

2. Equilibrium

Equilibrium constant K_c

K_c = equilibrium constant

For a generalised reaction



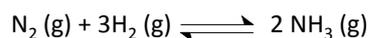
m, n, p, q are the stoichiometric balancing numbers

A, B, C, D stand for the chemical formula

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

[] means the equilibrium concentration

Example 1



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

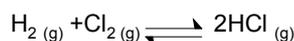
The unit of K_c changes and depends on the equation.

Working out the unit of K_c

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

$$K_c = \frac{[NH_3(g)]^2}{[N_2(g)][H_2(g)]^3} \rightarrow \text{Unit} = \frac{[\text{mol dm}^{-3}]^2}{[\text{mol dm}^{-3}][\text{mol dm}^{-3}]^3} \xrightarrow{\text{Cancel out units}} \text{Unit} = \frac{1}{[\text{mol dm}^{-3}]^2} \rightarrow \text{Unit} = [\text{mol dm}^{-3}]^{-2} \downarrow \text{Unit} = \text{mol}^{-2} \text{dm}^6$$

Example 2: writing K_c expression



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

Working out the unit

$$\text{Unit } K_c = \frac{[\text{mol dm}^{-3}]^2}{[\text{mol dm}^{-3}][\text{mol dm}^{-3}]} = \text{no unit}$$

Calculating K_c

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 0.6mol of Cl_2 . At equilibrium there were 0.2moles of HCl . Calculate K_c

| | H_2 | Cl_2 | HCl |
|-------------------|-------|--------|-------|
| Initial moles | 0.5 | 0.6 | 0 |
| Equilibrium moles | | | 0.2 |

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

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

Work out the moles at equilibrium for the reactants

moles of hydrogen at equilibrium = $0.5 - 0.1 = 0.4$

moles of reactant at equilibrium = initial moles – moles reacted

moles of chlorine at equilibrium = $0.6 - 0.1 = 0.5$

| | H_2 | Cl_2 | HCl |
|-------------------------------|------------------|------------------|------------------|
| Initial moles | 0.5 | 0.6 | 0 |
| Equilibrium moles | 0.4 | 0.5 | 0.2 |
| Equilibrium concentration (M) | $0.4/0.6 = 0.67$ | $0.5/0.6 = 0.83$ | $0.2/0.6 = 0.33$ |

If the K_c 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 K_c expression

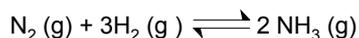
$$K_c = \frac{[HCl(g)]^2}{[H_2(g)][Cl_2(g)]} = \frac{0.33^2}{0.67 \times 0.83} = 0.196 \text{ no unit}$$

Work out the equilibrium concentrations

conc = moles / vol (in dm^3)

Finally put concentrations into K_c expression

Example 4



For the following equilibrium

Initially there were 1.5 moles of N_2 and 4 mole of H_2 in a 1.5 dm^3 container. At equilibrium 30% of the Nitrogen had reacted. Calculate K_c

| | N_2 | H_2 | NH_3 |
|-------------------|--------------|--------------|---------------|
| Initial moles | 1.5 | 4.0 | 0 |
| Equilibrium moles | | | |

30% of the nitrogen had reacted = $0.3 \times 1.5 = 0.45$ moles reacted.
Using the balanced equation 3×0.45 moles of H_2 must have reacted and 2×0.45 moles of NH_3 must have formed

Work out the moles at equilibrium for the reactants and products

moles of reactant at equilibrium = initial moles – moles reacted

moles of nitrogen at equilibrium = $1.5 - 0.45 = 1.05$ moles of hydrogen at equilibrium = $4.0 - 0.45 \times 3 = 2.65$

moles of product at equilibrium = initial moles + moles formed

moles of ammonia at equilibrium = $0 + (0.45 \times 2) = 0.9$

| | N_2 | H_2 | NH_3 |
|-------------------------------|------------------|-------------------|-----------------|
| Initial moles | 0.5 | 0.6 | 0 |
| Equilibrium moles | 1.05 | 2.65 | 0.9 |
| Equilibrium concentration (M) | $1.05/1.5 = 0.7$ | $2.65/1.5 = 1.77$ | $0.9/1.5 = 0.6$ |

Finally put concentrations into K_c expression

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

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

Work out the equilibrium concentrations

conc = moles/ vol (in dm^3)

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

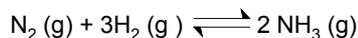
K_c only changes with temperature.

It does not change if pressure or concentration is altered.
A catalyst also has no effect on K_c

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

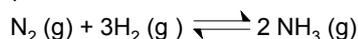


If temperature is increased the reaction will shift to oppose the change and move in the backwards endothermic direction. The position of equilibrium shifts left. The value of K_c gets smaller as there are fewer products

Effect of Pressure on position of equilibrium and K_c

The position of equilibrium will change if pressure is altered but the value of K_c stays constant

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



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