

## Appendix 2

### 13. Determining the value of $K_c$ for an equilibrium reaction – Student Sheet

In this experiment you have the opportunity to use the skills you have developed by doing volumetric analysis to determine an equilibrium constant.

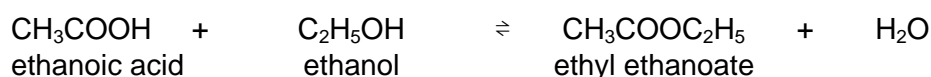
#### Intended lesson outcomes

By the end of this practical you should:

- be able to deduce an expression for the equilibrium constant,  $K_c$ ;
- be able to calculate a value for  $K_c$  from the concentrations of substances at equilibrium;
- be able to draw up a table for the results of your calculations;
- have further developed your skills in titration techniques;
- be aware of the use of a **control** in an experiment.

#### Background information

Ethyl ethanoate is an important solvent. One way of manufacturing it is shown by the following equation.






This is the equilibrium you are going to investigate. It takes a very long time for this reversible reaction to reach equilibrium. An acid catalyst, such as hydrochloric acid is used to speed up both the forward and reverse reactions. This enables the equilibrium to be reached in a much shorter time. The catalyst does not affect the value of  $K_c$ .

#### Question

- 1 What is the expression for  $K_c$  for this equilibrium?

#### Safety

	You must wear eye protection throughout this experiment.
	Glacial ethanoic acid is <b>corrosive</b> . (Mop up any spillages with plenty of water).
	Hydrochloric acid is <b>corrosive</b> .

#### Procedure

1. Your class is going to set up five boiling tubes with different starting concentrations of the various reactants. The volumes of each reactant you need to add are shown in the following table. Take great care with glacial ethanoic acid – it is pure ethanoic acid and is very corrosive.

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boiling tube	volume of liquid/cm <sup>3</sup>				
	HCl (1 mol dm <sup>-3</sup> )	glacial ethanoic acid CH <sub>3</sub> COOH	ethanol C <sub>2</sub> H <sub>5</sub> OH	ethyl ethanoate CH <sub>3</sub> COOC <sub>2</sub> H <sub>5</sub>	H <sub>2</sub> O
A (control)	2	0	0	0	18
B	2	6	6	6	0
C	2	0	6	6	6
D	2	6	0	6	6
E	2	6	6	0	6

- Stopper the boiling tubes, label them and leave for a week, immersed in a bucket of water. This will keep the temperature reasonably constant and allow equilibrium to be achieved.
- After a week has passed, the reactions in tubes B, C and D need **quenching**. This prevents the position of equilibrium from changing any more. Adding water, which makes the reaction mixture very dilute, does this and stops the catalytic effect of the hydrochloric acid.

Take 10.0 cm<sup>3</sup> from each boiling tube using a pipette and pipette filler. Make each up to 100.0 cm<sup>3</sup> in a volumetric flask using deionised water. (Remember to wash the pipette with a little deionised water and then twice with a very small volume of the liquid in the boiling tube.)

Label these flasks B, C and D to correspond to the contents of the original boiling tubes.

- For tube A, pour the entire contents into a 100.0 cm<sup>3</sup> volumetric flask and make up to the 100.0 cm<sup>3</sup> mark. Wash the boiling tube twice with a little deionised water and transfer the washings. Label this flask A.
- From each volumetric flask except A, withdraw 10.0 cm<sup>3</sup> using a clean pipette and pipette filler. Transfer the liquid to a conical flask and add two or three drops of phenolphthalein indicator. Titrate this with 0.01 mol dm<sup>-3</sup> NaOH(aq). Do a rough titration followed by two accurate readings that agree within 0.1 cm<sup>3</sup>.
- Pour the entire contents of volumetric flask A, the solution made up from tube A, into a 250 cm<sup>3</sup> conical flask. Rinse the volumetric flask twice with a little deionised water and transfer the washings to the conical flask. Add phenolphthalein indicator. In this case you are not going to be able to do a rough titration, so add 0.5 cm<sup>3</sup> at a time until the indicator changes colour.

### Processing your results

Although this looks complicated, if you take it by stages it should be straightforward.

#### Stage 1

You need to know how many moles of each reactant you started with. You can calculate this if you know the density of each liquid as mass = volume x density.

	CH <sub>3</sub> COOH	C <sub>2</sub> H <sub>5</sub> OH	CH <sub>3</sub> COOC <sub>2</sub> H <sub>5</sub>	H <sub>2</sub> O
density/ g cm <sup>-3</sup>	1.05	0.79	0.92	1.0

**Worked example** – tube B, ethanoic acid:

If 6.0 cm<sup>3</sup> of ethanoic acid is added, the mass of CH<sub>3</sub>COOH = 6.0 cm<sup>3</sup> x 1.05 g cm<sup>-3</sup> = 6.3 g  
*M<sub>r</sub>* CH<sub>3</sub>COOH = 60, therefore mol CH<sub>3</sub>COOH = 6.3/60 = 0.105 mol

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Now draw up a table to record the starting moles for the four liquids for each of the boiling tubes C, D and E and record the results of your calculations.

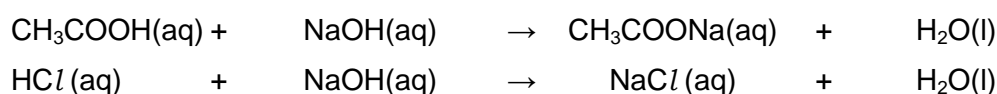
**Stage 2**

Calculate the number of moles of HCl in  $2.0 \text{ cm}^3$   $1 \text{ mol dm}^{-3}$  HCl. In this case you do not need to use density, as we already know its concentration.

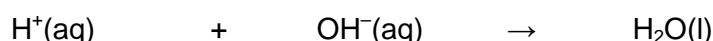
For the contents of tube A, all these moles were titrated against  $0.1 \text{ mol dm}^{-3}$  NaOH. Calculate the volume of NaOH you would expect to add to neutralise the hydrochloric acid present. This should be the actual volume you used because there was only water and HCl in the mixture. This is your control and it is used to prove that the concentration of HCl is not changed by the conditions of the experiment.

**Stage 3**

There are two neutralisation reactions arising in the titrations from volumetric flasks B, C and D. The end point is when both acids are neutralised.



Remember, every neutralisation reaction can be expressed ionically as:



- For each titration result from flasks B, C, D and E, calculate the number of moles of  $\text{H}^+(\text{aq})$  present.
- Before doing the titrations, you diluted the solutions in the boiling tubes. This was to quench the equilibrium mixtures. First of all you only pipetted half the  $20.0 \text{ cm}^3$  of the equilibrium mixture from the boiling tube into the volumetric flask. Then you only transferred a tenth of these moles from the volumetric flask into the conical flask. The result of these two actions means that you must multiply each of your answers in (a) by 20 to give the number of moles  $\text{H}^+(\text{aq})$  in each original equilibrium mixture.
- We know from Stage 2 that the moles of  $\text{H}^+$  due to the catalyst HCl remained unchanged. Now subtract the number of moles of  $\text{H}^+$  in each tube from your answer in (b). You now have the equilibrium moles of ethanoic acid. This amount may be greater or less than the number of moles you started with.

**Stage 4**

At this stage you can calculate the moles of each substance present at equilibrium. We will show how to do this using your results for tube B.

- Moles of ethanoic acid present at equilibrium = x mol** (this is the value from Stage 3)  
These moles have not reacted because they were still in the boiling tube. You started with 0.105 moles of ethanoic acid. This means that  $0.105 - x$  moles have reacted.
- Moles of ethanol at equilibrium =  $0.103 - (0.105 - x)$**

The equation tells us that 1 mole of ethanoic acid reacts with 1 mole of ethanol. So  $0.105 - x$  moles ethanoic acid must have reacted with  $0.105 - x$  ethanol.

Look back at the table you drew up in Stage 1. For tube B you should have calculated a value of 0.103 for the starting moles of ethanol.

Therefore the number of moles of ethanol at equilibrium is the starting value 0.103 moles, minus the number of moles that have reacted,  $0.105 - x$  moles, i.e. moles of ethanol at equilibrium =  $0.103 - (0.105 - x)$

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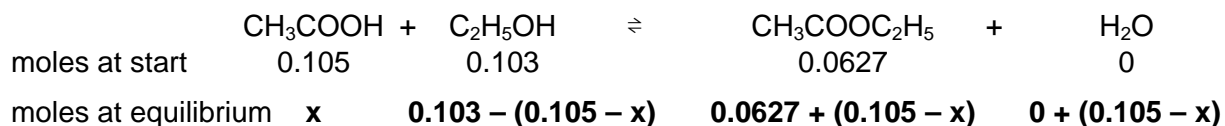
**(c) Moles of ethyl ethanoate at equilibrium = 0.0627 + (0.105 – x)**

We started with 0.0627 moles of ethyl ethanoate in tube B but because 0.105 – x moles ethanoic acid has reacted, 0.105 – x ethyl ethanoate must have been produced. The total ethyl ethanoate at equilibrium is the starting number of moles plus the number of moles produced.

**(d) Moles of water at equilibrium = 0 + (0.105 + x)**

This has increased by the same amount as ethyl ethanoate for the same reason. There was no water to start with in tube B.

In summary:

**Stage 5**

Using the number of moles of each substance at equilibrium in boiling tube B, we can calculate  $K_c$ .

The expression for  $K_c$  requires equilibrium *concentrations*. For the purposes of this calculation, all four substances are present in the same volume ( $20 \text{ cm}^3$ ), therefore these volumes cancel each other out when substituted into the equilibrium expression. So we can use the number of moles. This is not always the case.

Using your answer to Question 1 in the 'Background information' at the start of this sheet, calculate  $K_c$  for tube B.

Work through the stages again to calculate values for  $K_c$  for boiling tubes C, D and E and tabulate your results for equilibrium moles and  $K_c$  for each tube.

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### 13. Determining the value of $K_c$ for an equilibrium reaction – Teachers' Notes

This experiment shows how an equilibrium constant can be measured using techniques students have already learned in their practical course to date.

#### Learning outcomes

These are printed on the Student Sheet.

#### A suggested approach

The Student Sheet gives full details of how to carry out the practical. Setting up the equilibrium mixtures can be done towards the end of a theory lesson about equilibrium constants. With careful preparation, it is safe and need not take too long. Before this lesson, fill burettes with each of the five liquids set out in Table 13.1 on the Student Sheet. This makes handling glacial ethanoic acid much safer for the students. Each student, or group of students, only needs to prepare one mixture, or two at the most. These are left in a bucket of water to try to keep the temperature reasonably constant. If you have access to a thermostatically controlled water bath, so much the better, but this is not essential.

After a week (although 3 days may well suffice), the boiling tube mixtures are at equilibrium, so the practical lesson is about titrating them to determine the equilibrium concentrations. It is only necessary for each group to titrate one of the boiling tube mixtures. As a suggestion, you could always tell students you have titrated tube A for them, since the concentration of hydrochloric acid remains the same in this *control* tube.

#### Technical information

##### Notes

1. We suggest that one student, or group, makes up one or two of the boiling tube mixtures.
2. The table on the Student Sheet gives the volumes for each boiling tube.

Requirements per student/group:

#### Setting up the mixtures

Bucket of water

Thermometer  $-5^{\circ}\text{C}$  to  $50^{\circ}\text{C}$  for putting in the bucket of water

Boiling tubes with bungs

Access to burettes containing:

- deionised water
- glacial ethanoic acid (**corrosive**)
- ethanol
- ethyl ethanoate
- $1\text{ mol dm}^{-3}$  hydrochloric acid (**irritant**)

#### Titration of the equilibrium mixtures

$10.0\text{ cm}^3$  pipette

Pipette filler

$0.01\text{ mol dm}^{-3}$  NaOH(aq)

Phenolphthalein

$100\text{ cm}^3$  conical flask

$100\text{ cm}^3$  volumetric flask

Sheet of white paper or a white tile

Wash bottle and deionised/distilled water

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**Safety**

The main points are included on the Student Sheet but it is the teacher's responsibility to ensure that a full risk assessment is carried out prior to the practical session. MSDS sheets should be consulted so that the correct action can be taken in event of a spillage and/or accident.

You are strongly advised to set up the burette containing glacial ethanoic acid yourself.

**Processing the results**

Although the actual calculations look daunting, they are set out in a very logical way. However, students may still need help in understanding exactly what is happening at each stage. A sample calculation has been done in the case of boiling tube B. Students will be able to see the results of their calculations more clearly if they draw up suitable tables. These should have the appropriate row and column headings, including units. The  $K_c$  value expected is approximately 4.

The equilibrium concentrations of boiling tube mixtures C, D and E are as follows.

**Boiling tube C**

	$\text{CH}_3\text{COOH}$	+	$\text{C}_2\text{H}_5\text{OH}$	$\rightleftharpoons$	$\text{CH}_3\text{COOC}_2\text{H}_5$	+	$\text{H}_2\text{O}$
moles at start	0		0.103		0.0627		0.333
moles at equilibrium	x		$0.103 + x$		$0.0627 - x$		$0.333 - x$

**Boiling tube D**

	$\text{CH}_3\text{COOH}$	+	$\text{C}_2\text{H}_5\text{OH}$	$\rightleftharpoons$	$\text{CH}_3\text{COOC}_2\text{H}_5$	+	$\text{H}_2\text{O}$
moles at start	0.105		0		0.0627		0.333
moles at equilibrium	x		$(0.105 - x)$		$0.0627 + (0.105 - x)$		$0.333 + (0.105 - x)$

**Boiling tube E**

	$\text{CH}_3\text{COOH}$	+	$\text{C}_2\text{H}_5\text{OH}$	$\rightleftharpoons$	$\text{CH}_3\text{COOC}_2\text{H}_5$	+	$\text{H}_2\text{O}$
moles at start	0.105		0.103		0		0.333
moles at equilibrium	x		$0.103 - (0.105 - x)$		$(0.105 - x)$		$0.333 + (0.105 - x)$

**Answer to question on the Student Sheet**

$$1 \quad K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]}$$