2. Equilibrium

Equilibrium constant Kc

Kc = equilibrium constant

For a generalised reaction

$mA + nB \rightleftharpoons pC + qD$

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

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

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

[ ] means the equilibrium concentration

Example 1

$N_2(g) + 3H_2(g) \rightleftharpoons 2\,NH_3(g)$

$$Kc = \frac{[NH_{3\,(g)}]^2}{[N_2(g)]\,[H_2(g)]^3}$$

The unit of Kc changes and depends on the equation.

Working out the unit of Kc

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

$$Kc = \frac{[NH_{3\,(g)}]^2}{[N_2(g)]\,[H_2(g)]^3} \rightarrow \text{Unit} = \frac{[\text{mol dm}^{-3}]^2}{[\text{mol dm}^{-3}]\,[\text{mol dm}^{-3}]^3} \xrightarrow{\text{Cancel out units}} \text{Unit} = \frac{1}{[\text{mol dm}^{-3}]^2} \rightarrow \text{Unit} = [\text{mol dm}^{-3}]^{-2}$$

$\downarrow$

Unit = $mol^{-2}\,dm^{+6}$

Example 2: writing Kc expression

$H_{2\,(g)} + Cl_{2\,(g)} \rightleftharpoons 2HCl_{(g)}$

$$Kc = \frac{[HCl_{(g)}]^2}{[H_2(g)]\,[Cl_2(g)]} \xrightarrow{\text{Working out the unit}} \text{Unit Kc} = \frac{[\text{mol dm}^{-3}]^2}{[\text{mol dm}^{-3}]\,[\text{mol dm}^{-3}]}$$

= no unit

Calculating Kc

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

	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 0.2 moles of HCl has been formed it must have used up 0.1 of Cl_2 and 0.1 moles of H_2 (as 1:2 ratio)

Work out the moles at equilibrium for the reactants

moles of reactant at equilibrium = initial moles – moles reacted

moles of hydrogen at equilibrium = 0.5 – 0.1 = 0.4

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 Kc 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 kc expression

$$Kc = \frac{[HCl_{(g)}]^2}{[H_2(g)]\,[Cl_2(g)]}$$

Work out the equilibrium concentrations

conc = moles/ vol (in dm^3)

Finally put concentrations into Kc expression

$$Kc = \frac{0.33^2}{0.67 \times 0.83}$$

= 0.196 no unit

Example 4

For the following equilibrium $N_2 (g) + 3H_2 (g) \rightleftharpoons 2 NH_3 (g)$

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 Kc

	N_2	H_2	NH_3
Initial moles	1.5	4.0	0
Equilibrium moles			

30% of the nitrogen had reacted = 0.3 x1.5 = 0.45 moles reacted. Using the balanced equation 3 x 0.45 moles of H_2 must have reacted and 2x 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 x3 = **2.65**

moles of product at equilibrium = initial moles + moles formed

moles of ammonia at equilibrium = 0 + (0.45 x 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

Work out the equilibrium concentrations

conc = moles/ vol (in dm^3)

Finally put concentrations into Kc expression

$$Kc = \frac{[NH_{3\,(g)}]^2}{[N_2 (g)]\,[H_2 (g)]^3}$$

$$Kc = \frac{0.6^2}{0.7 \times 1.77^3} = \mathbf{0.0927\ mol^{-2}\ dm^{+6}}$$

Effect of changing conditions on value of Kc

The larger the Kc the greater the amount of products.
If Kc is small we say the equilibrium favours the reactants

Kc only changes with temperature.
It does not change if pressure or concentration is altered.
A catalyst also has no effect on Kc

Effect of Temperature on position of equilibrium and Kc

Both the position of equilibrium and the value of Kc will change it temperature is altered

In this equilibrium which is exothermic in the forward direction

$N_2 (g) + 3H_2 (g) \rightleftharpoons 2 NH_3 (g)$

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 Kc gets smaller as there are fewer products

Effect of Pressure on position of equilibrium and Kc

The position of equilibrium will change it pressure is altered but the value of Kc stays constant

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

$N_2 (g) + 3H_2 (g) \rightleftharpoons 2 NH_3 (g)$

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 Kc stays the same though as only temperature changes the value of Kc

Catalysts have no effect on the value of Kc or the position of equilibrium