(0.0433)} = 4.2 \times 10^{-8}$$

$4x^2 = 1.36 \times 10^{-10}$

$x = 5.839 \times 10^{-6}$

Therefore at equilibrium, $[NO(g)] = 2x = 2(5.839 \times 10^{-6}) = 1.2 \times 10^{-5}$ mol/L

**Check Your Solution**

Since $0.0433 - 5.839 \times 10^{-6}$ is essentially 0.0433 the first assumption was valid.
Checking the equilibrium values,

$$K_c = \frac{(1.2 \times 10^{-5})^2}{(0.075)(0.0433)} = 4.4 \times 10^{-8} \approx \text{original value of } K_c = 4.2 \times 10^{-8}$$

The answer has the correct unit (mol/L) and the correct number of significant digits (2).