

## AS AND A LEVEL

Delivery Guide

# CHEMISTRY A

**H032, H432**

For first teaching in 2015

## Theme: Equilibrium

Version 2

# AS AND A LEVEL CHEMISTRY A

Delivery guides are designed to represent a body of knowledge about teaching a particular topic and contain:

- Content: A clear outline of the content covered by the delivery guide;
- Thinking Conceptually: Expert guidance on the key concepts involved, common difficulties students may have, approaches to teaching that can help students understand these concepts and how this topic links conceptually to other areas of the subject;
- Thinking Contextually: A range of suggested teaching activities using a variety of themes so that different activities can be selected which best suit particular classes, learning styles or teaching approaches.

If you have any feedback on this Delivery Guide or suggestions for other resources you would like OCR to develop, please email [resources.feedback@ocr.org.uk](mailto:resources.feedback@ocr.org.uk).

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## A Level only

### 3.2.3 Chemical equilibrium

- (a) explanation that a dynamic equilibrium exists in a closed system when the rate of the forward reaction is equal to the rate of the reverse reaction and the concentrations of reactants and products do not change
- (b) Le Chatelier's principle and its application for homogeneous equilibria to deduce qualitatively the effect of a change in temperature, pressure or concentration on the position of equilibrium
- (c) explanation that a catalyst increases the rate of both forward and reverse reactions in an equilibrium by the same amount resulting in an unchanged position of equilibrium
- (d) the techniques and procedures used to investigate changes to the position of equilibrium for changes in concentration, temperature and in the presence of a catalyst
- (e) explanation of the importance to the chemical industry of a compromise between chemical equilibrium and reaction rate in deciding the operational conditions
- (f) expressions for the equilibrium constant,  $K_c$ , for homogeneous reactions and calculations of the equilibrium constant,  $K_c$ , from provided equilibrium concentrations
- (g) estimation of the position of equilibrium from the magnitude of  $K_c$ .

### 5.1.2 How far?

- (a) use of the terms mole fraction and partial pressure
- (b) calculation of quantities present at equilibrium, given appropriate data
- (c) the techniques and procedures used to determine quantities present at equilibrium
- (d) expressions for  $K_c$  and  $K_p$  for homogeneous and heterogeneous equilibria (see also 3.2.3 f)
- (e) calculations of  $K_c$  and  $K_p$ , or related quantities, including determination of units (see also 3.2.3 f)
- (f) (i) the qualitative effect on equilibrium constants of changing temperature for exothermic and endothermic reactions  
(ii) the constancy of equilibrium constants with changes in concentration, pressure or in the presence of a catalyst
- (g) explanation of how an equilibrium constant controls the position of equilibrium on changing concentration, pressure and temperature
- (h) application of the above principles in 5.1.2 How far? for  $K_c$ ,  $K_p$  to other equilibrium constants, where appropriate (see also 5.1.3 c etc).

### Approaches to teaching the content

Students will have studied reversible reactions at GCSE but the dynamic nature of the equilibrium state is first encountered at AS Level. This underlies an understanding of le Chatelier's principle, which refers to changes to a system which has already reached equilibrium. The calculation of  $K_c$  is now introduced at AS Level: showing students quantitatively why changes in one component require compensating changes in others can help them to put le Chatelier's principle in context as a quick and simple way of making predictions. The topic is revisited at A Level, when  $K_p$  is introduced together with the calculation of equilibrium quantities.

### Common misconceptions or difficulties students may have

Many students associate equilibrium with a static or balanced state, or even one in which the concentrations on both sides are equal. Underlying this is one of the biggest difficulties the chemistry student faces, distinguishing between macroscopic and molecular levels of description. Although there is no change in macroscopic properties at equilibrium, students have to realise they are still occurring at the molecular level.

Students often learn to apply le Chatelier's principle confidently but then struggle to adjust their thinking when confronted with the equilibrium constant. The introduction of  $K_c$  alongside le Chatelier may help mitigate this as a distinction can be drawn between the equilibrium **constant** and the position of equilibrium (which is affected by concentration). Equilibrium can be modelled as a seesaw, where  $K_c$  represents the position of the fulcrum. If  $K_c \gg 1$  the fulcrum will be well to the right and the seesaw balances when there are many more product molecules than reactants.

Many students will have difficulty accepting that a catalyst has no effect on the position of equilibrium, especially when we speak of enzymes catalysing "the digestion of starch into glucose", for example. Suitable analogies and consideration of enthalpy profiles with and without catalyst may help students to appreciate qualitatively that the catalyst speeds up both forward and backward reactions. The effects of pressure and temperature on  $K_c$  are also problematical: numerical examples can help students to understand that  $K_c$  is constant unless the temperature is changed.

A common error in exam questions is for students to use equilibrium amounts in calculations instead of converting them into concentrations or partial pressures.

### Conceptual links to other areas of the specification – useful ways to approach this topic to set students up for topics later in the course

Examples of equilibria involving redox reactions or complex ions provide convenient examples to refer back to at later stages of the course. The choice of conditions for some organic reactions can also be related to this topic: for example in the production of esters concentrated acid acts not only as a catalyst but also to absorb some of the water produced, shifting the position of equilibrium in favour of the condensation reaction.

Able students will be interested to see the underlying explanation of le Chatelier's principle when they encounter the effect of temperature on free energy change.

## Activities

### Dynamic Equilibrium Analogy: Video

[https://www.youtube.com/watch?v=hSxTNEyTogU&index=1&list=PL816Qsrt2Os394LaHIen\\_URp8ZFC90Kpg](https://www.youtube.com/watch?v=hSxTNEyTogU&index=1&list=PL816Qsrt2Os394LaHIen_URp8ZFC90Kpg)

This video uses the transfer of water between two containers as a simple analogy for dynamic equilibrium. It is one of a series of concise and well produced videos, most of which are relevant to the A Level course. The whole series is available at <https://www.youtube.com/user/richthornley>. This is potentially a very useful site as an ideas-bank for teachers and for independent use by students.

### Dynamic equilibrium simulation

<https://www3.fed.cuhk.edu.hk/chemistry/files/Simulation.pdf>

This student activity uses paperclips to simulate a dynamic equilibrium, which may be useful to help your students understand the concept.

### Reaching Equilibrium: Simulation

University of Colorado PhET project

<https://phet.colorado.edu/en/simulation/reversible-reactions>

The program displays a reaction enthalpy profile together with the numbers of reactant and product particles as the reaction progresses. Students can see how the numbers of each particle change as the reaction progresses and investigate the effect of changing the number of reactant or product particles, the enthalpy change and the temperature. They can also adjust the activation energy to simulate the addition of a catalyst.

### Equilibrium Spreadsheets

David Dice

<http://digipac.ca/chemical/equilibrium/>

This site provides a number of spreadsheets, with instructions, for modelling equilibrium. The high-school dance model helps to explain principles of equilibrium and the spreadsheet showing the interconversion of *E* and *Z*-but-2-ene generates graphs showing the rate at which equilibrium is approached and the position of equilibrium under various conditions.

### Rates and Equilibria

Royal Society of Chemistry: Chemistry for the Gifted & Talented

<http://www.rsc.org/learn-chemistry/resource/res00000646/rates-and-equilibria>

This activity is suitable for a small group of students working together. It involves organising information cards to explain the relationship between the rate of the forward and backward reaction and the equilibrium constant and the effect of temperature on  $K_c$ . This could appeal to mathematically-minded students as it helps to make links to the kinetics topic, emphasises the dynamic nature of equilibrium and shows them how to use  $K_c$  to explain the effects of concentration, temperature and pressure rather than simply applying le Chatelier's principle.

### Assessment for learning

Royal Society of Chemistry: LearnChemistry

<https://edu.rsc.org/resources/equilibria-starters-1-16andndash18/4010277.article> has a useful worksheets on le Chatelier's principle,  $K_c$  and  $K_p$ .

### Applying le Chatelier's Principle

Nuffield Advanced Science Data Book

<http://www.nationalstemcentre.org.uk/elibrary/resource/3402/nuffield-advanced-science-book-of-data-second-edition>

Pages 124–5 contain data for a selection of gaseous equilibria, including values for  $\Delta_r H$ . This information can be used to generate questions or could be projected onto the whiteboard for students to consider the effects of pressure and temperature on each equilibrium. The data includes values for each equilibrium constant at several temperatures.

The dynamic nature of equilibrium is probably best introduced by analogy, although visual examples can help students to understand some of the concepts. For example the equilibrium between liquid and vapour in a partially filled bottle of water, where droplets of condensation on the walls of the bottle show the reaction is proceeding in both directions and the stopper emphasises that equilibrium can only be reached in a closed system. The distribution of iodine between aqueous and hydrocarbon solvents provides striking evidence that equilibrium does not mean equality. Simulations can be used to show how the concentrations of reactant and product change as equilibrium is approached.

Simple test-tube experiments show that reactions are reversible under different conditions. For example, when ammonium chloride is heated it decomposes but the products recombine in the cooler conditions at the top of the tube. Equilibria in which there are significant amounts of products and reactants can be investigated practically at the qualitative or quantitative level, providing opportunities to fulfil some of the practical requirements of the course and allow students to determine values for  $K_c$ .

**Effect of Concentration on Equilibrium**

Royal Society of Chemistry: LearnChemistry

<http://www.rsc.org/learn-chemistry/resource/res00001710/an-equilibrium-involving-chromate-vi-and-dichromate-vi-ions>

and

<http://www.rsc.org/learn-chemistry/resource/res00001711/an-equilibrium-involving-copper-ii-ions>

These simple equilibria form a good introduction to reactions which can be reversed by increasing or decreasing reactant concentrations. The reaction of copper(II) ions with ammonia is encountered elsewhere in the specification.

**Le Chatelier's Principle: the effects of concentration and temperature**

Royal Society of Chemistry: LearnChemistry

<http://www.rsc.org/learn-chemistry/resource/res00000001/the-equilibrium-between-two-coloured-cobalt-species>

This is a colourful demonstration of the equilibrium between aqueous cobalt(II) ions and the  $[\text{CoCl}_4]^{2-}$  complex. The colour changes at different temperatures are particularly striking and Le Chatelier's principle can be used to deduce which direction is exothermic and hence which ligand forms stronger bonds.

**Le Chatelier's Principle: the effect of pressure**

Royal Society of Chemistry: LearnChemistry

<http://www.rsc.org/learn-chemistry/resource/res00001728/equilibria-involving-carbon-dioxide-in-aqueous-solution>

This quick and simple demonstration can be carried out in five minutes at the start or end of a lesson. A few  $\text{cm}^3$  of soda water is put in a plastic syringe and the pressure is reduced. The soda water becomes less acidic as  $\text{CO}_2$  leaves the solution. The demonstration also links with work on acidity.

**Le Chatelier's Principle: the effects of temperature and pressure**

Microscale Gas Chemistry

<http://mattson.creighton.edu/NOx/index.html> (Preparation and experiments 6 and 11)

The equilibrium between nitrogen oxides is a textbook favourite and this site describes a safe and simple way in which it can be demonstrated. The reaction is carried out in plastic  $60 \text{ cm}^3$  syringes with Luer-Lok fittings to which gas-tight caps can be attached. A small quantity of nitrogen monoxide is generated in a syringe then reacted with oxygen to produce a mixture of brown  $\text{NO}_2$  and colourless  $\text{N}_2\text{O}_4$ . Changing the pressure or temperature causes the mixture to become darker or lighter as more or less  $\text{NO}_2$  is formed. The demonstration is also described at <http://www.rsc.org/learn-chemistry/resource/res00001739/le-chateliers-principle-the-equilibrium-between-nitrogen-dioxide-and-dinitrogen-tetroxide>, although the method for preparing  $\text{NO}_2$  is more time-consuming.

**Equilibrium and the Chemical Industry**

University of York: Centre for Industry Education Collaboration

<http://www.essentialchemicalindustry.org/processes/catalysis-in-industry.html>

This website provides values for the equilibrium yield of ammonia at different temperatures and pressures and explains how the conditions for the Haber process were chosen. Students can use the site to find out about the reactions involved in the manufacture of nitric acid or sulfuric acid, for example, and see how the conditions used reflect considerations of equilibrium and rate.

**Introducing  $K_c$  – Learner Resource 1**

The worksheet uses data from two different equilibria to develop the relationship between the stoichiometric equation and the formula for the equilibrium constant.

**Determining  $K_c$  for a redox equation**

Royal Society of Chemistry: Microscale Chemistry

<http://www.rsc.org/learn-chemistry/resource/res00000537/measuring-an-equilibrium-constant>

A solution containing  $\text{Ag}^+$  and  $\text{Fe}^{2+}$  ions is left to reach equilibrium and the concentration of silver ions is then determined by titration. Since the initial concentrations are known the amount of  $\text{Fe}^{2+}$  remaining at equilibrium can be deduced, and hence the quantity which has been oxidised. A microscale technique uses  $1 \text{ cm}^3$  aliquots so the experiment can be carried out quickly and relatively cheaply. This is a good example of a heterogeneous equilibrium since one of the products is solid silver.

**Determining  $K_c$  for the hydrolysis of an ester – Learner Resource 2**

Students prepare mixtures containing various amounts of water, ethyl ethanoate, ethanoic acid and ethanol, together with  $\text{HCl}$  catalyst. Equilibrium is approached from various starting points and the mixtures are then titrated with alkali. By subtracting the amount of alkali needed to neutralise the catalyst students can determine the amount of ethanoic acid present at equilibrium and hence deduce the amounts of the other reactants. In this example  $K_c$  has no units but it is good practice for students to convert their values into concentrations before completing the calculation.

<https://www.ocr.org.uk/Images/179794-equilibrium-activity-game-and-pieces-.pdf>

# Learner resource 1

## Introducing the Equilibrium Constant

### Distribution of ammonia between two immiscible solvents

The table below shows the results of an experiment in which trichloromethane ( $\text{CHCl}_3$ ) was mixed with aqueous ammonia solution and shaken. The two solvent layers were allowed to separate and the concentration of ammonia in each layer was then measured. The experiment was repeated with different amounts of ammonia.

Concentration of ammonia in $\text{CHCl}_3$ layer ( $\text{mol dm}^{-3}$ )	Concentration of ammonia in aqueous layer ( $\text{mol dm}^{-3}$ )
0.0574	1.35
0.0429	1.05
0.0348	0.825
0.0261	0.648
0.0208	0.513

- How could the concentration of ammonia be determined?
- The results show that ammonia is more soluble in water than in trichloromethane. Why is ammonia so soluble in water?
- What do you notice about the ratio of the concentrations of ammonia in each layer?

The ratio  $[\text{NH}_3]_{\text{aq}} / [\text{NH}_3]_{\text{CHCl}_3}$  is called the partition coefficient for the equilibrium. It is constant unless the temperature is changed.

### The equilibrium between hydrogen, iodine and hydrogen iodide

Hydrogen gas was mixed with iodine vapour and heated until no more hydrogen iodide was produced. The amounts of each substance at equilibrium were as follows:

$[\text{H}_2]$ $\text{mol dm}^{-3}$	$[\text{I}_2]$ $\text{mol dm}^{-3}$	$[\text{HI}]$ $\text{mol dm}^{-3}$
$4.56 \times 10^{-3}$	$0.74 \times 10^{-3}$	$13.54 \times 10^{-3}$
$3.56 \times 10^{-3}$	$1.25 \times 10^{-3}$	$15.59 \times 10^{-3}$
$2.25 \times 10^{-3}$	$2.34 \times 10^{-3}$	$16.85 \times 10^{-3}$

- Write an equation for the reversible reaction between hydrogen and iodine to make hydrogen iodide.
- What do you notice about the ratio  $[\text{HI}] / [\text{H}_2] [\text{I}_2]$  ?
- What about the ratio  $[\text{HI}]^2 / [\text{H}_2] [\text{I}_2]$  ? Which formula gives the better constant?

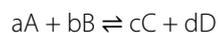
The following results were obtained for an experiment in which hydrogen iodide was heated. It was carried out at the same temperature as the previous experiment.

$[\text{H}_2]$ $\text{mol dm}^{-3}$	$[\text{I}_2]$ $\text{mol dm}^{-3}$	$[\text{HI}]$ $\text{mol dm}^{-3}$
$0.48 \times 10^{-3}$	$0.48 \times 10^{-3}$	$3.53 \times 10^{-3}$
$0.50 \times 10^{-3}$	$0.50 \times 10^{-3}$	$3.66 \times 10^{-3}$
$1.14 \times 10^{-3}$	$1.14 \times 10^{-3}$	$8.41 \times 10^{-3}$

- Can you explain why the concentrations of hydrogen and iodine are the same in each experiment?
- Calculate the ratio  $[\text{HI}]^2 / [\text{H}_2] [\text{I}_2]$ . Do you get similar values to the previous experiment?

### The equilibrium constant

Experiments like these have shown that for an equilibrium



the ratio  $[\text{C}]^c [\text{D}]^d / [\text{A}]^a [\text{B}]^b$  is constant at any particular temperature. It is called the equilibrium constant and has the symbol  $K_c$ .

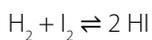
- Explain why the equilibrium constant for the equilibrium  $\text{H}_2 + \text{I}_2 \rightleftharpoons 2 \text{HI}$  has no units.
- Write an expression for the equilibrium constant for the reaction  $\text{N}_2\text{O}_4 \rightleftharpoons 2 \text{NO}_2$
- What units does this equilibrium constant have if the concentrations are in  $\text{mol dm}^{-3}$ ?

## Learner resource 1 Teacher Answers

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Ammonia concentration is determined by titrating aliquots of each layer against standard HCl.

Ammonia molecules are able to form hydrogen bonds with water molecules, making ammonia very soluble in water. The ratio of concentrations in each layer is approximately constant.



The ratio varies from 4.0 to 3.1 in just these three examples. It is clearly not constant.

The ratio is fairly constant at about 54.

The concentrations of hydrogen and iodine are identical in the second set of experiments because they are formed in equal amounts by the decomposition of hydrogen iodide.

The second set of data also gives values for the equilibrium constant of about 54.

$K_c$  has no units when the quantities on the top and the bottom of the expression cancel.

$$K_c = [\text{NO}_2]^2 / [\text{N}_2\text{O}_4]$$

Units of  $K_c$  are  $\text{mol dm}^{-3}$ . (If the equation was written in the opposite direction the units would be  $\text{dm}^3 \text{mol}^{-1}$ .)

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