Inquiry question: How can the position of equilibrium be described and what does the equilibrium constant represent?

Students:

- deduce the equilibrium expression (in terms of Keq) for homogeneous reactions occurring in solution
- perform calculations to find the value of Keq and concentrations of substances within an equilibrium system, and use these values to make predictions on the direction in which a reaction may proceed
- qualitatively analyse the effect of temperature on the value of Keq
- conduct an investigation to determine Keq of a chemical equilibrium system, for example:
  - Keq of the iron(III) thiocyanate equilibrium
- explore the use of Keq for different types of chemical reactions, including but not limited to:
  - dissociation of ionic solutions
  - dissociation of acids and bases

Equilibrium Expression

Homogeneous reactions in terms of equilibrium reactions are when all of the substances are in the **same** state of matter: meaning all of them are either solids, liquids, gases or aqueous solutions.

Heterogeneous reactions in terms of equilibrium reactions are when the substances are of **different** state of mater. Some can be solids, liquids, gases and/or aqueous solutions.

An example of homogeneous reactions would be the Haber process:

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

The equilibrium expression would be $Keq = \frac{[NH3]^2}{[N2][H2]^3}$

The equilibrium expression in terms of K_{eq} (equilibrium constant) calculates and shows whether the equilibrium favours the products or the reactants. In other words, it calculates the ratio of the species – how much there are of products in relation to the reactants. The higher the K_{eq} value, then the higher the ratio of **products to reactants**. The lower the K_{eq} value, the higher the ratio of **reactants to products**.

Hence, the general equation of an equilibrium expression is:

$$K_{eq} = \frac{[\text{products}]}{[\text{reactants}]}$$

$$aA + bB \rightleftharpoons cC + dD$$

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

Note: K_c can be used interchangeably with K_{eq}. In terms of the chemical equation, the upper-case letters are the species, and the lower-case letters are for the co-efficient of the species (how many moles are present).

In terms of the equilibrium expression, you will note that C and D are products, and they are multiplied to its 'power'. The multiplication of all the 'relevant' products are then divided by the multiplication of all the 'relevant' reactants (to its power).

Note: Only aqueous solutions and gases play a role in the equilibrium expressions and only they are included in the mathematical equation/calculation.

Taking a look at an example:

$$CH_{4(g)} + 2O_{2(g)} \rightleftharpoons CO_{2(g)} + 2H_2O_{(l)}$$

$$K_{eq} = \frac{[CO_2]\cancel{[H_2O]^2}}{[CH_4][O_2]^2}$$

This is your equilibrium expression. Remember that you must use the co-efficient as the power. Also, since water is a liquid, it does not take a role in this equilibrium expression.

Another example:

$$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

$$K_c = \frac{[CaO][CO_2]}{[CaCO_3]}$$

However, both $CaCO_3$ and CaO are solids and do not play a role in this expression. The final expression is:

$K_{eq} = [CO_2]$

Calculating Equilibrium Constant

In the exam, you may well be asked to calculate the equilibrium constant of different equations.

The way to calculate the equilibrium constant is to:

1. Determine the equilibrium expression
2. Use the RICE table to calculate the concentration of substances at equilibrium
3. Use these values to input into the equilibrium expression to calculate the constant

Let's use the Haber Process example again:

Question: The initial concentration of N_2 was 1.0 mol and H_2 was 2.0 mol in a 2.0L container. At equilibrium, the concentration of NH_3 was at 0.60 mol. Find the equilibrium constant.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

$$Keq = \frac{[NH3]^2}{[N2][H2]^3}$$

	$N_{2(g)}$	$H_{2(g)}$	$NH_{3(g)}$
Ratio	1	3	2
Initial concentration (mol L^{-1})	0.5	1	0
Change in concentration (mol L^{-1})	- 0.15	- 0.45	+ 0.3
Equilibrium concentration (mol L^{-1})	0.35	0.55	0.3

Temperature and Equilibrium Constant

Endothermic reactions:

- If temperature is increased, the equilibrium will shift to the RHS (LCP), and more of the products are present.
    - Therefore, this will increase the equilibrium constant as the numerator is greater.
- If temperature is decreased, the equilibrium will shift to the LHS (LCP), and more of the reactants are present.
    - Therefore, this will decrease the equilibrium constant as the denominator is greater.

Exothermic reactions:

- If temperature is increased, the equilibrium will shift to the LHS (LCP), and more of the reactants are greater.
    - Therefore, this will decrease the equilibrium constant as the denominator is greater.
- If temperature is decreased, the equilibrium will shift to the RHS (LCP), and more of the products are greater.
    - Therefore, this will increase the equilibrium constant as the numerator is greater.

Dissociation of Acids and Bases

The equilibrium constant for the dissociation of acids is represented as K_a. The equilibrium constant for the dissociation of bases is represented as K_b.

You will learn more about this in module 6.

For now, remember the equilibrium constant equation will be the same as before.

K_a = [products] / [reactants]

K_b = [products] / [reactants]