$K_c$ and $K_p$ are the equilibrium constants of gaseous mixtures. However, the difference between the two constants is that $K_c$ is defined by molar concentrations, whereas $K_p$ is defined by the partial pressures of the gasses inside a closed system. The equilibrium constants do not include the concentrations of single components such as liquids and solid, and they do not have any units.

### How To Write Gas Equilibrium Constants

Here are some easy steps on writing Gas Equilibrium Constants (this is the same for finding $K_c$, $K_p$, $K_{sp}$, $Q$ and etc.):

1. In equilibrium equations, even though the arrows point both ways ($\rightleftharpoons$) we usually associate the left as reactants and the right as products.
2. The products are on the TOP of the fraction (the numerator).
3. The reactants are on the BOTTOM of the fraction (the denominator).
4. The concentrations of the products and reactants are always raised to the power of their coefficient in the balanced chemical equation.
5. If any of the reactants or products are solids or liquids, their concentrations are equal to one because they are pure substances. (see Solubility Product Constant, $K_{sp}$)

The standard example of writing Gas Equilibrium Constants are:

\[
\begin{align*}
\text{aA + bB} & \rightleftharpoons \text{cC + dD} \\
K_c &= \dfrac{[C]^c[D]^d}{[A]^a[B]^b} \\
K_p &= \dfrac{(C)^c(D)^d}{(A)^a(B)^b}
\end{align*}
\]

**Example 1: Thermal Decomposition of $NH_4SH_{(s)}$**

Consider the thermal decomposition of $NH_4SH_{(s)}$:

\[
NH_4SH_{(s)} \rightleftharpoons NH_3(g) + H_2S(g)
\]

This also is related to $K_{sp}$

\[
K_c = \dfrac{[NH_3][H_2S]}{[NH_4SH]}
\]

but since $NH_4SH_{(s)}$ is a solid, we get:

\[
K_c = [NH_3][H_2S]
\]

As for $K_p$, it is the same as $K_c$, but instead of brackets [ ], $K_p$ uses parentheses ( ):

\[
K_p = (NH_3)(H_2S)
\]
Example 2: Hydrogen and Iodine

Consider the double replacement reaction of hydrogen and iodine gas:

\[
H_2 (g) + I_2 (g) \rightleftharpoons 2HI (g)
\]

\[
K_c = \frac{[HI]^2}{[H_2][I_2]}
\]

\[
K_p = \frac{(HI)^2}{(H_2)(I_2)}
\]

---

**Definition of \((K_c)\) and \((K_p)\)**

\((K_c)\) is an equilibrium constant in terms of molar **concentrations** and is usually defined as:

\[
K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}
\]

in the general reaction,

\[
aA + bB \rightleftharpoons cC + dD
\]

If a large \((K_c)\) is formed then there are more products formed. Inversely, a small \((K_c)\) indicates that the reaction favors the reactants.

\((K_p)\) is an equilibrium constant in terms of **partial pressures** and is usually defined as:

\[
K_p = \frac{(C)^c(D)^d}{(A)^a(B)^b}
\]

for the general reaction

\[
aA + bB \rightleftharpoons cC + dD
\]

**Homogeneous Equilibria**: Reactants/Products all in a single phase. For example:

\[
A_-(g) + B_-(g) \rightleftharpoons C_-(g) + D_-(g)
\]

**Heterogeneous Equilibria**: Reactants/Products in more than one phase. For example:

\[
A_-(s) + B_-(g) \rightleftharpoons C_-(g) + D_-(s)
\]
The relationship between the two equilibrium constants are:

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

or

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

where,

- \( \Delta n \) = (Total moles of gas on the products side) - (Total moles of gas on the reactants side). Hence \( \Delta n = (d + c) - (a + b) \)
- \( R \) is the gas constant found in the ideal gas law \((0.0821 \ \text{mole/Liter/degree Kelvin})\)

Relating Gas Equilibrium Constants to Equilibrium (K)

The value of K depends on whether the solution being calculated for is using concentrations or partial pressures. The gas equilibrium constants relate to the equilibrium (K) because they are both derived from the ideal gas law \((PV = nRT)\).

- \( K_c \) is the concentration of the reaction, it is usually shown as:
  \[ \frac{[C][D]}{[A][B]} \]

- \( K_p \) is the amount of partial pressure in the reaction, usually shown as:
  \[ \frac{p(C)p(D)}{p(A)p(B)} \]

As we have seen above, \( K_p = K_c (RT)^{\Delta n} \), we can derive this formula from the Ideal Gas Law.

We know that \( K_c \) is in terms Molarity \((\text{mole/Liter})\), and we can also arrange the Ideal Gas Law \((PV = nRT)\) as: \( \frac{n}{L} = \frac{P}{RT} \). We can replace \( K_c \) with Molarity: the equation become, \( K_p = K_c (RT)^{\Delta n} \)

We know that Partial Pressure is directly proportional to Concentration:

\( P = \frac{n}{L} \) Pressure can be in units of: Pascal (Pa), Atmosphere (atm), or Torr.

Therefore we can replace \( K_c \) with Molarity: the equation become, \( K_p = K_c (RT)^{\Delta n} \)

\( (RT)^{\Delta n} = \frac{[C][D]}{[A][B]} \)

Also: \( \frac{n}{L} = \frac{P}{RT} \), can be shown as \( K_c = K_p (RT) \)

K is also written the same as \( K_c \) and \( K_p \):

\( aA + bB \rightleftharpoons cC + dD \)

\( K = \frac{[C]^c[D]^d}{[A]^a[B]^b} \)
Example 3

\[ 2 \text{NOBr}_{(g)} \rightleftharpoons 2 \text{NO}_{(s)} + \text{Br}_2 (g) \]

Given: NOBr = 0.46 M
NO = 0.1 M
Br₂ = 0.3 M

To set up \(K_c\), it is \(\dfrac{\text{Products}}{\text{Reactants}}\)

\[ K_c = \dfrac{[\text{NO}]^2 \cdot [\text{Br}_2]}{[\text{NOBr}]^2} \]

\[ K_c = \dfrac{[0.1]^2 \cdot [0.3]}{[0.46]^2} \]

Answer: \(K_c = 0.0142\) M

Example 4

Examples of \(K_p\)

This is an example of \(K_p\) done by Dr. Enderle on youtube.
Example 5

N₂O₄ (l) is an important component of rocket fuel. At 25 °C, N₂O₄ is a colorless gas that partially dissociates into NO₂. The color of an equilibrium mixture of these 2 gases depends on their relative proportions, which are dependent on temperature. Equilibrium is established in the reaction \( \text{N}_2\text{O}_4 (g) \rightleftharpoons 2\text{NO}_2 (g) \) at 25 °C.

Given: 3.00 L container
7.64 g N₂O₄
1.56 g NO₂

What is the \( K_c \) for this reaction?

Step 1: Convert grams to moles

\[
mol \text{N}_2\text{O}_4 = 7.64 \text{ g} \times \left( \frac{1 \text{ mol N}_2\text{O}_4}{92.01 \text{ g}} \right) = 8.303 \times 10^{-2} \text{ mol}
\]
\[
mol \text{NO}_2 = 1.56 \text{ g} \times \left( \frac{1 \text{ mol NO}_2}{46.01 \text{ g}} \right) = 3.391 \times 10^{-2} \text{ mol}
\]

Step 2: Convert moles to Molarity (moles/L)

\[
[N_2O_4] = \frac{8.303 \times 10^{-2} \text{ mol N}_2\text{O}_4}{3.00 \text{ L}} = 0.0277 \text{ M}
\]
\[
[NO_2] = \frac{3.391 \times 10^{-2} \text{ mol NO}_2}{3.00 \text{ L}} = 0.0113 \text{ M}
\]

Step 3: Write the Equilibrium constant for \( K_c \):

\[
K_c = \left( [\text{NO}_2] \right)^2 / [\text{N}_2\text{O}_4] = \left( 0.0113 \right)^2 / 0.0277 = 4.61 \times 10^{-3}
\]

Given: N₂ = 0.79 moles
O₂ = 0.21 moles
Temp = 2500 K

When equilibrium is established the mole percent of Nitrogen Oxide (NO) at 1.8%

Calculate \( K_p \) for the reaction

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

1st step: Create an ICE table:

\[
\begin{align*}
\text{N}_2(g) & + \text{O}_2(g) \rightleftharpoons 2\text{NO}_2(g)
\end{align*}
\]
<table>
<thead>
<tr>
<th>ICE</th>
<th>( \text{(N_2)} )</th>
<th>( \text{(O_2)} )</th>
<th>( \text{(\text{NO})} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.79 mol</td>
<td>0.21 mol</td>
<td>0 mol</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>-x</td>
<td>+2x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>(0.79 - x)</td>
<td>(0.21 - x)</td>
<td>(2x + 0)</td>
</tr>
</tbody>
</table>

2nd Step: Find the Mole Percent (%) of products

\[
X_{\text{NO}} = 0.018
\]

\[
X_{\text{total}} = (0.79-x) + (0.21-x) + (2x) = 1
\]

\[
X_{\text{NO}} = \frac{2x}{X_{\text{total}}}
\]

\[
0.018 = \frac{2x}{1}
\]

\[
x = 0.009
\]

3rd Step: Set up an equation

\[
K_p = \frac{p(\text{\(\text{NO}\)})^2}{p(\text{\(\text{N}_2\)})p(\text{\(\text{O}_2\)})}
\]

\[
PV = NRT
\]

\[
\frac{\left[ n(\text{\(\text{NO}\)})(RT) \right]^2}{V^2} \frac{\left[ n(\text{\(\text{N}_2\)})(RT) \right]}{V} \frac{n(\text{\(\text{O}_2\)})(RT)}{V}
\]

[Volume Cancels out]

\[
\frac{\left[ n(\text{\(\text{NO}\)})(RT) \right]^2}{n(\text{\(\text{N}_2\)})n(\text{\(\text{O}_2\)})}
\]

4th Step: Plug in values

\[
\left[ K_p \right] = \frac{(2x)^2}{(0.79-x)(0.21-x)}
\]

\[
x = 0.009
\]

\[
K_p = 2.1 \times 10^{-3}
\]

Template:ExampleEnd

How the Gas Equilibrium Constants Relate to Reaction Quotient (Q)

The process of finding the Reaction Quotient (Q_c) is the same as finding K_c and K_p, where the products of the reaction is divided by the reactants of the reaction \(\frac{\left(\frac{\text{products}}{\text{reactants}}\right)}{\text{at any time not necessarily at}}\)
If a problem asks you to find which way the reaction will shift in order to achieve equilibrium, and K is given, you would have to calculate for Q and compare the two numbers.

When comparing K and Q:

- **K < Q**: Since there are more products than reactants, the reaction will produce more reactants to reach equilibrium, the reaction favors the reactants.
- **K > Q**: Since there are more reactants than products, the reaction will produce more products to reach equilibrium, the reaction favors the products.
- **K = Q**: There is no change in the products nor reactants, so equilibrium is achieved.

A trick to remember to which the reaction will favor is:
Put:
K \_ Q (in alphabetical order! - or it will not work)

- **K < Q**: K \(\leftarrow\) Q
  The reaction will favor the reactants because reactants are on the left of the equation.

- **K > Q**: K \(\rightarrow\) Q
  The reaction will favor the products because products are on the right of the equation.

- **K = Q**: NO CHANGE

(See Relationship **Between K and Q** for more information)

**Example 7: Relating K to Q**
\[
\text{CO (g) + H}_2\text{O (g) } \rightleftharpoons \text{CO}_2 (g) + \text{H}_2 (g)
\]

Given: \(K_c = 1.00\) at about 1100 K

CO = 1.00 mol H\(_2\)O = 1.00 mol CO\(_2\) = 2.00 mol H\(_2\) = 2.00 mol

Compared with their initial amounts, which of the substances will be present in a greater amount and which is in a lesser amount when equilibrium is established?

**Step 1: Write out the expression for \(Q_c\)**

\[
Q_c = \dfrac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]} \]

**Step 2: Plug in the number of Molarity, since volume is not given, assume it is 1 Liters**

\[
Q_c = \dfrac{[2.00][2.00]}{[1.00][1.00]} = 4.00
\]

**Step 3: Compare \(K_c\) with \(Q_c\)**

\[K_c = 1.00\) (unitless) \(Q_c = 4.00\) (unitless)\]

\[K_c < Q_c\]

\[1.00 < 4.00\]

Therefore, the reaction will shift to the LEFT towards the reactants.

**Example 8**

A mixture of hydrogen, iodine, and hydrogen iodide, each at 0.0020 M, was introduced into a container heated to 783 Kelvins. At this temperature \((K_c = 46)\), Predict if more HI or less will be formed.

**Step 1: Write out the reaction**

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

**Step 2: Write out the expression for \(Q_c\)**

\[
Q_c = \dfrac{[\text{HI}]^2}{[\text{I}_2][\text{H}_2]} \]

**Step 3: Plug in the Molarity given**
Molarity = 0.0020 M

\[ Q_c = \frac{(0.0020)^2}{0.0020 \cdot 0.0020} \]
\[ Q_c = 1.00 \]

**Step 4: Compare \( K_c \) with \( Q_c \)**

\[ K_c = 46 \text{ (unitless)} \]
\[ Q_c = 1.00 \text{ (unitless)} \]

\[ K_c > Q_c \]
\[ 46 > 1.00 \]

The reaction will shift to the RIGHT towards the products. Therefore, more HI will be produced.

---

**Activity - A Brief Overview**

When calculating equilibrium constants, solids and liquids are mostly omitted because they are considered pure substances. Pure substances do not have a change in concentration. Though pure substances are omitted in equilibrium constants, they are still necessary components of the chemical equation (see Activity).

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**References**


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**Outside Links**

3. [http://www.youtube.com/watch?v=pnZpn9N8XbU](http://www.youtube.com/watch?v=pnZpn9N8XbU) (The youtube video about: \( K_c \) vs. \( K_p \))
4. [http://www.youtube.com/watch?v=W7NyGbI2Q1Q&feature=related](http://www.youtube.com/watch?v=W7NyGbI2Q1Q&feature=related) (Dr. Enderle’s example on \( K_p \))
5. http://library.thinkquest.org/C00666...brium/gas.html (On how to write equilibrium constants)
7. http://www.tutorvista.com/content/ch...n-quotient.php
9. http://www.chemguide.co.uk/physical/equilibria/kp.html#top (more information about K_p)
10. http://www.chemguide.co.uk/physical/equilibria/kc.html (more information about K_c)

Problems

Below are practice problems for K_c and K_p:

K_c

1. Gaseous Hydrogen Iodide is placed in a closed container at 425 °C, Where it partially decomposes to Hydrogen and Iodine:

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

   The following are given:

   \[ [\text{HI}] = 3.53 \times 10^{-3} \text{ M} \]
   \[ [\text{H}_2] = 4.79 \times 10^{-4} \text{ M} \]
   \[ [\text{I}_2] = 4.79 \times 10^{-4} \text{ M} \]

   What is the value of K_c at this temperature?

2. Write the K_c for the reaction and state where the reaction is Homogeneous or Heterogeneous.
   a) \( \text{N}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2\text{NO (g)} \)
   b) \( \text{FeO (s)} + \text{H}_2 (g) \rightleftharpoons \text{Fe (s)} + \text{H}_2\text{O (g)} \)

3. Determine values of K_c from the K_p value given: (number 7 from p. 655 in the textbook)

   \( 2\text{NO (g)} + \text{O}_2 (g) \rightleftharpoons 2\text{NO}_2 (g); \ K_p = 1.48 \times 10^4 \) at 184 °C

K_p

1. Write the K_p for the reaction and state where the reaction is Homogeneous or Heterogeneous.
   a) \( 2\text{C}_2\text{H}_4 (g) + 2\text{H}_2\text{O (g)} \rightleftharpoons 2\text{C}_2\text{H}_6 (g) + \text{O}_2 (g) \)
   b) \( \text{Ti (s)} + 2\text{Cl}_2 (g) \rightleftharpoons \text{TiCl}_4 (g) \)

2. Determine values of K_p from the K_c value given: (number 8 from p. 655 in the textbook)
2H₂S (g) + CH₄ (g) ⇌ 4H₂ (g) + CS₂ (g); K_c = 5.27 * 10⁻⁸ at 973 K.

3. The two common chlorides of Phosphorus, PCl₃ and PCl₅, both important in the production of other phosphorous compounds, coexist in equilibrium through: (number 17 from p. 655 in the textbook)

PCl₃ (g) + Cl₂ (g) ⇌ PCl₅ (g)

At 250 °C, an equilibrium mixture in a 2.50 L flask contains 0.105 g PCl₅, 0.220 g PCl₃, and 2.12 g Cl₂. What are the values of (a) Kc and (b) Kp for this reaction?

**Answers**

**K_c**

1. 2HI (g) ⇌ H₂ (g) + I₂ (g)

   [HI] = 3.53 * 10⁻³ M [H₂] = 4.79 * 10⁻⁴ M [I₂] = 4.79 * 10⁻⁴ M

   \[ K_c = \frac{[H_2][I_2]}{[HI]^2} \]

   \[ K_c = \frac{(4.79 * 10^{-4} M) (4.79 * 10^{-4} M)}{(3.53 * 10^{-3} M)^2} \]

   \[ K_c = \frac{2.29441 * 10^{-7} M^2}{1.24609 * 10^{-5} M^2} \]

   \[ K_c = 1.841 * 10^{-2} \]

2. a) N₂ (g) + O₂ (g) ⇌ 2NO (g)

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

   The reaction is a homogeneous reaction because the reactants/products all have the same phase.

   b) FeO (s) + H₂ (g) ⇌ Fe (s) + H₂O (g)

   \[ K_c = \frac{[H_2O]}{[H_2]} \]

   FeO and Fe are solids so they are not included in equilibrium constants.

   The reaction is a heterogeneous reaction because the reactants/products have different phases.

3. Converting to K_c from K_p (number 7 from p. 655 in the textbook)

2NO (g) + O₂ (g) ⇌ 2NO₂ (g)

\[ K_p = 1.48 * 10^4 \text{ at } 184 \text{ °C} \]

We know that \[ K_p = K_c \left(\frac{RT}{\Delta n}\right) \], we are given K_p but not K_c, you can rearrange the equation to: \[ K_c = \]
\[
\frac{K_p}{(RT)^{-\Delta{n}}}
\]

Which can also be written as: \[K_c = K_p \cdot (RT)^{\Delta{n}}\]

Now that we have your formula, we need to convert 184 °C to Kelvin, K = 184 + 273 = 457 K

\[K_c = \frac{[NO_2]^2}{[NO]^2[O_2]}\]

\[-\Delta{n} = (\text{total number of moles of products}) - (\text{total number of moles in reactants})\]

\[-\Delta{n} = (2) - (3) = -1\]

\[\Delta{n} = -(-1)\]

\[R = 0.08206 \left(\frac{\text{Liter} \cdot \text{Atm}}{\text{Mole} \cdot \text{Kelvin}}\right)\]

Now, plug in all the numbers we found: \[K_c = K_p \cdot (RT)^{\Delta{n}}\]

\[K_c = (1.48 \times 10^4) \cdot (0.08206)(457K)^{-1}\]

\[K_c = 5.5 \times 10^5\]

[it would be the same if you used this equation: \(K_c = K_p \cdot (RT)\)]

**K_p**

1. a) \(2\text{C}_2\text{H}_4 \ (g) + 2\text{H}_2\text{O} \ (g) \rightleftharpoons 2\text{C}_2\text{H}_6 \ (g) + \text{O}_2 \ (g)\)

\[K_p = \frac{[C_2\text{H}_6]^2[\text{O}_2]}{[C_2\text{H}_4]^2[\text{H}_2\text{O}]^2}\]

The reaction is a homogeneous reaction because the reactants/products all have the same phase.

b) \(\text{Ti} \ (s) + 2\text{Cl}_2 \ (g) \rightleftharpoons \text{TiCl}_4 \ (g)\)

\[K_p = \frac{[\text{TiCl}_4]}{[\text{Cl}_2]^2}\), Ti is a solid so it is not included in equilibrium constants.

The reaction is a heterogeneous reaction because the reactants/products have different phases.

2. Find \(K_p\), when \(K_c\) is given: (number 8 from p. 655 in the textbook)

\(2\text{H}_2\text{S} \ (g) + \text{CH}_4 \ (g) \rightleftharpoons 4\text{H}_2 \ (g) + \text{CS}_2 \ (g)\)

\[K_c = 5.27 \times 10^{-8} \text{ at } 973 \text{ K.}\]

\[K_p = K_c \cdot (RT)^{\Delta{n}}\), since Temperature is already converted to Kelvin and \(R = 0.08206 \left(\frac{\text{Liter} \cdot \text{Atm}}{\text{Mole} \cdot \text{Kelvin}}\right)\)

We need to find \(\Delta{n}\): \[K_p = \frac{[\text{H}_2]^4[\text{CS}_2]}{[\text{H}_2\text{S}]^2[\text{CH}_4\text{O}]^2}\]

\[\Delta{n} = (\text{total number of moles of products}) - (\text{total number of moles in reactants})\]
\( \Delta n = (5) - (3) = 2 \)

We can plug in our numbers: \( K_p = (5.27 \times 10^{-8})(0.08206)(973) \)
\( K_p = 3.6 \times 10^{-4} \)

3. The two common chlorides of Phosphorus, PCl\(_3\) and PCl\(_5\), both important in the production of other phosphorous compounds, coexist in equilibrium through: (number 17 from p. 655 in the textbook)

PCl\(_3\) (g) + Cl\(_2\) (g) \( \rightleftharpoons \) PCl\(_5\) (g)

At 250 °C, an equilibrium mixture in a 2.50 L flask contains 0.105 g PCl\(_5\), 0.220 g PCl\(_3\), and 2.12 g Cl\(_2\). What are the values of:

(a) \( K_c \)

[We need to convert grams to Molarity (mol/L), so we multiply grams with the molar mass and divide by Liters.]

PCl\(_5\) = \( \dfrac{0.105 \text{ g}}{2.50 \text{ L}} \times \dfrac{1 \text{ mol}}{137.3 \text{ g}} \) (Molar Mass) = 2.0173 \times 10^{-4} M

PCl\(_3\) = \( \dfrac{0.220 \text{ g}}{2.50 \text{ L}} \times \dfrac{1 \text{ mol}}{137.3 \text{ g}} \) (Molar Mass) = 6.4093 \times 10^{-4} M

Cl\(_2\) = \( \dfrac{2.12 \text{ g}}{2.50 \text{ L}} \times \dfrac{1 \text{ mol}}{70.9 \text{ g}} \) (Molar Mass) = 0.0119605

\( K_c = \dfrac{[PCl_5]}{[Cl_2][PCl_3]} \)
\( K_c = \dfrac{2.0173 \times 10^{-4}}{0.0119605 \times 6.4093 \times 10^{-4}} \)
\( K_c = 26.32 \)

(b) \( K_p \)

\( K_p = K_c (RT) \)

We need to find: \( \Delta n \) and \( K \), \( R = 0.08206 \) \( \dfrac{\text{L atm}}{\text{mol K}} \)

K = 250 \( \circ \) C + 273 = 523 K
\( \Delta n = (1) - (2) = -1 \)
\( K_p = (26.32) [(0.08206)(523)]^{(-1)} \)
\( K_p = 0.6133 \)
Contributors

- Rubi Medrano, Sharon Wei (UCD)