Chem Factsheet

Chemical Equilibria: K_c

To succeed in this topic you need to:-
* understand the concepts of 'dynamic equilibrium' and Le Chatalier's Principle (Factsheet 09)

After working through this Factsheet you will be able to
* understand the concept of 'K_c' and write its mathematical expression from the balanced chemical equation.
* work out the units of K_c
* perform calculations using the K_c expression

What is K_c?

The 'c' in K_c stands for 'concentration' and it is always expressed in mol dm^{-3} (moles/litre = moles/dm^3). To show this square brackets,[ ], are used around the species involved.

K_c is a constant numerical value for any equilibrium reaction but is always quoted at a particular temperature. This is because **only temperature affects the value of K_c**. In other words, you can add or remove reactants and products, change the pressure of the system or add a catalyst and the position of equilibrium will change (according to Le Chatalier's Principle) to keep K_c a constant value.

> **Exam hint:-** *What affects the K_c value is invariably tested in questions!*

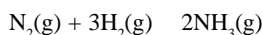## Writing K_c expressions

Look at the example below:

$$w A + x B \ û \ y C + z D$$

$$K_c = \frac{[C]_{eq}^{\ y} \times [D]_{eq}^{\ z}}{[A]_{eq}^{\ w} \times [B]_{eq}^{\ x}}$$

The example illustrates the rules for writing K_c expressions:

> **Rules**
> 1. 'Products over 'reactants' (right-hand side on top)
> 2. Numbers in front become powers
> 3. '+' becomes '×'

Another example:

$$N_2(g) + 3H_2(g) \ û \ 2NH_3(g)$$

$$K_c = \frac{[NH_3]^2_{eq}}{[N_2]_{eq} \times [H_2]^3_{eq}}$$

Every K_c has its own unit and this must be worked out every time (see box on working with powers). For the example above:

$$K_c = \frac{(mol\ dm^{-3})^2}{(mol\ dm^{-3}) \times mol\ dm^{-3})^3} = \frac{(mol\ dm^{-3})^2}{(mol\ dm^{-3})^4}$$

$$= \frac{1}{(mol\ dm^{-3})^2} = (mol\ dm^{-3})^{-2} = mol^{-2}dm^6$$

Sometimes the expression cancels down to '1' i.e. 'no units'. **You must write 'no units' in these cases or the examiner will assume you have not tried to work out the units.**

K_c is a 'concentration term' so can **only** be used for **solutions** and **gases**. You cannot have concentrations of solids, so these are ignored when writing the K_c expression.

The following examples illustrate the K_c expression for homogeneous and heterogeneous equilibria:

Example 1: $2SO_2(g) + O_2(g) \ û \ 2SO_3(g)$

(a homogeneous reaction i.e. all species in the same phase)

$$K_c = \frac{[SO_3]_{eq}^{\ 2}}{[SO_2]_{eq}^2 \times [O_2]_{eq}} \qquad units = mol^{-1}dm^3$$

Example 2: $CaCO_3(s) \ û \ CaO(s) + CO_2(g)$

(a heterogeneous reaction i.e. the species are in different phases)

$$K_c = [CO_2]_{eq} \qquad units = mol\ dm^{-3}$$

(NB. $CaCO_3$ and CaO are solids)

> **Exam hint:-** *The concept of K_c, and its mathematical treatment, is a fundamental part of the A2 course. Questions commonly try to 'catch out' candidates by including equations of heterogeneous reactions i.e. state symbols for solids and liquids which do <u>not</u> appear on K_c expressions.*

Question 1 at the end of this Factsheet provides further practice on writing K_c expressions and working out the units of K_c

> **Working with powers**
> To work with powers, you need to know these rules:
> * If something doesn't have a power, it means the power is 1
>   e.g. mol = mol^1
> * If you have a power in the bottom of a fraction, you can bring it to the top by changing the sign
>   e.g. $1/mol^2 = mol^{-2}$ $1/dm^{-3} = dm^3$
> * Anything to the power 0 is 1
>   e.g. mol^0 = 1
> * Multiply the numbers $\Rightarrow$ add the powers
>   e.g. $dm^{-3} \times dm^{-3} = dm^{-3+-3} = dm^{-6}$
> * Divide the numbers $\Rightarrow$ subtract the powers
>   e.g. $mol^2/mol^3 = mol^{2-3} = mol^{-1}$
> * To find a power of a power, multiply the powers
>   e.g. $(dm^{-3})^2 = dm^{-3 \times 2} = dm^{-6}$
>
> When simplifying an expression, it is usually easiest to work in this order:
>
> 1. Work out any powers eg $(mol\ dm^{-3})^2 = mol^2dm^{-6}$
> 2. Simplify the whole of the top of the fraction, and the whole of the bottom of the fraction eg $moldm^{-3} \times moldm^{-3} = mol^{1+1}dm^{-3+-3} = mol^2dm^{-6}$
> 3. Divide the top of the fraction by the bottom, (i.e subtract powers)
>   e.g. $(mol\ dm^{-3})/(mol^2dm^{-6}) = mol^{1-2}dm^{-3--6} = mol^{-1}dm^3$

K_c values and the effect of temperature

Since K_c is a number, its value tells you the about position of equilibrium.

1. Low K_c value e.g. $K_c = 2.4 \times 10^{-1}$

   Since $K_c = \dfrac{[products]}{[reactants]}$

   there must be more reactants than products so the equilibrium lies to the left

2. High K_c value e.g. $K_c = 36.7$

   Since $K_c = \dfrac{[products]}{[reactants]}$

   there must be more products than reactants so the equilibrium lies to the right.

How the K_c value is changed by changes in temperature depends on whether the forward reaction is exothermic or endothermic:

e.g. $N_2(g) + 3H_2(g) \,û\, 2NH_3(g)$ $\triangle H = -92kJmol^{-1}$

The forward reaction is exothermic so if the temperature is raised the equilibrium moves to the left (Le Chatalier's) i.e. more reactants, so the K_c value gets less. The reverse is true if the temperature is reduced the equilibrium moves to the right producing more products, so the K_c value increases.

The same logic applies to endothermic reactions.
Question (2) at the end of this Factsheet provides practice on working out the effect of temperature on K_c values

K_c calculations
There are two general types of calculations using the K_c expression:
1. finding the value of K_c given the amounts of species present,
2. finding the concentrations of species given the K_c value.

Finding the value of K_c
Method 1
You are provided with the equilibrium amounts in the question.

Example 1.
This question refers to the following system in equilibrium:

$$2A(g) + B(g) \,û\, 2C(g)$$
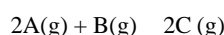
At equilibrium there were found to be 3×10^{-3} mol of A, 2×10^{-3} mol of B and 6.2×10^{-3} mol of C with a total volume of 2 dm^3

1. Write the K_c expression:

$$K_c = \frac{[C]_{eq}^2}{[A]_{eq}^2[B]_{eq}}$$

2. Substitute the values in

$$K_c = \frac{\left[\frac{6.2 \times 10^{-3}}{2}\right]^2}{\left[\frac{3 \times 10^{-3}}{2}\right]^2 \times \left[\frac{2 \times 10^{-3}}{2}\right]}$$

NB. [ ] = mol dm^{-3}, hence the need to use the 2dm^3 from the question to obtain 'per 1dm^3'.

3. Calculate the answer and units $K_c = 4.27 \times 10^3$ mol^{-1} dm^3

This type of calculation is fairly straightforward as long as the K_c expression is written correctly and the conversion using the volume is remembered.

Method 2
You need to work out the amounts at equilibrium first, then find K_c. This is the most common type of question! The method is outlined below:

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1. Write out the equation
2. Put the starting amounts under the equation
3. Put the equilibrium amount(s) from the question under the starting amounts
4. Use the information from the equation to calculate the 'missing' equilibrium amount(s), and find the concentrations
5. Write the K_c expression
6. Substitute the equilibrium amounts (using concentrations)
7. Calculate the K_c value and its units.

Example 1.
This problem refers to the equilibrium : $A(g) + B(g) \,û\, C(g)$

0.50 mol of A and 0.40 mol of B were mixed in a 10 dm^3 container. At equilibrium there were 0.2 mol of A. What is the value of K_c?

1.   A + B û   C   equation

2.   0.50   0.40   -   starting amounts

3.   0.20   ?   ?   equilibrium amounts

4. If only 0.20 mol of A remains then 0.30 mol has converted to C. Since A:B:C is 1:1:1 then 0.30 mol of B must also convert to C. So 0.30 mol of C is produced, and 0.1 mol of B remains.
   So we have

       A + B û C
       0.20   0.10   0.30
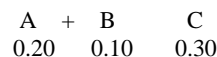
   Now find the concentrations by dividing by the volume (10dm^3):
   A: 0.020   B: 0.010   C: 0.030

5. $K_c = \dfrac{[C]_{eq}}{[A]_{eq} \times [B]_{eq}}$

6. $K_c = \dfrac{0.030}{0.020 \times 0.010}$

7. $K_c = 150$ mol^{-1} dm^3

Example 2.
This problem refers to the equilibrium $2R(g) \,û\, S(g) + T(g)$

2 mol of R was allowed to decompose into S and T until equilibrium was reached. At equilibrium there was 0.5 mol of S and the total volume was 2 dm^3

Answer   1.   2R   û   S   + T
          2.   start   2   -   -
          3.   eqm   ?   0.5   ?

4. Since S : T is 1 : 1, we must have 0.5 moles of T at equilibrium. Since R : S is 2 : 1, we need 1 mole of R to make 0.5 moles of S So there is 1 mole of R left

               2R   û   S   +   T
     eqm   1   0.5   0.5
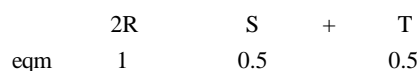
5. $K_c = \dfrac{[S]_{eq} \times [T]_{eq}}{[R]_{eq}^2}$

6,7 $K_c = 0.5 \times 0.5/1 = 0.25$ (no units)

Question (3) and (4) at the end of this Factsheet provides further practice on finding K_c.

Finding equilibrium concentrations from K_c values
This type of calculation involves being given the value of K_c and some data about the concentrations of the species in the equilibrium reaction.

Method 1
You are provided with the K_c value and some of the equilibrium concentrations

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Example
This question refers to the following equilibrium system:

$$N_2O_4(g) \; ⇌ \; 2NO_2(g)$$

At 250°C, the K_c value is 13.7 mol dm^{-3}, and in the equilibrium mixture there is 0.4 mol dm^{-3} of N_2O_4. Calculate the concentration of NO_2 in the equilibrium mixture.

1. Write the K_c expression

$$K_c = \frac{[NO_2]_{eq}^2}{[N_2O_4]_{eq}}$$

2. Substitute the values from the question into the equation

$$13.7 = \frac{[NO_2]_{eq}^2}{0.4}$$

3. Calculate the answer, including the units

$[NO_2]_{eq}^2 = 0.4 \times 13.7 = 5.48$

$[NO_2] = \sqrt{5.48} = 2.34$ mol dm^{-3}

Method 2
You are provided with the K_c value and the **starting** amount of a species.

Example
For the following equilibrium system, $K_c = 0.1$

$$X(g) \; ⇌ \; Y(g)$$
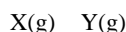
If 10 moles of X were allowed to reach equilibrium, how many moles of Y would there be in the equilibrium mixture?

1. Set the equilibrium amount of a product to be x. Work out the equilibrium amounts of all the other substances in terms of x.

If x moles of Y are produced, then x moles of X must be used up.
So we have:

	X(g)	⇌ Y(g)
start	10	0
eqm	10 − x	x

2. Convert to concentrations

Since we do not know the volume, call it V.
So: $[X] = (10 - x)/V$
$[Y] = x/V$

3. Write the K_c expresion and substitute in:

$$K_c = \frac{[Y]_{eq}}{[X]_{eq}}$$

$$0.1 = \frac{x/V}{(10-x)/V}$$

4. Rearrange the equation to find x

First cancel the Vs: $\qquad 0.1 = \dfrac{x}{(10-x)}$

Then multiply up: $\quad 0.1(10-x) \quad = x$
Expand brackets: $\quad 1 - 0.1x \quad = x$
Rearrange to find x: $\quad 1 \quad = 1.1x$
$\qquad\qquad\qquad x = 1/1.1 = 0.91$ moles

Exam Hint:- *You will **never** have to solve a quadratic equation using the quadratic equation formula. If you ever find yourself with such an equation, you have made a mistake!*

Practice Questions
1. For each of the following equilibrium reactions write the K_c expression and calculate the units:

   (a) $2NH_3(g) ⇌ N_2(g) + 3H_2(g)$
   (b) $2SO_2(g) + O_2(g) ⇌ 2SO_3(g)$
   (c) $C(s) + O_2(g) ⇌ CO_2(g)$
   (d) $N_2O_4(g) ⇌ 2NO_2(g)$
   (e) $2HI(g) ⇌ H_2(g) + I_2(g)$
   (f) $MgCO_3(s) ⇌ MgO(s) + CO_2(g)$
   (g) $C_2H_5OH(l) + CH_3COOH(l) ⇌ CH_3COOC_2H_5(l) + H_2O(l)$
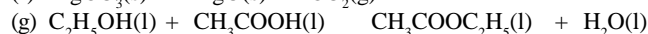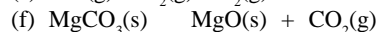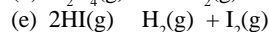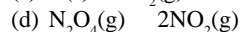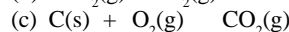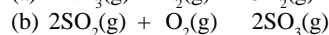
2. (a) The reaction: $2A(g) ⇌ B(g) + C(g)$
   has $K_c = 2.67 \times 10^{-2}$ at 350 K.
   When heated to 450K the K_c value changes to 0.98.
   What can be deduced about the equilibrium from this information?

   (b) The question is about the reaction,

   $$X(g) ⇌ 2Y(g) \qquad\qquad \Delta H = -15 \text{ kJ mol}^{-1}$$

   The K_c value is 7.4 mol dm^{-3} at 300K.
   If the temperature is increased how do the following change?
   (i) The value of K_c
   (ii) The rate of reaction

3. Calculate the value of K_c for each of the following and its units
   (a) For the reaction: $2HI (g) ⇌ H_2(g) + I_2(g)$
   At equilibrium there are 1.8×10^{-2} mol dm^{-3} of hydrogen iodide and the hydrogen and iodine are both at a concentration of 3.2×10^{-3} mol dm^{-3}.

   (b) At the equilibrium position for the reaction,

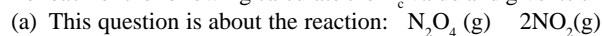 $$W(g) ⇌ X(g) + Y(g)$$
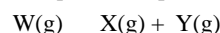
   the amounts of W, X & Y are 0.25 moles, 0.01 moles of 0.17 moles respectively. The total volume is 1.2 dm^3.

   (c) When equilibrium is reached for the reaction,

   $$2NH_3(g) ⇌ \qquad N_2(g) + 3H_2(g)$$

   There are 0.020 mol NH_3, 0.015 mol N_2 and 0.050 mol H_2 in a total volume of 750cm^3.

4. For each of the following calculate the K_c value and give its units:
   (a) This question is about the reaction: $N_2O_4 (g) ⇌ 2NO_2(g)$
   0.4 moles of N_2O_4 is allowed to decompose. At equilibrium there is found to be 0.1 mole of N_2O_4 in the 0.5 dm^3 gaseous mixture.

   (b) In the reaction: $2SO_2(g) + O_2(g) ⇌ 2SO_3(g)$
   2 mol SO_2 and 2 mol O_2 react together until the equilibrium position is reached. In a total volume of 10 dm^3 there is found to be 1 mole of SO_3.

   (c) When 1×10^{-2} mol ethanoic acid reacts with 2×10^{-2} mol ethanol according to the equation,

   $$CH_3COOH(l) + C_2H_5OH(l) ⇌ CH_3COOC_2H_5(l) + H_2O(l)$$

   there is found to be 5×10^{-3} moles of water in the equilibrium mixture.

   (d) When 0.4 mol M is mixed with 0.4 mol N they react together according to the following equation : $2M(g) + N(g) ⇌ 4P(g)$
   In a total volume of 2 dm^3 there is found to be 0.3 mol.

   (e) This question is about the equilibrium : $C(s) + O_2(g) ⇌ CO_2 (g)$
   1×10^{-2} mol oxygen reacts with carbon and the system reaches the equilibrium position. In the total volume of 200 cm^3 there are 7.5×10^{-3} moles of carbon dioxide.

5. (a) K_c is 15.6 mol^{-1} dm^3 at 370°C for the system

     $PCl_3(g) + Cl_2(g) \; ⇌ \; PCl_5(g)$
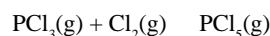
At equilibrium, the concentration of PCl_5 is 0.6 mol dm^{-3}. What are the concentrations of the PCl_3 and Cl_2 if there were equimolar amounts of each at the start?
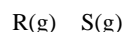
(b) For the equilibrium system

     $A(g) \; ⇌ \; 2B(g)$
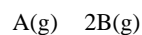
$K_c = 0.02$ mol dm^{-3} at 170° C.
At equilibriu,m the concentration of B is 0.12 mol dm^{-3}.
What is the equilibrium conentration of A?

(c) K_c is 0.88 at 200°C for the system

     $R(g) \; ⇌ \; S(g)$

If there were 2 moles of R at the start, how many moles of R would there be at equilibrium?

(d) For the system

     $F(g) + G(g) \; ⇌ \; H(g) + J(g)$
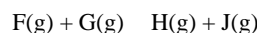
$K_c = 1$ at 120°C
Six moles of F and six moles of G were used at the start. How many moles of J will there be at equilibrium?

Answers

1. (a) $K_c = \dfrac{[N_2]_{eq} \times [H_2]_{eq}^{\,3}}{[NH_3]_{eq}^{\,2}}$ mol^{2}dm^{-6}

(b) $K_c = \dfrac{[SO_3]_{eq}^{\,2}}{[SO^2]_{eq}^{\,2} \times [O_2]_{eq}}$ mol^{-1}dm^3

(c) $K_c = \dfrac{[CO_2]_{eq}}{[O_2]_{eq}}$ (C(s)!!) No units

(d) $K_c = \dfrac{[NO_2]_{eq}^{\,2}}{[N_2O_4]_{eq}}$ mol dm^{-3}

(e) $K_c = \dfrac{[H_2]_{eq} \times [I_2]_{eq}}{[HI]_{eq}^{\,2}}$ No units

(f) $K_c = [CO_2]_{eq}$ (solids!!) mol dm^{-3}

(g) $K_c = \dfrac{[CH_3COOC_2H_5]_{eq} \times [H_2O]_{eq}}{[C_2H_5OH]_{eq} \times [CH_3COOH]_{eq}}$ No units

2. (a) the forward reaction is endothermic

(b) (i) Decreases/gets less

     (ii) Increases

3. (a) 3.16×10^{-2} no units
(b) 5.67×10^{-3} mol dm^{-3}
(c) 8.2×10^{-3} mol^{2}dm^{-6}

4. (a) 1.8 mol dm^{-3} ($[N_2O_4]_{eq} = 0.05$, $[NO_2]_{eq} = 0.3$)

(b) 6.67 mol$^-$ dm^3 ($[SO_2]_{eq} = 0.1$, $[O_2]_{eq} = 0.15$, $[SO_3]_{eq} = 0.1$)

(c) 0.33 no units   ($K_c = (0.5 \times 10^{-2})^2 / (0.5 \times 10^{-2} \times 1.5 \times 10^{-2})$)

(d) 0.025 mol dm^{-3} ($[M]_{eq} = 0.15$, $[N]_{eq} = 0.175$, $[P]_{eq} = 0.1$)

(e) 3.00 no units   C is a solid so does not appear to K_c express.

     ($[O_2]_{eq} = 1.25 \times 10^{-2}$, $[CO_2]_{eq} = 3.75 \times 10^{-2}$)

5. (a) $15.6 = \dfrac{0.6}{x^2}$

$15.6x^2 = 0.6$
$x^2 = 0.6/15.6 = 0.03846$
$x = \sqrt{0.03846} = 0.196$ moles

(b) $0.02 = \dfrac{0.12^2}{A}$

$0.02A = 0.12^2$
$A = 0.12^2/0.02 = 0.72$ moles

(c) $0.88 = \dfrac{x}{2-x}$

$0.88 (2 - x) = x$
$1.76 - 0.88x = x$
$1.76 = 1.88x$
$x = 1.76/1.88 = 0.936$ moles

(d) At equilibrium, we have x moles of J and H and $6 - x$ moles of F and G

$1 = \dfrac{x^2}{(6 - x)(6 - x)}$

$(6 - x)(6 - x) = x^2$
$36 - 12x + x^2 = x^2$
$36 = 12x$
$x = 3$ moles