Chapter 7

Reversible Reactions and Chemical Equilibrium

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Write the equilibrium expression for each homogeneous reaction.

1. Problem
The reaction between ethanol and ethanoic acid to form ethyl ethanoate and water:
\[ \text{CH}_3\text{CH}_2\text{OH}(\ell) + \text{CH}_3\text{COOH}(\ell) \rightleftharpoons \text{CH}_3\text{COOCH}_2\text{CH}_3(\ell) + \text{H}_2\text{O}(\ell) \]

What Is Required?
You need to find an expression for \( K \).

What Is Given?
You know the balanced chemical equation.

Plan Your Strategy
\( K_c \) is a fraction with product concentrations in the numerator and concentrations of reactants in the denominator. Each concentration term must be raised to the power of the coefficient in the balanced equation.

Act on Your Strategy
\[ K_c = \frac{[\text{CH}_3\text{COOCH}_2\text{CH}_3][\text{H}_2\text{O}]}{[\text{CH}_3\text{CH}_2\text{OH}][\text{CH}_3\text{COOH}]} \]

Check Your Solution
Each chemical formula is enclosed in square brackets, indicating concentration. The products are in the numerator, and the reactants in the denominator. There are no coefficients in the balanced equation, and therefore no power terms in the equilibrium expression (other than the value “1”).

2. Problem
The reaction between nitrogen gas and oxygen gas at high temperatures:
\[ \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \]

What Is Required?
You need to find an expression for \( K \).

What Is Given?
You know the balanced chemical equation.

Plan Your Strategy
\( K_c \) is a fraction with product concentrations in the numerator and concentrations of reactants in the denominator. Each concentration term must be raised to the power of the coefficient in the balanced equation.

Act on Your Strategy
\[ K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} \]
Check Your Solution
The products are in the numerator, and the reactants in the denominator. Each chemical formula is enclosed in square brackets, and the concentration of NO is raised to the power of the coefficient in the chemical equation.

3. Problem
The reaction between hydrogen gas and oxygen gas to form water vapour:

\[ 2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g) \]

What Is Required?
You need to find an expression for \( K_c \).

What Is Given?
You know the balanced chemical equation.

Plan Your Strategy
The expression for \( K_c \) is a fraction. The concentration of 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 the coefficient in the balanced equation.

Act on Your Strategy
\[
K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]} 
\]

Check Your Solution
The products are in the numerator, and the reactants in the denominator. Each chemical formula is enclosed in square brackets, and the concentrations of \( H_2O \) and \( H_2 \) are raised to the power of their respective coefficient in the chemical equation.

4. Problem
The reduction-oxidation equilibrium of iron and iodine ions in aqueous solution:

\[ 2Fe^{3+}(aq) + 2I^-(aq) \rightleftharpoons 2Fe^{2+}(aq) + I_2(aq) \]

Note: You will learn about reduction-oxidation reactions in the next unit.

What Is Required?
You need to find an expression for \( K_c \).

What Is Given?
You know the balanced chemical equation.

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 the 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 products are in the numerator, and the reactants in the denominator. Each chemical formula is enclosed in square brackets, and the concentrations of \( Fe^{3+} \), \( I^- \), and \( Fe^{2+} \) are raised to the power of their respective coefficient in the chemical equation.

5. Problem
The oxidation of ammonia (one of the reactions in the production of nitric acid):

\[ 4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \]
What Is Required?
You need to find an expression for $K_c$.

What Is Given?
You know the balanced chemical equation.

Plan Your Strategy
The expression for $K_c$ is a fraction. The concentrations of 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 the coefficient in the balanced equation.

Act on Your Strategy
$$K_c = \frac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^3}$$

Check Your Solution
The products are in the numerator, and the reactants in the denominator. Each chemical formula is enclosed in square brackets, and the concentration of each chemical is raised to the power of its coefficient in the chemical equation.

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6. Problem
The following reaction took place in a sealed flask at 250˚C.
$$\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$$
At equilibrium, the gases in the flask had the following concentrations:
$[\text{PCl}_5] = 1.2 \times 10^{-2}$ mol/L, $[\text{PCl}_3] = 1.5 \times 10^{-2}$ mol/L, and $[\text{Cl}_2] = 1.5 \times 10^{-2}$ mol/L. Calculate the value of $K_c$ at 250˚C.

What Is Required?
You need to calculate the value of $K_c$.

What Is Given?
You have the chemical equation and the concentration of each substance at equilibrium.

Plan Your Strategy
Write the equilibrium expression. Then substitute the equilibrium molar concentrations into the expression.

Act on Your Strategy
The equilibrium expression is
$$K_c = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$
Substitute the equilibrium molar concentrations into the expression.
$$K_c = \frac{(1.5 \times 10^{-2}) \times (1.5 \times 10^{-2})}{(1.2 \times 10^{-2})}$$
$$K_c = 1.9 \times 10^{-2}$$

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 molar concentrations at equilibrium were substituted into the expression.
7. Problem
Iodine and bromine react to form iodine monobromide, IBr.

\[ \text{I}_2(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{IBr}(\text{g}) \]

At 250°C, an equilibrium mixture in a 2.0 L flask contained 0.024 mol of I\(_2\)(g), 0.050 mol of Br\(_2\)(g), and 0.38 mol of IBr(g). What is the value of \(K_c\) for the reaction at 250°C?

**What Is Required?**
You need to calculate the value of \(K_c\).

**What Is Given?**
You have the chemical equation, the amount of each substance at equilibrium, and the volume of the flask.

**Plan Your Strategy**

**Step 1** Calculate the molar concentration of each compound at equilibrium.

\[
\begin{align*}
[I_2] &= \frac{0.024 \text{ mol}}{2.0 \text{ L}} = 0.012 \text{ mol/L} \\
[Br_2] &= \frac{0.050 \text{ mol}}{2.0 \text{ L}} = 0.025 \text{ mol/L} \\
[IBr] &= \frac{0.38 \text{ mol}}{2.0 \text{ L}} = 0.19 \text{ mol/L}
\end{align*}
\]

**Step 2** Write the equilibrium expression. Then substitute the equilibrium molar concentrations into the expression.

\[
K_c = \frac{[IBr]^2}{[I_2][Br_2]}
\]

\[
= \frac{(0.19)^2}{(0.012)(0.025)}
\]

\[
= 1.2 \times 10^2
\]

**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 molar concentrations at equilibrium were substituted into the expression. The answer has the correct number of significant digits.

8. Problem
At high temperatures, carbon dioxide gas decomposes into carbon monoxide and oxygen gas. At equilibrium, the gases have the following concentrations:

\[
[\text{CO}_2(\text{g})] = 1.2 \text{ mol/L}, \quad [\text{CO}(\text{g})] = 0.35 \text{ mol/L}, \quad \text{and} \quad [\text{O}_2(\text{g})] = 0.15 \text{ mol/L}.
\]

Determine \(K_c\) at the temperature of the reaction.

**What Is Required?**
You need to calculate the value of \(K_c\).

**What Is Given?**
You know the reactant and products, and the concentration of each substance at equilibrium.
Plan Your Strategy

**Step 1** Write the chemical equation.

**Step 2** Write the equilibrium expression. Then substitute the equilibrium molar concentrations into the expression for \( K_c \).

**Act on Your Strategy**

**Step 1** The chemical equation is

\[ 2\text{CO}_2(\text{g}) \rightleftharpoons 2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \]

**Step 2**

\[
K_c = \frac{[\text{CO}]^2[\text{O}_2]}{[\text{CO}_2]^2} = \frac{(0.35)^2(0.15)}{(1.2)^2} = 0.013
\]

**Check Your Solution**

The chemical equation is balanced. 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 molar concentrations at equilibrium were substituted into the expression. The answer has the correct number of significant digits.

9. Problem

Hydrogen sulfide is a pungent, poisonous gas. At 1400 K, an equilibrium mixture was found to contain 0.013 mol/L hydrogen, 0.046 mol/L sulfur in the form of \( \text{S}_2(\text{g}) \), and 0.18 mol/L hydrogen sulfide. Calculate the equilibrium constant, at 1400 K, for the following reaction.

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

**What Is Required?**

You need to calculate the value of \( K_c \).

**What Is Given?**

You have the chemical equation and the amount of each substance at equilibrium.

**Plan Your Strategy**

Write the equilibrium expression. Then substitute the equilibrium molar concentrations into the expression.

**Act on Your Strategy**

\[
K_c = \frac{[\text{H}_2]^2[\text{S}_2]}{[\text{H}_2\text{S}]^2} = \frac{(0.013)^2(0.046)}{(0.18)^2} = 2.4 \times 10^{-4}
\]

**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 molar concentrations at equilibrium were substituted into the expression.
10. Problem
Methane, ethyne, and hydrogen form the following equilibrium mixture.

\[ 2\text{CH}_4(g) \rightleftharpoons \text{C}_2\text{H}_2(g) + 3\text{H}_2(g) \]

While studying this reaction mixture, a chemist analysed a 4.0 L sealed flask at
1700°C. The chemist found 0.46 mol of \( \text{CH}_4(g) \), 0.64 mol of \( \text{C}_2\text{H}_2(g) \), and 0.92 mol
of \( \text{H}_2(g) \). What is the value of \( K_c \) for the reaction at 1700°C?

What Is Required?
You need to calculate the value of \( K_c \).

What Is Given?
You have the chemical equation and the amount of each substance at equilibrium
in a 4.0 L flask.

Plan Your Strategy
Step 1 Calculate the molar concentration of each compound at equilibrium.
Step 2 Write the equilibrium expression. Then substitute the equilibrium molar
concentrations into the expression.

Act on Your Strategy
Step 1 The reaction takes place in a 4.0 L flask. Calculate the molar concentrations
at equilibrium.
\[
[\text{CH}_4] = \frac{0.46 \text{ mol}}{4.0 \text{ L}} = 0.115 \text{ mol/L}
\]
\[
[\text{C}_2\text{H}_2] = \frac{0.64 \text{ mol}}{4.0 \text{ L}} = 0.16 \text{ mol/L}
\]
\[
[\text{H}_2] = \frac{0.92 \text{ mol}}{4.0 \text{ L}} = 0.23 \text{ mol/L}
\]
Step 2 Write the equilibrium expression. Substitute the equilibrium molar
concentrations into the expression.
\[
K_c = \frac{[\text{C}_2\text{H}_2][\text{H}_2]^3}{[\text{CH}_4]^2}
\]
\[
= \frac{(0.16)(0.23)^3}{(0.115)^2}
\]
\[
= 0.15
\]

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
responding coefficients in the chemical equation. The molar concentrations at
equilibrium were substituted into the expression. The answer has the correct number
of significant digits.

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11. Problem
At 25°C the value of \( K_c \) for the following reaction is 82.

\[ \text{I}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{ICl}_2(g) \]

0.83 mol of \( \text{I}_2(g) \) and 0.83 mol of \( \text{Cl}_2(g) \) are placed in a 10 L container at 25°C.

What are the concentrations of the three gases at equilibrium?

What Is Required?
You need to find \([\text{I}_2]\), \([\text{Cl}_2]\) and \([\text{ICl}_2]\) at equilibrium.

What Is Given?
You are given the chemical equation. You know the initial amount of each gas,
the volume of the container, and the equilibrium constant.
Plan Your Strategy

**Step 1** Calculate the initial concentrations.

**Step 2** Set up an ICE table. Insert the initial concentrations you calculated in Step 1 in your ICE table. Let the change in molar concentrations of the reactants be \( x \). Use the stoichiometry of the chemical equation to write and record expressions for the equilibrium concentrations.

**Step 3** Write the equilibrium expression. Substitute the expressions for the equilibrium concentrations into the expression. Solve the equilibrium expression for \( x \).

**Step 4** Calculate the equilibrium concentration of each gas.

Act on Your Strategy

**Step 1** The initial amount of \( \text{I}_2 \) is equal to the initial amount of \( \text{Cl}_2 \).

\[
[\text{I}_2] = [\text{Cl}_2] = \frac{0.83 \text{ mol}}{10 \text{ L}} = 0.083 \text{ mol/L}
\]

**Step 2** Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>I(<em>2)(</em>{\text{g}})</th>
<th>+</th>
<th>Cl(<em>2)(</em>{\text{g}})</th>
<th>⇋</th>
<th>2ICl(_{\text{g}})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.083</td>
<td></td>
<td>0.083</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td></td>
<td>(-x)</td>
<td></td>
<td>(+2x)</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.083 (-x)</td>
<td></td>
<td>0.083 (-x)</td>
<td></td>
<td>2(x)</td>
</tr>
</tbody>
</table>

**Step 3**

\[
K_c = \frac{[\text{ICl}]^2}{[\text{I}_2][\text{Cl}_2]}
\]

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

\[
= \frac{(2x)^2}{(0.083 - x)^2}
\]

The right side of the equation is a perfect square.

\[
\pm 9.06 = \frac{(2x)}{(0.083 - x)}
\]

Solving for both values,

\( x = 0.068 \) and \( x = 0.11 \)

**Step 4** The value \( x = 0.11 \) is physically impossible because it is a larger concentration than was initially placed in the container. To achieve equilibrium, the initial \([\text{I}_2]\) and \([\text{Cl}_2]\) must decrease.

\[
[\text{I}_2] = [\text{Cl}_2] = 0.083 - 0.068 = 0.015 \text{ mol/L}
\]

\[
[\text{ICl}] = 2 \times 0.068 = 0.14 \text{ mol/L}
\]

Check Your Solution

The equilibrium expression has product concentrations in the numerator and reactant concentrations in the denominator. Check \( K_c \):

\[
K_c = \frac{(0.14)^2}{(0.015)^2} = 87
\]

This is close to the given value, \( K_c = 82 \). The difference is due to mathematical rounding.

12. Problem

At a certain temperature, \( K_c = 4.0 \) for the following reaction.

\[2\text{HF}_{\text{g}} \rightleftharpoons \text{H}_2\text{(g)} + \text{F}_2\text{(g)}\]

A 1.0 L reaction vessel contained 0.045 mol of \( \text{F}_2\)\(_{\text{g}}\) at equilibrium. What was the initial amount of HF in the reaction vessel?
What Is Required?
You need to find the initial amount of HF\(_{(g)}\) placed in the reaction vessel.

What Is Given?
You are given the chemical equation. You know the amount of F\(_2(g)\) at equilibrium, the volume of the container, and the equilibrium constant.

Plan Your Strategy

**Step 1** Calculate the concentration of F\(_2(g)\) at equilibrium and use the chemical equation to find the equilibrium concentration of H\(_2(g)\).

**Step 2** Set up an ICE table. Insert the equilibrium concentrations you calculated in Step 1 in your ICE table. Let the change in molar concentrations of the reactants be \(x\). Use the stoichiometry of the chemical equation to write and record expressions for the equilibrium concentrations.

**Step 3** Write the equilibrium expression. Substitute the expressions for the equilibrium concentrations into the expression. Solve the equilibrium expression for the initial [HF].

**Step 4** Calculate the initial amount of HF\(_{(g)}\).

Act on Your Strategy

**Step 1** The volume of the reaction vessel is 1.0 L. Therefore, 
\[
[F_2] = \frac{0.045 \text{ mol}}{1.0 \text{ L}} = 0.045 \text{ mol/L}
\]
\[
[H_2] = [F_2] = 0.045 \text{ mol/L}
\]

**Step 2** Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>2HF(_{(g)})</th>
<th>⇋</th>
<th>H(_2(g))</th>
<th>+</th>
<th>F(_2(g))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>[HF]</td>
<td>=</td>
<td>0</td>
<td>+</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>−2x</td>
<td>+x</td>
<td>+x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>([HF] − 2x)</td>
<td>0.045</td>
<td>0.045</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Therefore, \(x = 0.045\)

**Step 3** \(K_c = \frac{[H_2][F_2]}{[HF]^2}\)

\[
4.0 = \frac{(0.045)(0.045)}{([HF] − 0.090)^2}
\]

\[
= \frac{(0.045)^2}{([HF] − 0.090)^2}
\]

The right side of the equation is a perfect square.

\[
±2.0 = \frac{(0.045)}{([HF] − 0.090)}
\]

Solving for both values gives

[H\(_F\)] = 0.112 and [HF] = 0.0675

**Step 4** The value [HF] = 0.0675 is physically impossible because subtracting the change in concentration, 0.090, results in a negative value for the initial concentration of HF\(_{(g)}\) placed in the container. Therefore, [HF] = 0.11 mol/L.

Because the volume of the reaction vessel is 1.0 L, the initial amount of HF\(_{(g)}\) placed in the reaction vessel was 0.11 mol.
Check Your Solution
The equilibrium expression has product concentrations in the numerator and reactant concentrations in the denominator. Check \(K_c\): 
\[
K_c = \frac{(0.045)^2}{(0.11 - 0.090)^2} = 5.1
\]
This is close to the given value, \(K_c = 4.0\). The difference is due to mathematical rounding.

13. Problem
A chemist was studying the following reaction.
\[
\text{SO}_2(g) + \text{NO}_2(g) \rightleftharpoons \text{NO}(g) + \text{SO}_3(g)
\]
In a 1.0 L container, the chemist added \(1.7 \times 10^{-1}\) mol of \(\text{SO}_2(g)\) to \(1.1 \times 10^{-1}\) mol of \(\text{NO}_2(g)\). The value of \(K_c\) for the reaction at a certain temperature is 4.8. What is the equilibrium concentration of \(\text{SO}_3(g)\) at this temperature?

What Is Required?
You need to find the equilibrium \([\text{SO}_3]\).

What Is Given?
You are given the chemical equation. You know the equilibrium constant for the reaction, \(K_c = 4.8\). You also know the initial amounts of \(\text{SO}_2(g)\) and \(\text{NO}_2(g)\), and the volume of the container.

Plan Your Strategy
Step 1 Calculate the initial concentrations of \(\text{SO}_2(g)\) and \(\text{NO}_2(g)\).
Step 2 Set up an ICE table. Insert the equilibrium concentrations you calculated in Step 1 in your ICE table. Let the change in molar concentrations of the reactants be \(x\). Use the stoichiometry of the chemical equation to write and record expressions for the equilibrium concentrations.
Step 3 Write the equilibrium expression. Substitute the expressions for the equilibrium concentrations into the expression. Rearrange the equilibrium expression into the form of a quadratic equation for \(x\). Solve the quadratic equation for \(x\).
Step 4 Substitute \(x\) into the equilibrium line of the ICE table to find the equilibrium concentration of \([\text{SO}_3]\).

Act on Your Strategy
Step 1 The volume of the reaction vessel is 1.0 L. Therefore,
\[
[\text{SO}_2] = \frac{1.7 \times 10^{-1}\text{ mol}}{1.0 \text{ L}} = 1.7 \times 10^{-1}\text{ mol/L}
\]
\[
[\text{NO}_2] = \frac{1.1 \times 10^{-1}\text{ mol}}{1.0 \text{ L}} = 1.1 \times 10^{-1}\text{ mol/L}
\]
Step 2 Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>(\text{SO}_2(g))</th>
<th>(\text{NO}_2(g))</th>
<th>(\text{NO}(g))</th>
<th>(\text{SO}_3(g))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>(1.7 \times 10^{-1})</td>
<td>(1.1 \times 10^{-1})</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(-x)</td>
<td>(+x)</td>
<td>(+x)</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>((1.7 \times 10^{-1}) - x)</td>
<td>((1.1 \times 10^{-1}) - x)</td>
<td>(x)</td>
<td>(x)</td>
</tr>
</tbody>
</table>

Step 3 \(K_c = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]}\)
\[
4.8 = \frac{(x)(x)}{(0.17 - x)(0.11 - x)}
\]
This equation must be rearranged into a quadratic equation. 

3.8x^2 − 1.344x + 0.08976 = 0

Recall that a quadratic equation of the form \( ax^2 + bx + c = 0 \) has the following solution.

\[
x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}
\]

Therefore,

\[
x = \frac{-(-1.344) \pm \sqrt{1.806 - 1.364}}{7.6}
\]

\[
x = 1.344 \pm 0.6648
\]

\[
x = 0.089 \text{ and } x = 0.264
\]

The value \( x = 0.264 \) is not physically possible. It would result in negative concentration of SO\(_2\) and NO\(_2\) at equilibrium. The concentration of each substance at equilibrium is found by substituting \( x = 0.089 \) into the last line of the ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>SO(_2)(g) + NO(_2)(g) ⇌ NO(_2)(g) + SO(_3)(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equilibrium</td>
<td>(0.17 - 0.089) (0.11 - 0.089) 0.089 0.089</td>
</tr>
</tbody>
</table>

Applying the rules for subtraction involving measured values:

[SO\(_2\)] = 0.08 mol/L
[NO\(_2\)] = 0.02 mol/L
[NO] = 0.089 mol/L
[SO\(_3\)] = 0.089 mol/L

At equilibrium the concentration of SO\(_3\)(g) is 0.089 mol/L.

**Check Your Solution**

The equilibrium expression has product concentrations in the numerator and reactant concentrations in the denominator. Check \( K_c \):

\[
K_c = \frac{(0.089)^2}{(0.08)(0.02)} = 5
\]

Solving the quadratic equation gave a value of \( x \) known only to two decimal places. Subtracting the value of \( x \) from [SO\(_2\)] and [NO\(_2\)] gave values only good to one significant figure. The calculated value of \( K_c \) is equal to the given value, 4.8, within the error introduced by rounding.

14. **Problem**

Phosgene, COCl\(_2\)(g), is an extremely toxic gas. It was used during World War I. Today, it is used to manufacture pesticides, pharmaceuticals, dyes, and polymers. It is prepared by mixing carbon monoxide and chlorine gas.

\[
CO(g) + Cl_2(g) ⇌ COCl_2(g)
\]

0.055 mol of CO\(_(g)\) and 0.072 mol of Cl\(_2\)(g) are placed in a 5.0 L container. At 870 K, the equilibrium constant is 0.20. What are the equilibrium concentrations of the mixture at 870 K?

**What Is Required?**

You need to find the equilibrium concentrations of CO, Cl\(_2\) and COCl.

**What Is Given?**

You are given the chemical equation. You know the equilibrium constant for the reaction, \( K_c = 0.20 \). You also know the initial amounts of CO\(_(g)\) and Cl\(_2\)(g), and the volume of the container.
Plan Your Strategy

Step 1  Calculate the initial concentrations of CO\(_{(g)}\) and Cl\(_2\)(g).

Step 2  Set up an ICE table. Insert the equilibrium concentrations you calculated in Step 1 in your ICE table. Let the change in molar concentrations of the reactants be \(x\). Use the stoichiometry of the chemical equation to write and record expressions for the equilibrium concentrations.

Step 3  Write the equilibrium expression. Substitute the expressions for the equilibrium concentrations into the expression. Rearrange the equilibrium expression into the form of a quadratic equation for \(x\). Solve the quadratic equation for \(x\).

Step 4  Substitute \(x\) into the equilibrium line of the ICE table to find the equilibrium concentrations of CO, Cl\(_2\), and COCl\(_2\).

Act on Your Strategy

Step 1  The volume of the reaction vessel is 5.0 L. Therefore,

\[
[\text{CO}] = \frac{0.055 \text{ mol}}{5.0 \text{ L}} = 0.011 \text{ mol/L}
\]

\[
[\text{Cl}_2] = \frac{0.072 \text{ mol}}{5.0 \text{ L}} = 0.0144 \text{ mol/L}
\]

Step 2  Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>CO(_{(g)})</th>
<th>+</th>
<th>Cl(_2)(g)</th>
<th>(\rightleftharpoons)</th>
<th>COCl(_2)(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.011</td>
<td>0.0144</td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(-x)</td>
<td>+x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>((0.011 - x))</td>
<td>((0.0144 - x))</td>
<td>(x)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Step 3  \(K_c = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}\)

\[
0.20 = \frac{x}{(0.011-x)(0.0144-x)}
\]

This equation must be rearranged into a quadratic equation

\[
x^2 - 5.025x + 1.58 \times 10^{-4} = 0
\]

Recall that a quadratic equation of the form \(ax^2 + bx + c = 0\) has the following solution.

\[
x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}
\]

Therefore,

\[
x = \frac{-(5.025) \pm \sqrt{25.25 - 6.32 \times 10^{-5}}}{2.0}
\]

\[
x = \frac{5.025 \pm 5.025}{2.0}
\]

\[
x = 5.02 \text{ or } x = 0.0
\]

The value \(x = 5.02\) is not physically possible. It would result in negative concentration of CO and Cl\(_2\) at equilibrium. The concentration of reactants remains essentially unchanged with very little formation of COCl\(_2\)(g).

Check Your Solution

Calculate the equilibrium \([\text{COCl}_2]\) to check that it is a very small value.

\[
K_c = 0.20 = \frac{[\text{COCl}_2]}{0.011 \times 0.014}
\]

\[
[\text{COCl}_2] = 3.1 \times 10^{-5} \text{ mol/L}
\]

The equilibrium concentration of \([\text{COCl}_2]\) is a very small value.
15. Problem

Hydrogen bromide decomposes at 700 K.

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

0.090 mol of HBr is placed in a 2.0 L reaction vessel and heated to 700 K. What is the equilibrium concentration of each gas?

What Is Required?

You need to find \([\text{HBr}], [\text{H}_2], \text{and} [\text{Br}_2]\) at equilibrium.

What Is Given?

You are given the chemical equation. You know the initial amount of HBr, the volume of the container, and the equilibrium constant.

Plan Your Strategy

**Step 1** Calculate the initial concentration of HBr.

**Step 2** Set up an ICE table. Insert the initial concentrations you calculated in Step 1 in your ICE table. Let the change in molar concentrations of HBr be \(x\). Use the stoichiometry of the chemical equation to write and record expressions for the equilibrium concentrations.

**Step 3** Write the equilibrium expression. Substitute the expressions for the equilibrium concentrations into the expression. Solve the equilibrium expression for \(x\).

**Step 4** Calculate the equilibrium concentration of each gas.

Act on Your Strategy

**Step 1**

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

**Step 2**

Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>2\text{HBr}(g)</th>
<th>⇌</th>
<th>\text{H}_2(g)</th>
<th>+</th>
<th>\text{Br}_2(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.045</td>
<td></td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>−2(x)</td>
<td>+(x)</td>
<td>+(x)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.045 − 2(x)</td>
<td>(x)</td>
<td>(x)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Step 3**

\[ K_c = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2} \]

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

The right side of the equation is a perfect square.

\[ \pm 6.48 \times 10^{-5} = \frac{x}{0.045 - 2x} \]

Solving for both values,

\(x = +2.92 \times 10^{-6}\) and \(x = −2.92 \times 10^{-6}\)

**Step 4**

The value \(x = −2.92 \times 10^{-6}\) is physically impossible because it would result in negative concentration of both \(\text{H}_2\) and \(\text{Br}_2\).

Therefore, \([\text{H}_2] = [\text{Br}_2] = 2.9 \times 10^{-6} \text{ mol/L}\).

\([\text{HBr}] = 0.045 - (2 \times 2.9 \times 10^{-6}) = 0.045 \text{ mol/L}\)

Check Your Solution

The equilibrium expression has product concentrations in the numerator and reactant concentrations in the denominator. The equilibrium concentrations are in mol/L.

Check \(K_c\):

\[ K_c = \frac{(2.9)^2}{(0.045)^2} = 4.2 \times 10^{-9} \]

This is equal to the given value of \(K_c\).
16. Problem
For the reaction $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$, 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 $\text{HI}(\text{g})$ into its component gases?

What Is Required?
Choose the temperature that favours the greater dissociation of $\text{HI}$.

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

Plan Your Strategy
$\text{HI}$ is on the right in the chemical equation. A greater dissociation of hydrogen iodide corresponds to a smaller value of $K_c$.

Act on Your Strategy
The value of $K_c$ is smaller at 1100 K. The position of equilibrium lies more to the left and favours the dissociation of hydrogen iodide at the higher temperature.

Check Your Solution
The question asked to choose the conditions for a smaller concentration of a product. $K_c$ is a fraction with product terms in the numerator, so the smallest value of $K_c$ must correspond to the smallest concentration of product.

17. Problem
Three reactions, and their equilibrium constants, are given below.

   1. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$ \quad $K_c = 4.7 \times 10^{-31}$
   2. $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ \quad $K_c = 1.8 \times 10^{-6}$
   3. $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ \quad $K_c = 0.025$

Arrange these equations in the order of their tendency to form products.

What Is Required?
Place the reactions in the order of their tendency to form products.

What Is Given?
$K_c$ is given for each reaction.

Plan Your Strategy
The equilibrium constant, $K_c$, is a ratio of product concentrations divided by reactant concentrations. The larger the value of $K_c$, the greater the tendency to form products.

Act on Your Strategy
The largest value of $K_c$ is 0.025, then $1.8 \times 10^{-6}$, and finally, $4.7 \times 10^{-31}$. From greatest to least tendency to form products, the order is III, II, I.

Check Your Solution
The question asked about the tendency to form products. $K_c$ is a ratio of product concentrations divided by reactant concentrations. The largest value of $K_c$ corresponds to the greatest tendency to form products.
18. Problem
Identify each reaction as essentially going to completion or not taking place.

(a) \(N_2(g) + 3Cl_2(g) \rightleftharpoons 2NCl_3(g) \quad K_c = 3.0 \times 10^{11}\)

(b) \(2CH_4(g) \rightleftharpoons C_2H_6(g) + H_2(g) \quad K_c = 9.5 \times 10^{-13}\)

(c) \(2NO(g) + 2CO(g) \rightleftharpoons N_2(g) + 2CO_2(g) \quad K_c = 2.2 \times 10^{59}\)

What Is Required?
For each reaction, determine whether the equilibrium mixture consists mostly of products or reactants.

What Is Given?
\(K_c\) is given for each reaction.

Plan Your Strategy
The equilibrium constant, \(K_c\), is a ratio of product concentrations divided by reactant concentrations. A large value of \(K_c\) corresponds to mostly products, and a small value of \(K_c\) indicates mostly reactants.

Act on Your Strategy
Reaction (a): \(K_c = 3.0 \times 10^{11}\). At equilibrium, mostly products are present and the reaction essentially goes to completion.

Reaction (b): \(K_c = 9.5 \times 10^{-13}\). At equilibrium, mostly reactants are present and there is essentially no reaction.

Reaction (c): \(K_c = 2.2 \times 10^{59}\). At equilibrium, mostly products are present and the reaction goes to completion.

Check Your Solution
\(K_c\) is a ratio of product concentrations divided by reactant concentrations. Reactions that go to completion have very large values for \(K_c\). Reactions that do not take place have very small values for \(K_c\).

19. Problem
Most metal ions combine with other ions in solution. For example, in aqueous ammonia, silver(I) ions are in equilibrium with different complex ions.

\([Ag(H_2O)_2]^+_{(aq)} + 2NH_3_{(aq)} \rightleftharpoons [Ag(NH_3)_2]^+_{(aq)} + 2H_2O_{(l)}\)

At room temperature, \(K_c\) for this reaction is \(1 \times 10^7\). Which of the two silver complex ions is the more stable? Explain your reasoning.

What Is Required?
Decide which silver ion is more stable.

What Is Given?
\(K_c\) is given for the reaction.

Plan Your Strategy
The value of \(K_c\) indicates the relative concentrations of products and reactants at equilibrium.

Act on Your Strategy
\(K_c\) for this reaction is \(1 \times 10^7\). Therefore, the concentration of products is much greater than the concentration of reactants. Consequently, \([Ag(NH_3)_2]^+_{(aq)}\) must be the more stable silver ion.

Check Your Solution
\(K_c\) is a ratio of product concentrations divided by reactant concentrations. \(K_c\) for the reaction is a large value, therefore the concentration of silver ion on the product side of the equation must be greater than the concentration of silver ion on the reactant side of the equation.
20. Problem
Consider the following reaction.
\[ \text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{HCl}(\text{g}) \]
\[ K_c = 2.4 \times 10^{33} \text{ at } 25^\circ \text{C} \]
HCl(\text{g}) is placed into a reaction vessel. To what extent do you expect the equilibrium mixture to dissociate into H\textsubscript{2}(\text{g}) and Cl\textsubscript{2}(\text{g})?

What Is Required?
Describe the extent of dissociation of HCl at 25°C.

What Is Given?
\[ K_c = 2.4 \times 10^{33} \text{ at } 25^\circ \text{C} \]

Plan Your Strategy
HCl is on the right in the chemical equation. A small dissociation of HCl corresponds to a large [HCl] and a large value of \( K_c \).

Act on Your Strategy
The value of \( K_c \) is large. The position of equilibrium lies to the right and very little HCl dissociates at 25°C.

Check Your Solution
\( K_c \) is a fraction with product terms in the numerator. A large value of \( K_c \) corresponds to large concentration of product terms. Because HCl is on the product side of the equation, essentially no dissociation takes place.

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21. Problem
The following equation represents the equilibrium reaction for the dissociation of phosgene gas.
\[ \text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{Cl}_2(\text{g}) \]
At 100°C, the value of \( K_c \) for this reaction is \( 2.2 \times 10^{-8} \). The initial concentration of COCl\textsubscript{2}(\text{g}) in a closed container at 100°C is 1.5 mol/L. What are the equilibrium concentrations of CO(\text{g}) and Cl\textsubscript{2}(\text{g})?

What Is Required?
You need to find [CO] and [Cl\textsubscript{2}] at equilibrium.

What Is Given?
You are given the chemical equation. You know the value of \( K_c \), and the initial concentration of COCl\textsubscript{2}(\text{g}).

Plan Your Strategy
Step 1 Divide the initial concentration of COCl\textsubscript{2}(\text{g}) by \( K_c \) to determine whether you can ignore the change in concentration.
Step 2 Set up an ICE table. Let \( x \) represent the change in [COCl\textsubscript{2}].
Step 3 Write the equilibrium expression. Substitute the equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for \( x \).
Step 4 Calculate [CO] and [Cl\textsubscript{2}] at equilibrium.

Act on Your Strategy
Step 1
\[
\frac{\text{initial concentration}}{K_c} = \frac{1.5}{2.2 \times 10^{-8}} = 6.8 \times 10^7
\]
Because this is well above 500, the change in [COCl\textsubscript{2}] can be ignored.
Step 2  Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>COCl₂(g)</th>
<th>⇌</th>
<th>CO(g)</th>
<th>+</th>
<th>Cl₂(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.5</td>
<td></td>
<td>0</td>
<td>+</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>−x</td>
<td>+x</td>
<td></td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.5 − x ≈ 1.5</td>
<td>x</td>
<td>x</td>
<td>x</td>
<td></td>
</tr>
</tbody>
</table>

Step 3  

\[
K_c = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]} = 2.2 \times 10^{-8}
\]

\[
x = \sqrt{3.3 \times 10^{-8}} = 1.8 \times 10^{-4}
\]

Step 4  

\[[\text{CO}] = [\text{Cl}_2] = 1.8 \times 10^{-4} \text{ mol/L}\]

Check Your Solution
The assumption that \(x\) is negligible compared with the initial concentration is valid because, using the rules for subtracting measured quantities,

\[
1.5 - (1.8 \times 10^{-4}) = 1.5.
\]

Next, check the equilibrium values:

\[
\frac{(1.8 \times 10^{-4})^2}{1.5} = 2.2 \times 10^{-8}
\]

This is equal to the equilibrium constant.

22. Problem
Hydrogen sulfide is a poisonous gas with a characteristic, offensive odour. It dissociates at 1400˚C, with \(K_c\) equal to \(2.4 \times 10^{-4}\).

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

4.0 mol of \(\text{H}_2\text{S}\) is placed in a 3.0 L container. What is the equilibrium concentration of \(\text{H}_2(g)\) at 1400˚C?

What Is Required?
You need to find the concentration of \(\text{H}_2(g)\) at equilibrium.

What Is Given?
You are given the chemical equation. You know the value of \(K_c\), and the initial amount of \(\text{H}_2(g)\). The volume of the container is given.

Plan Your Strategy
Step 1  Calculate the initial molar concentration of \(\text{H}_2\text{S}(g)\).
Step 2  Divide the initial concentration of \(\text{H}_2\text{S}(g)\) by \(K_c\) to determine whether you can ignore the change in concentration.
Step 3  Set up an ICE table. Let \(x\) represent the change in \([\text{H}_2\text{S}(g)]\).
Step 4  Write the equilibrium expression. Substitute the equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for \(x\).
Step 5  Calculate \([\text{H}_2]\) at equilibrium.

Act on Your Strategy
Step 1  \([\text{H}_2\text{S}] = \frac{4.0 \text{ mol}}{3.0 \text{ L}} = 1.33 \text{ mol/L}\)
Step 2  

\[
\frac{\text{Initial concentration}}{K_c} = \frac{1.33}{2.4 \times 10^{-4}} = 5.5 \times 10^3
\]

Because this is well above 500, the change in \([\text{H}_2\text{S}]\) can be ignored.
Step 3  Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>H₂S(g)</th>
<th>⇄</th>
<th>2H₂(g)</th>
<th>+</th>
<th>S₂(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.33</td>
<td>0</td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>−x</td>
<td>+2x</td>
<td>+x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.33−x</td>
<td>2x</td>
<td>x</td>
<td></td>
<td>3√7.98×10⁻⁵</td>
</tr>
</tbody>
</table>

Step 4  

\[ K_c = \frac{[H_2]^2[S_2]}{[H_2S]} \]

\[ 2.4 \times 10^{-4} = \frac{(2x)^2(x)}{1.33} \]

\[ x = \sqrt{7.98 \times 10^{-5}} \]

\[ = 4.3 \times 10^{-2} \]

Step 4  

\[ [H_2] = 2 \times (4.3 \times 10^{-2}) = 8.6 \times 10^{-2} \text{ mol/L} \]

Check Your Solution

The assumption that \( x \) is negligible compared with the initial concentration is valid because, using the rules for subtracting measured quantities,

\[ 1.3 - (4.3 \times 10^{-2}) = 1.3 \]

Next, check the equilibrium values:

\[ \frac{(8.6 \times 10^{-2})^2(4.3 \times 10^{-2})}{1.3} = 2.4 \times 10^{-4} \]

This is equal to the equilibrium constant.

23. Problem

At a certain temperature, the value of \( K_c \) for the following reaction is \( 3.3 \times 10^{-12} \).

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

A certain amount of nitrogen trichloride, \( \text{NCl}_3(g) \), is put in a 1.0 L reaction vessel at this temperature. At equilibrium, \( 4.6 \times 10^{-4} \) mol of \( \text{N}_2(g) \) is present. What amount of \( \text{NCl}_3(g) \) was placed in the reaction vessel?

What Is Required?

You need to find the initial amount of \( \text{NCl}_3(g) \) placed in the reaction vessel.

What Is Given?

You are given the chemical equation and the value of \( K_c \). You know the volume of the reaction vessel. You are told the amount of \( \text{N}_2(g) \) at equilibrium.

Plan Your Strategy

Step 1  Calculate the equilibrium molar concentration of \( \text{N}_2(g) \).

Step 2  \( K_c \) is very small. Assume you can ignore the change in concentration of \( \text{NCl}_3(g) \) and set up an ICE table. Let \( x \) represent the change in [\( \text{N}_2 \)].

Step 3  Write the equilibrium expression. Substitute the equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for [\( \text{NCl}_3 \)].

Step 4  Determine the amount of \( \text{NCl}_3(g) \) put in the reaction vessel.
Act on Your Strategy

Step 1  The volume of the reaction vessel is 1.0 L. Therefore, the equilibrium 

\[ [N_2] = 4.6 \times 10^{-4} \text{ mol/L}. \]

Step 2  Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>2NCl_3(g)</th>
<th>⇄</th>
<th>N_2(g)</th>
<th>+</th>
<th>3Cl_2(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>[NCl_3]</td>
<td>0</td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-2x</td>
<td>+x</td>
<td>+3x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>[NCl_3] − 2x ≈ [NCl_3]</td>
<td>4.6 \times 10^{-4}</td>
<td>3x = 3(4.6 \times 10^{-4})</td>
<td></td>
<td>= 1.38 \times 10^{-3}</td>
</tr>
</tbody>
</table>

Step 3  

\[ K_c = \frac{[N_2][Cl_3]^3}{[NCl_3]^2} \]

\[ 3.3 \times 10^{-12} = \frac{(4.6 \times 10^{-4})(1.38 \times 10^{-3})^3}{[NCl_3]^2} \]

\[ [NCl_3] = \sqrt{3.66 \times 10^{-1}} \]

\[ = 6.1 \times 10^{-1} \text{ mol/L} \]

Step 4  The reaction vessel has a volume of 1.0 L. The amount of NCl_3(g) put in the vessel must have been 6.1 \times 10^{-1} \text{ mol}.

Check Your Solution

First, check your assumption that \( x \) is negligible compared with the initial concentration of NCl_3(g). Using the rules for subtracting measured quantities, 

\[ 6.1 \times 10^{-1} - 2(4.6 \times 10^{-3}) = 6.1 \times 10^{-1} \].

Next, check the equilibrium values:

\[ \frac{(4.6 \times 10^{-4})(1.4 \times 10^{-3})^3}{(6.1 \times 10^{-1})^2} = 3.4 \times 10^{-12} \]

This is equal to the equilibrium constant within mathematical rounding error.

24. Problem

At a certain temperature, the value of \( K_c \) for the following reaction is \( 4.2 \times 10^{-8} \).

\[ N_2(g) + O_2(g) ⇄ 2NO(g) \]

0.45 mol of N_2(g) and 0.26 mol of O_2(g) are put in a 6.0 L reaction vessel. What is the equilibrium concentration of NO(g) at this temperature?

What Is Required?

You need to find the concentration of NO(g) at equilibrium.

What Is Given?

You are given the chemical equation. You know the value of \( K_c \), and the initial amounts of N_2(g) and O_2(g). The volume of the container is given.

Plan Your Strategy

Step 1  Calculate the initial molar concentrations of N_2(g) and O_2(g).

Step 2  Divide the smallest initial concentration by \( K_c \) to determine whether you can ignore the change in concentration.

Step 3  Set up an ICE table. Let \( x \) represent the change in [N_2(g)].

Step 4  Write the equilibrium expression. Substitute the equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for \( x \).

Step 5  Calculate [NO] at equilibrium.
Act on Your Strategy

Step 1

\[ [N_2] = \frac{0.45 \text{ mol}}{6.0 \text{ L}} = 0.075 \text{ mol/L} \]

\[ [O_2] = \frac{0.26 \text{ mol}}{6.0 \text{ L}} = 0.043 \text{ mol/L} \]

Step 2

The smallest initial concentration is \([O_2]\).

\[
\begin{align*}
\text{initial concentration} & = \frac{0.043}{4.2 \times 10^{-8}} \\
& = 1.0 \times 10^6
\end{align*}
\]

Because this is well above 500, the changes in \([O_2]\) and \([N_2]\) can be ignored.

Step 3

Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>N_2(g)</th>
<th>+</th>
<th>O_2(g)</th>
<th>⇌</th>
<th>2NO(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.075</td>
<td>0.043</td>
<td></td>
<td>0</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>−x</td>
<td>−x</td>
<td></td>
<td>+2x</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.075−x ≈ 0.075</td>
<td>0.043 −x ≈ 0.043</td>
<td>2x</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Step 4

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

\[
4.2 \times 10^{-8} = \frac{(2x)^2}{(0.075)(0.043)}
\]

\[
x = \sqrt{3.39 \times 10^{-11}}
\]

\[
x = 5.8 \times 10^{-6}
\]

Step 5

\[ [NO] = 2 \times (5.8 \times 10^{-6}) = 1.2 \times 10^{-5} \text{ mol/L} \]

Check Your Solution

The assumption that \(x\) is negligible compared with the initial concentration is valid because, using the rules for subtracting measured quantities, \(0.043 - (5.8 \times 10^{-6}) = 0.043\). Next, check the equilibrium values:

\[
\left(1.2 \times 10^{-5}\right)^2 \div (0.075)(0.043) = 4.5 \times 10^{-8}
\]

This is equal to the equilibrium constant, within the error introduced by mathematical rounding.

25. Problem

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

\[ 2CO_{2(g)} \rightleftharpoons 2CO_{(g)} + O_2_{(g)} \]

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 need to find the concentration of \(CO_2\), \(CO\), and \(O_2\) at equilibrium.

What Is Given?

You are given the chemical equation. You know the value of \(K_c\), and the initial amount of \(CO_2\). The volume of the container is given.
Plan Your Strategy

Step 1 Calculate the initial molar concentration of \( \text{CO}_2(g) \).

Step 2 Divide the initial concentration of \( \text{CO}_2(g) \) by \( K_c \) to determine whether you can ignore the change in concentration.

Step 3 Set up an ICE table. Let \( x \) represent the change in \( [\text{O}_2(g)] \).

Step 4 Write the equilibrium expression. Substitute the equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for \( x \).

Step 5 Calculate the concentration of each gas at equilibrium.

Act on Your Strategy

Step 1 \([\text{CO}_2] = \frac{3.0 \text{ mol}}{5.0 \text{ L}} = 0.60 \text{ mol/L} \)

Step 2 \[
\frac{\text{initial concentration}}{K_c} = \frac{0.60}{2.0 \times 10^{-6}} = 3.0 \times 10^5
\]

Because this is well above 500, the change in \( [\text{CO}_2] \) can be ignored.

Step 3 Set up an ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>( 2\text{CO}_2(g) )</th>
<th>( \rightleftharpoons )</th>
<th>( 2\text{CO(g)} )</th>
<th>+</th>
<th>( \text{O}_2(g) )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.60</td>
<td>0</td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>(-2x)</td>
<td>(+2x)</td>
<td>(+x)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>(0.60 - 2x \approx 0.60)</td>
<td>(2x)</td>
<td>(x)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Step 4
\[
K_c = \frac{[\text{CO}]^2[\text{O}_2]}{[\text{CO}_2]^2}
\]
\[
2.0 \times 10^{-6} = \frac{(2x)^2(x)}{0.60}
\]
\[
x = \sqrt{3.0 \times 10^{-7}}
\]
\[
= 6.7 \times 10^{-4}
\]

Step 4 At equilibrium:
\[\text{[CO}_2] = 0.60 \text{ mol/L} \]
\[\text{[CO]} = 2(6.7 \times 10^{-3}) = 1.3 \times 10^{-2} \text{ mol/L} \]
\[\text{[O}_2] = 6.7 \times 10^{-3} \text{ mol/L} \]

Check Your Solution

The assumption that \( x \) is negligible compared with the initial concentration is valid because, using the rules for subtracting measured quantities,
\[0.60 - 2(6.7 \times 10^{-3}) = 0.59 \approx 0.60\] . Next, check the equilibrium values:
\[
\frac{(1.3 \times 10^{-2})^2(6.7 \times 10^{-3})}{0.60} = 1.9 \times 10^{-6}
\]

This is equal to the equilibrium constant, within the error introduced by mathematical rounding.

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26. Problem

The following reaction takes place inside a cylinder with a moveable piston.

\[2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g)\]
At room temperature, the equilibrium concentrations inside the cylinder are 
\[ [\text{NO}_2] = 0.0206 \text{ mol/L} \] and 
\[ [\text{N}_2\text{O}_4] = 0.0724 \text{ mol/L} \].

(a) Calculate the value of \( K_c \).

(b) Calculate the concentration of each gas at the moment that the piston is used to halve the volume of the reacting mixture. Assume that the temperature remains constant.

(c) Determine the value of \( Q_c \) when the volume is halved.

(d) Predict the direction in which the reaction will proceed to re-establish equilibrium.

What Is Required?
Calculate \( K_c \). Find \([\text{NO}_2]\) and \([\text{N}_2\text{O}_4]\) when the volume inside the cylinder is halved and calculate \( Q_c \). Predict the direction the reaction will proceed to re-establish equilibrium.

What Is Given?
You have the chemical equation and the equilibrium concentrations of \( \text{NO}_2(g) \) and \( \text{N}_2\text{O}_4(g) \).

Plan Your Strategy

(a) Write the equilibrium expression. Substitute the equilibrium concentrations into the equilibrium expression and calculate the value of \( K_c \).

(b) Calculate \([\text{NO}_2]\) and \([\text{N}_2\text{O}_4]\) when the volume of the cylinder is halved.

(c) Calculate the value of \( Q_c \).

(d) Compare the value of \( Q_c \) with the value of \( K_c \). Decide whether the system is at equilibrium and, if not, in which direction the reaction will go.

Act on Your Strategy

(a) \[ K_c = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} \] 
\[ = \frac{0.0724}{(0.0206)^2} \] 
\[ = 171 \]

(b) The volume of the cylinder is halved.
Therefore, 
\[ [\text{NO}_2] = \frac{0.0206 \text{ mol/L}}{0.5} = 0.0412 \text{ mol/L} \]
\[ [\text{N}_2\text{O}_4] = \frac{0.0724 \text{ mol/L}}{0.5} = 0.145 \text{ mol/L} \]

(c) \[ Q_c = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} \] 
\[ = \frac{0.1448}{(0.0412)^2} \] 
\[ = 85.3 \]

(d) \( Q_c < K_c \). The system is not at equilibrium and will proceed by moving to the right.

Check Your Solution
Check your calculations of \( K_c \) and \( Q_c \). The value of \( Q_c \) is less than that of \( K_c \), so you should expect the reaction to re-establish equilibrium by forming more product.

27. Problem
Ethyl acetate is an ester that can be synthesized by reacting ethanoic acid (acetic acid) with ethanol. At room temperature, the equilibrium constant for this reaction is 2.2. 
\[ \text{CH}_3\text{COOH}_\ell + \text{CH}_3\text{CH}_2\text{OH}_\ell \rightleftharpoons \text{CH}_3\text{COOCH}_2\text{CH}_3\ell + \text{H}_2\text{O}_\ell \]
Various samples were analyzed. The concentrations are given in the table below. Decide whether each sample is at equilibrium. If it is not at equilibrium, predict the direction the reaction will proceed to establish equilibrium.

<table>
<thead>
<tr>
<th>Sample</th>
<th>[CH$_3$COOH] (mol/L)</th>
<th>[CH$_3$CH$_2$OH] (mol/L)</th>
<th>[CH$_3$COOCH$_2$CH$_3$] (mol/L)</th>
<th>[H$_2$O] (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a)</td>
<td>0.10</td>
<td>0.10</td>
<td>0.10</td>
<td>0.10</td>
</tr>
<tr>
<td>(b)</td>
<td>0.084</td>
<td>0.13</td>
<td>0.16</td>
<td>0.28</td>
</tr>
<tr>
<td>(c)</td>
<td>0.14</td>
<td>0.21</td>
<td>0.33</td>
<td>0.20</td>
</tr>
<tr>
<td>(d)</td>
<td>0.063</td>
<td>0.11</td>
<td>0.15</td>
<td>0.17</td>
</tr>
</tbody>
</table>

**What Is Required?**
You must calculate $Q_c$ for each sample and interpret the value.

**What Is Given?**
You have the chemical equation. You know the value of $K_c$ is 2.2. You also know the concentrations of the reacting substances in each sample.

**Plan Your Strategy**
Write the expression for $Q_c$, and then calculate its value for each sample. Compare the value of $Q_c$ with the value of $K_c$. Decide whether the system is at equilibrium and, if not, in which direction the reaction will go.

**Act on Your Strategy**

$$Q_c = \frac{[\text{CH}_3\text{COOCH}_2\text{CH}_3][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{CH}_3\text{CH}_2\text{OH}]}$$

(a) $Q_c = \frac{(0.10)(0.10)}{(0.10)(0.10)} = 1.0$

$Q_c < K_c$. The reaction proceeds to the right.

(b) $Q_c = \frac{(0.16)(0.28)}{(0.084)(0.13)} = 4.1$

$Q_c > K_c$. The reaction proceeds to the left.

(c) $Q_c = \frac{(0.33)(0.20)}{(0.14)(0.21)} = 2.2$

$Q_c = K_c$. The reaction is at equilibrium.

(d) $Q_c = \frac{(0.15)(0.17)}{(0.063)(0.11)} = 3.7$

$Q_c > K_c$. The reaction proceeds to the left.

**Check Your Solution**
Check your calculations of $Q_c$. When the value of $Q_c$ is less than that of $K_c$, the reaction proceeds to the right because more product must be formed to reach equilibrium.

**28. Problem**
In the past, methanol was obtained by heating wood without allowing the wood to burn. The products were collected, and methanol (sometimes called “wood alcohol”) was separated by distillation. Today methanol is manufactured by reacting carbon monoxide with hydrogen gas.

$$\text{CO(g)} + 2\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_3\text{OH(g)}$$

At 210°C, $K_c$ for this reaction is 14.5. A gaseous mixture at 210°C contains the following concentrations of gases: [CO] = 0.25 mol/L, [H$_2$] = 0.15 mol/L, and [CH$_3$OH] = 0.36 mol/L. What will be the direction of the reaction if the gaseous mixture reaches equilibrium?
What Is Required?
You must calculate $Q_c$ and interpret its value.

What Is Given?
You have the chemical equation. You know the value of $K_c$ is 14.5. You also know the concentrations of the gases: $[\text{CO}] = 0.25 \text{ mol/L}$, $[\text{H}_2] = 0.15 \text{ mol/L}$, and $[\text{CH}_3\text{OH}] = 0.36 \text{ mol/L}$.

Plan Your Strategy
Write the expression for $Q_c$, and then calculate its value. Compare the value of $Q_c$ with the value of $K_c$. Decide whether the system is at equilibrium and, if not, in which direction the reaction will go.

Act on Your Strategy
$$Q_c = \frac{[\text{CH}_3\text{OH}][\text{CO}][\text{H}_2]^2}{2}$$
$$= \frac{0.36}{(0.25)(0.15)^2}$$
$$= 64$$

$Q_c > K_c$ The system is not at equilibrium. The reaction proceeds to the left.

Check Your Solution
Check your calculations of $Q_c$. When the value of $Q_c$ is greater than that of $K_c$, the reaction proceeds to the left because less product must be formed to reach equilibrium.

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29. Problem
Consider the following reaction.
$$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) + 52 \text{ kJ} \rightleftharpoons 2\text{HI}(\text{g})$$
In which direction does the equilibrium shift if there is an increase in temperature?

What Is Required?
You need to determine whether increasing the temperature causes the equilibrium to shift to the left or to the right.

What Is Given?
You have the chemical equation. You know that heat is shown on the left of the equation. Therefore, the reaction is endothermic.

Plan Your Strategy
Use the chemical equation to determine the shift in equilibrium that will minimize the change.

Act on Your Strategy
The temperature is increased. Therefore, the equilibrium must shift in the direction in which the reaction is endothermic. From left to right, the reaction is endothermic. The reaction shifts to the right if the temperature is increased.

Check Your Solution
An increase in temperature must result in a shift in the reaction that minimizes the temperature increase. The shift must be in the direction in which the reaction is endothermic.
30. Problem
A decrease in the pressure of each equilibrium system below is caused by increasing
the volume of the reaction chamber. In which direction does the equilibrium shift?
(a) \( \text{CO}_2(g) + \text{H}_2(g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2\text{O}(g) \)
(b) \( 2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g) \)
(c) \( 2\text{CO}_2(g) \rightleftharpoons 2\text{CO}_2(g) + \text{O}_2(g) \)
(d) \( \text{CH}_4(g) + 2\text{H}_2\text{S}(g) \rightleftharpoons \text{CS}_2(g) + 4\text{H}_2(g) \)

What Is Required?
You need to determine whether decreasing the pressure causes the equilibrium to shift
to the left or the right, or whether it has no effect.

What Is Given?
You have each chemical equation.

Plan Your Strategy
Use the chemical equation to determine the shift in equilibrium that will minimize
the 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.
    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
An increase in volume must result in a decrease in pressure. 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.

31. Problem
The following reaction is exothermic.
\( 2\text{NO}(g) + 2\text{H}_2(g) \rightleftharpoons \text{N}_2(g) + 2\text{H}_2\text{O}(g) \)
In which direction does the equilibrium shift as a result of each change?
(a) removing the 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
or the right, or whether it has no effect.

What Is Given?
You have the chemical equation. You know the reaction is exothermic.

Plan Your Strategy
Identify the change. Then use the chemical equation to determine the shift in
equilibrium that will minimize the change.
Act on Your Strategy
(a) \([H_2]\) is reduced. Therefore, the equilibrium must shift to increase \([H_2]\).
The reaction shifts to the left.
(b) The pressure is increased by decreasing the volume of the reaction vessel.
Therefore, 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. The position of equilibrium does not change.
(d) The temperature increases. Therefore, the equilibrium must shift in the direction in which the reaction is endothermic. From left to right, the reaction is exothermic. Therefore, 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.

Check Your Solution
Check the changes. Any change that affects the equilibrium reaction must result in a shift that minimizes it.

32. Problem
In question 31, 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?
You have the chemical equation. You know the reaction is exothermic.

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 \(K_c\).
Therefore, 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.

33. Problem
Toluene, \(C_7H_8\), is an important organic solvent. It is made industrially from methyl cyclohexane.
\[
C_7H_{14(g)} \rightleftharpoons C_7H_8(g) + 3H_2(g)
\]
The forward reaction is endothermic. State three different changes to an equilibrium mixture of these reacting gases that would shift the equilibrium 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?
You have the chemical equation. You know the reaction is endothermic.

Plan Your Strategy
Each change must shift the reaction to the right. Use the chemical equation to identify changes that will shift the equilibrium to the right as the change is minimized.
Act on Your Strategy
- The equilibrium will shift to the right if any chemical on the left is added. Therefore, increasing \([C_7H_{14}]\) increases the production of toluene.
- The equilibrium will shift to the right if any chemical on the right is removed. Therefore, decreasing either \([C_7H_8]\) or \([H_2]\), or both, increases the production of toluene.
- The number of gas molecules increases as the reaction proceeds from left to right. Therefore, decreasing the pressure of the gases in the reaction vessel increases the production of toluene.
- The reaction is endothermic from left to right. Therefore, increasing the temperature increases the production of toluene.

Check Your Solution
Check each change. Minimizing each change must result in a shift in the reaction to the right.