2. Equilibrium

**Equilibrium constant Kc**

For a generalised reaction

\[ mA + nB \rightleftharpoons pC + qD \]

\( m, n, p, q \) are the stoichiometric balancing numbers

\( A, B, C, D \) stand for the chemical formula

\[ \frac{[C]^{p} [D]^{q}}{[A]^{m} [B]^{n}} \]

\( K_c = \text{equilibrium constant} \)

**Example 1**

\[ N_2 (g) + 3H_2 (g) \rightleftharpoons 2NH_3 (g) \]

\[ K_c = \frac{[NH_3 (g)]^2}{[N_2 (g)][H_2 (g)]^3} \]

The unit of Kc changes and depends on the equation.

**Working out the unit of Kc**

Put the unit of concentration (mol dm\(^{-3}\)) into the Kc equation

\[ K_c = \frac{[NH_3 (g)]^2}{[N_2 (g)][H_2 (g)]^3} \]

**Example 2:** writing Kc expression

\[ H_2 (g) + Cl_2 (g) \rightleftharpoons 2HCl (g) \]

**Calculating Kc**

Most questions first involve having to work out the equilibrium moles and then concentrations of the reactants and products. Usually the question will give the initial amounts (moles) of the reactants, and some data that will help you work out the equilibrium amounts.

**Calculating the moles at equilibrium**

moles of reactant at equilibrium = initial moles – moles reacted

moles of product at equilibrium = initial moles + moles formed

**Example 3**

For the following equilibrium

\[ H_2 (g) + Cl_2 (g) \rightleftharpoons 2HCl (g) \]

In a container of volume 600cm\(^3\) there were initially 0.5mol of \( H_2 \) and 0.6 mol of \( Cl_2 \).

At equilibrium there were 0.2 moles of \( HCl \). Calculate Kc

<table>
<thead>
<tr>
<th>( H_2 )</th>
<th>( Cl_2 )</th>
<th>( HCl )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial moles</td>
<td>0.5</td>
<td>0.6</td>
</tr>
<tr>
<td>Equilibrium moles</td>
<td>0.4</td>
<td>0.5</td>
</tr>
</tbody>
</table>

It is often useful to put the mole data in a table.

Using the balanced equation if 0.2 moles of \( HCl \) has been formed it must have used up 0.1 of \( Cl_2 \) and 0.1 moles of \( H_2 \) (as 1:2 ratio)

\[ \text{moles of hydrogen at equilibrium} = 0.5 - 0.1 = 0.4 \]

\[ \text{moles of chlorine at equilibrium} = 0.6 - 0.1 = 0.5 \]

If the Kc has no unit then there are equal numbers of reactants and products. In this case you do not have to divide by volume to work out concentration and equilibrium moles could be put straight into the Kc expression

\[ K_c = \frac{[HCl (g)]^2}{[H_2 (g)][Cl_2 (g)]} \]

\[ K_c = \frac{0.33^2}{0.67 \times 0.83} = 0.196 \text{ no unit} \]
Example 4

For the following equilibrium

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

Initially there were 1.5 moles of \( \text{N}_2 \) and 4 mole of \( \text{H}_2 \) in a 1.5 dm\(^3\) container. At equilibrium 30% of the Nitrogen had reacted. Calculate \( K_c \)

<table>
<thead>
<tr>
<th></th>
<th>( \text{N}_2 )</th>
<th>( \text{H}_2 )</th>
<th>( \text{NH}_3 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.5</td>
<td>4.0</td>
<td>0</td>
</tr>
<tr>
<td>Equilibrium</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[
30\% \text{ of the nitrogen reacted} = 0.3 \times 1.5 = 0.45 \text{ moles reacted.}
\]

Using the balanced equation \( 3 \times 0.45 \) moles of \( \text{H}_2 \) must have reacted and \( 2 \times 0.45 \) moles of \( \text{NH}_3 \) must have formed

Work out the moles at equilibrium for the reactants and products

\[
\text{moles of reactant at equilibrium} = \text{initial moles} - \text{moles reacted}
\]

\[
\text{moles of nitrogen at equilibrium} = 1.5 - 0.45 = 1.05 \quad \text{moles of hydrogen at equilibrium} = 4.0 - 0.45 \times 3 = 2.65
\]

\[
\text{moles of ammonia at equilibrium} = 0 + (0.45 \times 2) = 0.9
\]

<table>
<thead>
<tr>
<th></th>
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</table>

Finally put concentrations into \( K_c \) expression

\[
K_c = \frac{[\text{NH}_3 (g)]^2}{[\text{N}_2 (g)] [\text{H}_2 (g)]^3}
\]

\[
K_c = \frac{0.6^2}{0.7 \times 1.77^3} = 0.0927 \text{ mol}^{-2} \text{ dm}^{+6}
\]

Effect of changing conditions on value of \( K_c \)

The larger the \( K_c \) the greater the amount of products. If \( K_c \) is small we say the equilibrium favours the reactants

Effect of Temperature on position of equilibrium and \( K_c \)

Both the position of equilibrium and the value of \( K_c \) will change if temperature is altered

In this equilibrium which is exothermic in the forward direction

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

If temperature is increased the reaction will shift to oppose the change and move in the backwards endothermic direction. The position of equilibrium shifts left. The value of \( K_c \) gets smaller as there are fewer products.

Effect of Pressure on position of equilibrium and \( K_c \)

The position of equilibrium will change if pressure is altered but the value of \( Kc \) stays constant as \( K_c \) only varies with temperature

In this equilibrium which has fewer moles of gas on the product side

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

If pressure is increased the reaction will shift to oppose the change and move in the forward direction to the side with fewer moles of gas. The position of equilibrium shifts right. The value of \( K_c \) stays the same though as only temperature changes the value of \( K_c \).

Catalysts have no effect on the value of \( K_c \) or the position of equilibrium as they speed up both forward and backward rates by the same amount.