8942 version 3
Characterise the nature of chemical systems at equilibrium
Level 2 Credits 2

Element 1
PC 1.1 Equilibrium constant expressions, K (or K_c) are written for *given* chemical equations.

Reactions don't stop when they come to equilibrium. There is no net change in the concentrations of the reactants or products, as the forward and reverse reactions are in balance at equilibrium. The reaction appears to stop on the macroscopic scale. Chemical equilibrium is a *dynamic* balance between opposing forces, the forward and reverse reactions.

<u>Writing Equilibrium Constant Expressions</u>

Even though chemical reactions at equilibrium occur in both directions, the reagents on the right side of the equation are assumed to be the "products" of the reaction and the reagents on the left side of the equation are assumed to be the "reactants."

reactants products

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

The ratio described is a constant, called the equilibrium constant (K).

"C" shows it is written in terms of concentration.

The square brackets represent the concentration of the substance within the brackets.

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

The products of the reaction are always written above the line (the numerator).

The reactants are always written below the line (the denominator).

The indices, e.g. b, the powers that you have to raise the concentrations to, are the numbers in front of each of the substances in the balanced chemical equation.

For homogeneous* systems (the only type that you will meet at Level 2 NCEA), the equilibrium constant expression contains a term for every reactant and every product of the reaction (*A homogeneous equilibrium has everything present in the same state or phase. E.g. reactions where everything is a gas or reactions where everything is present in the same aqueous solution).

Examples:

$2NO_2(g) \rightleftharpoons N_2O_4(g)$	$K_C = \dfrac{[N_2O_4(g)]}{[NO_2(g)]^2}$
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	$K_C = \dfrac{[NH_3(g)]^2}{[N_2(g)]\,[H_2(g)]^3}$
$2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$	$K_C = \dfrac{[N_2(g)]\,[H_2(g)]^3}{[NH_3(g)]^2}$
$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons [FeSCN]^{2+}(aq)$	$K_C = \dfrac{[\,[FeSCN]^{2+}(aq)]}{[Fe^{3+}(aq)]\,[SCN^{-}(aq)]}$

Common mistakes!

- Writing the expression upside down eg $\dfrac{[reactants]}{[products]}$ ✘
- Leaving in the +

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

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

FAQ

- K or K_C? *Either is fine for Level 2 NCEA for all the examples you will meet. If you are asked to write K expressions use K, if K_C then write K_C!*
- Do I need the state symbols? *Not for this Unit Standard but if you are putting them in, put them ALL in, and put them in the right place, inside the] bracket eg* $[H_2(g)]^3$ ✓ $[H_2)](g)^3$ ✘, $[\,[FeSCN]^{2+}(aq)]$

PC 1.2 Changes in equilibrium position are predicted as a result of a change in conditions.

You will be applying Le Chatelier's Principle to help you work out what happens when you change the conditions in a reaction in dynamic equilibrium.

If a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium moves to counteract the change.

Questions will either ask:

Effect on amount of named reactant / product - This will be "increase", "decrease" or "is unchanged" OR Effect on the equilibrium position - This will be "shift to left", "shift to right" or "no change on equilibrium position". Which one you use will depend upon the question that is asked.

Change	Effect
Change of temperature	o Increase in temperature favours the endothermic reaction o Decrease in temperature favours the exothermic reaction
Change of concentration	o Increase in conc. of a reactant (or decrease in concentration of a product) favours the formation of product / shifts equilibrium position to the right o Decrease in conc. of a reactant (or increase in concentration of a product) favours the formation of reactants / shifts equilibrium position to the left
Change of pressure	o An increase in pressure (volume is reduced) favours the reaction producing the smaller number of moles of gas o A decrease in pressure (volume is increased) favours the reaction producing the larger number of moles of gas o If the number of moles of gas are the same on each side then the amount / equilibrium position will be unchanged
Addition / removal of catalyst	o Amount of reactant / product / equilibrium position is unchanged. (But equilibrium is reached much faster)

Example 1

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g) \quad \Delta H = +\ 192\ kJ$$

How would the amount of $SO_2(g)$ alter if:

- o The temperature was increased – *the amount of $SO_2(g)$ would increase*
- o The pressure was increased - *the amount of $SO_2(g)$ would decrease*
- o A catalyst was added – the amount of *$SO_2(g)$ would be unchanged*
- o Some O_2 was removed – *the amount of $SO_2(g)$ would increase*

How would the equilibrium position alter if:

- o The temperature was increased – *it would shift to the right*
- o The pressure was increased - *it would shift to the left*
- o A catalyst was added – *it would be unchanged*
- o Some O_2 was removed – *it would shift to the right*

Example 2

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g) : \ -\Delta H$$

How will the amount of H_2 be affected by each of the following?

- o Some CO(g) is added – *the amount of $H_2(g)$ would increase*
- o The volume of the system is increased - *the amount of $H_2(g)$ would be unchanged (2 moles gas on each side)*

How will the equilibrium position alter if:

- o The temperature is decreased - *it would shift to the right*
- o Some $H_2O(g)$ is removed - *it would shift to the left*
- o Addition of more catalyst - *it would be unchanged*