

CHEMICAL EQUILIBRIUM PART 2

EQUILIBRIUM CONSTANT K_c

The equilibrium constant, K_c , is a number that indicates the ratio of the concentrations of products and reactants at equilibrium.

It is a constant for a particular reaction at constant temperature.

For an equilibrium with a balanced equation like this: $aA + bB \rightleftharpoons cC + dD$
 $K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ where $[A]$ means the concentration of A at equilibrium

Hydrogen and iodine react reversibly to form hydrogen iodide like this: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

(a) Write an expression for K_c .

Reactants on bottom

Products on top (power 2 from balanced equation)

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

(b) The table shows the concentrations of reactants and products at equilibrium. Use the data to calculate a value for K_c . Give your answer to an appropriate number of significant figures.

Chemical	H_2	I_2	HI
Concentration (mol dm ⁻³)	0.405	0.357	0.290

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{[0.290]^2}{0.405 \times 0.357} = 0.582$$

All data is given to three significant figures – so you must give your answer to three significant figures.

There are no units as the units on the top and bottom of the fraction are the same (mol² dm⁻⁶) and so cancel out

When K_c is tiny, the amount of product at equilibrium is close to zero

Example 1

CALCULATING K_c

You may be asked to construct an expression for K_c as shown on this sheet. Then you may be given concentrations of all reactants and all products at equilibrium and asked to calculate a value for K_c .

Or you may be given a value for K_c along with the concentrations of some reactants and products at equilibrium and asked to calculate the equilibrium concentration of other reactants or products.

You may also be asked how changes in temperature affect the value of K_c (remember catalysts or changing concentrations have no effect on K_c).

INTERPRETING K_c

The value of K_c tells you the position of equilibrium, i.e. the amounts of reactants and products present at equilibrium.

Large K_c values mean lots of product – equilibrium lies to the right

Small K_c values mean lots of reactant – equilibrium lies to the left

Very large: $K_c > 1000$ indicates a reaction is nearly complete

Very small: $K_c < 0.001$ indicates a reaction not really happening

Example 3

Nitrogen and oxygen react reversibly to form nitrogen monoxide: $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

(a) Write an expression for K_c .

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

-ve indices mean small values

(b) At 100 °C K_c for this reaction is 6.0×10^{-28} . What does this value tell you about the reaction of nitrogen and oxygen at this temperature?

Because K_c is so small, the reaction is effectively not happening

(c) At 3000 °C, K_c for this reaction is 0.0020. In an equilibrium mixture where the concentration of oxygen was 0.20 mol dm⁻³ and the concentration of nitrogen was 0.80 mol dm⁻³, what would the concentration of nitrogen monoxide be?

$$K_c = 0.0020 = \frac{[NO]^2}{[N_2][O_2]} = \frac{[NO]^2}{0.20 \times 0.80} \quad \text{so } [NO]^2 = 0.0020 \times 0.20 \times 0.80 = 3.2 \times 10^{-4}$$

$$[NO] = \sqrt{3.2 \times 10^{-4}} = 0.018 \text{ mol dm}^{-3}$$

Square root of both sides

Example 4

In the Haber process, ammonia is formed from nitrogen and hydrogen: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
 At 500 °C, $K_c = 179 \text{ dm}^3 \text{ mol}^{-1}$. The forwards reaction is exothermic ($\Delta H = -ve$).

(a) Write an expression for K_c

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

$$\text{units} = \frac{[\text{mol dm}^{-3}]^2}{[\text{mol dm}^{-3}][\text{mol dm}^{-3}]^3} = \frac{[\text{mol dm}^{-3}]^2}{[\text{mol dm}^{-3}]^4}$$

(b) Would raising the temperature cause K_c to increase or decrease? Would this give a higher or lower yield of ammonia? Explain your answer.

At higher temperature, K_c would have a lower value and the yield of ammonia would be less.

According to Le Chatelier, as temperature increases for exothermic reactions, the position of equilibrium moves left. This means that there is less NH_3 in the equilibrium mixture and K_c decreases.

(c) Why is this reaction carried out in industry at high temperatures (often over 1000 °C)?

Although high temperature produces a low yield, it causes the reaction to reach equilibrium faster. A compromise between rate and position of equilibrium is made.

Example 2

Iodomethane reacts reversibly with water like this: $CH_3I(aq) + H_2O(l) \rightleftharpoons CH_3OH_2^+(aq) + I^-(aq)$
 Under certain conditions, $K_c = 1.89 \text{ mol dm}^{-3}$

(a) Write an expression for K_c .

$$K_c = \frac{[CH_3OH_2^+][I^-]}{[CH_3I]}$$

$[H_2O]$ does not appear – since H_2O is the solvent it is in a huge excess and $[H_2O]$ is approximately constant

(b) An equilibrium mixture has equal amounts of $CH_3OH_2^+$ and I^- . Both have a concentration of 0.250 mol dm⁻³. Calculate the concentration of CH_3I in the equilibrium mixture.

$$K_c = 1.89 = \frac{[CH_3OH_2^+][I^-]}{[CH_3I]} = \frac{0.250 \times 0.250}{[CH_3I]}$$

$$[CH_3I] = \frac{0.250 \times 0.250}{1.89} = 0.0331 \text{ mol dm}^{-3}$$

Answer has three significant figures (same as data)