\(K_p\) is the equilibrium constant calculated from the partial pressures of a reaction equation. It is used to express the relationship between product pressures and reactant pressures. It is a unitless number, although it relates the pressures.

**Introduction**

In calculating \(K_p\), the partial pressures of gases are used. The partial pressures of pure solids and liquids are not included. To use this equation, it is beneficial to have an understanding of partial pressures and mole fractions.

**Partial Pressures:** All of the partial pressures add up to the total pressure, as shown in the equation (Dalton's law)

\[
P_{\text{total}} = P_A + P_B + \ldots \tag{1}
\]

Individual gases maintain their respective pressures when combined. However, their individual pressures add up to the total pressure in the system.

**Mole Fractions:** Although mole fractions can be used for non-gases, this description is referring specifically to those involving gases. This fraction expresses the moles of an individual gas compared to the total moles. It is used for reactions in which more than one gas is involved. In general the equation for finding the mole fraction of a gas (represented by A) is

\[
x_A = \frac{\text{mol}_A}{\text{mol}_{\text{total}}} \tag{2}
\]

Note

If the mole fraction is multiplied by 100, it becomes a percent expressing the mole amount for the individual gas.

When combined with the total pressure, this will equal the partial pressure (for gas A). The equation is thus

\[
P_A = x_A \times P_{\text{total}} \tag{3}
\]

**Calculations**

Partial pressures are used for calculating the pressure constant. Consider this general equation for a gas-phase reaction:

\[aA + bB \rightleftharpoons cC + dD \tag{4}\]

The formula uses the partial pressures, \(P_A\), \(P_B\), \(P_C\), and \(P_D\), raised to exponents equal to their respective coefficients in the chemical equation. \(K_p\) is calculated as follows:

\[
K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b} \tag{5}
\]

Note

This expression is similar to \((K_c)\); however, \((K_c)\) is calculated from molar concentrations. To prevent confusion, do not use brackets around partial pressures.
\(K_p\) can also be obtained from \(K_c\) (the concentration equilibrium constant), temperature, change in moles, and the gas constant, as shown:

\[
K_p = K_c (RT)^{\Delta n}
\]  \(\text{(6)}\)

Calculating \(K_c\) is not discussed in this article; this equation simply relates the two. For more information see the page about \(K_c\).

References


Practice Problems

1. Give the equilibrium constant expression for the generation of water from oxygen and hydrogen gas:

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

2. What is \(K_p\) given that the initial pressure of \((\text{SO}_3)\) is 3 atm (and zero for the other species) for this reversible reaction:

\[
(2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3)
\]

3. What is \(K_p\) for this reaction given the partial pressures at equilibrium are \((P_{\text{H}_2} = 0.9968\text{ atm})\), and \((P_{\text{O}_2} = 1.105\text{ atm})\):

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

4. The mole fractions of a closed vessel of the following gases are: \(\text{Cl}_2: 0.243\), \(\text{O}_2: 0.274\), and \(\text{Cl}_2\text{O}_5:0.483\). If the following equilibrium is established with a total pressure is 3 atm. What is \(K_p\)?

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

5. Find \(K_p\) for this reaction

\[
2\text{N}_2\text{O}_5(g) \rightleftharpoons \text{O}_2(g) + 4\text{NO}_2(g)
\]

N: 14.01 g/mol, O: 16 g/mol, total pressure: 4 atm.

Answers

1. For this basic equation, all of the products and reactants are gases. The equation will be:
2. Set up an ICE table. If \( x \) is the final pressure of \( O_2 \).

<table>
<thead>
<tr>
<th>ICE</th>
<th>(2SO_2)</th>
<th>(O_2)</th>
<th>(2SO_3)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0</td>
<td>0</td>
<td>3</td>
</tr>
<tr>
<td>Change</td>
<td>+2(x)</td>
<td>+(x)</td>
<td>-2(x)</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>2(x)</td>
<td>(x)</td>
<td>3-2(x)</td>
</tr>
</tbody>
</table>

so \( K_p \) will be

\[
K_p = \frac{(3-2x)^2}{(2x)^2(x)} = \frac{9 - 12x + 4x^2}{4x^3}
\]

Solving for \( K_p \) requires knowledge of the equilibrium pressure of any of the three species in this reaction.

3. Hint: At first this equation may look identical to Problem 1, but look at the equation again. This time the \( H_2O \) is a liquid this time and not a gas. Therefore, the liquid water is not included in the equation. The equation is

\[
K_p = \frac{1}{0.9968^2 \times 1.105} = 0.9108
\]

4. First find the partial pressures

\[
P_{Cl_2}: 0.243 \times 3 \text{ atm} = 0.729 \text{ atm}
\]
\[
P_{O_2}: 0.274 \times 3 \text{ atm} = 0.822 \text{ atm}
\]
\[
P_{Cl_2O_5}: 0.483 \times 3 \text{ atm} = 1.449 \text{ atm}
\]

Now construct the expression of \( K_p \) for this specific reaction

\[
K_p = \frac{(P_{(Cl_2)})^2 \times (P_{(O_2)})^5}{(P_{(Cl_2O_5)})^2}
\]

\[
K_p = \frac{(0.729)^2 \times (0.822)^5}{(1.449)^2} = 0.095
\]

5. First find the mole fractions:

\[
N_{2O_5} = 2(14.01 \times 2 + 16 \times 5) = 216.04 \text{ mol}
\]
\[
O_2 = 16 \times 2 = 32 \text{ mol}
\]
\[
NO_2 = 4(14.01 + 16 \times 2) = 184.04 \text{ mol}
\]

mole fractions- \( N_2O_5 \): 0.5, \( O_2 \): 0.0741, \( NO_2 \): 0.425

Next find partial pressures:
\(N_2O_5 = 0.5 \times 4 \text{ atm} = 2 \text{ atm}\)

\(O_2 = 0.0741 \times 4 \text{ atm} = 0.2964 \text{ atm}\)

\(NO_2 = 0.425 \times 4 \text{ atm} = 1.7 \text{ atm}\)

Last, find \(K_p\):

\[
K_p = \frac{(0.2964)(1.7)^4}{(2)^2} = 0.619
\]

**Contributors**

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