Determination of the equilibrium constant for an Esterification reaction

Evaluation of the Experiment (Skill E)

Overall, I feel that the accuracy of the result was quite good. Because the equilibrium constant is a constant, the value of K_c should be the same for each experiment. All of the K_c values calculated for the four experiments came to approximately 6 (no units). The results are summarised below.

	INITIAL VOLUMES (in cm^3)				
Experiment	Ethanoic acid	Ethanol	Ester	Water	K_c
1	75	75	75	25	**3.07**
2	75	75	0	100	6.36
3	35	0	35	180	6.47
4	0	75	75	100	6.15

We can see from these results that although the value of K_c is roughly the same for all of them it is quite small for experiment one. I believe that this is an anomalous result.

There were two main areas of limitation in this experiment. The first concerns the fact that the temperature of the experiment was not reliably regulated. K_c is not affected by a change in concentrations because the concentrations at equilibrium will always remain in the same proportions so K_c remains the same. However K_c does change with temperature. It is difficult to maintain the temperature at room temperature because this is continually changing by small but significant degrees. These slight changes meant that each experiment could potentially have been conducted at a slightly different temperature resulting in different values of K_c.

The second limitation is that the experiment did not take place in a closed system. Sodium hydroxide (NaOH) was added externally. Because the temperature was quite high the reaction was still proceeding. This means that the NaOH neutralised the acid in the equilibrium. Removing the acid caused a shift in the equilibrium to make more acid. This again makes our value of Kc quite unreliable because on adding the NaOH we have actually shifted the equilibrium.

There were two main sources of error of measurement in the experiment. These came from taking measurements using the burette and pipette. Let us deal with the burette to begin with. The burette can be read accurately to ± 0.05cm^3. This gives an overall accuracy of 0.1cm^3. To calculate the percentage uncertainty we have with each volume collected for each experiment, we divide the potential error by the volume and multiply by 100, i.e. for the burette and a volume of 1cm^3 0.1/1 x 100.

The percentage uncertainties for the individual volumes are shown below:

Experiment	Volume (cm3) of NaOH	Percentage uncertainty (%)
1	18.40	0.54
2	15.47	0.65
3	18.09	0.55
4	6.80	1.47

As we can see clearly from the above data, experiments 1, 2, and 3 have quite high volumes of NaOH. These give low percentage uncertainties ranging between 0.54 and 0.65 percent. However for experiment four the volume is much lower, only 6.8cm^3. This gives a percentage uncertainty just less than a percent more than the large volumes. The conclusion is that for larger collected volumes the results are far more reliable, in fact they are nearly 170% more reliable using the burette.
The second source of error in measurement was in the use of the pipette. This was used to gather a 1cm^3 sample of the equilibrium mixture. The pipette is accurate to ± 0.01cm^3 giving an overall accuracy of 0.02cm^3. Using the same formula as before the percentage of uncertainty is:

$$0.02/1 \times 100 = 2\%$$

In conclusion, we can see from the two calculations that using the burette and titrating gives a much lower percentage of uncertainty than using the pipette. This is despite the fact that the pipette on its own is actually more accurate. The fact is that although the pipette is more accurate over one value, it can only measure miniscule volumes and – as we have already discussed – smaller volumes confer higher percentages of uncertainty. Therefore the burette is more reliable (the degree of reliability depending on the volume titrated) but the pipette is more accurate.

In light of the above observation, it would seem logical to increase the volume of the equilibrium reaction used in order to give a more reliable measurement. A suitable measurement might be 10cm^3. However, we cannot simply measure out 10 sets of 1cm^3 samples using the same 1cm^3 pipette as this would multiply out the level of accuracy 10 times. If this were the case, the level of accuracy would decrease from 0.02cm^3 to just 0.2cm^3. In order to ensure that the level of accuracy remains the same we would have to use one pipette that is able to measure 10cm^3 of sample. However, this would have to be very long and thin to ensure reliability. This could be seen as impractical.

On the count of the two other limitations mentioned above, namely the fact that the temperature was difficult to keep constant and that not having a closed system meant that adding NaOH introduced a shift in equilibrium. Both of these limitations can be minimised if not solved by using an ice bath at 0°C to put the sample into.
The temperature of an ice bath at this temperature is very easy to regulate. If the temperature rises slightly then more ice can be added to lower it again. In this way the temperature is kept constant for all of the experiments.
The fact that the experiment is not conducted in a closed environment means that the equilibrium shift cannot be completely stopped but it can be minimised by using the ice-bath. If we cool the experiment from room temperature quickly to the temperature of the ice then the reaction will be frozen at the equilibrium at room temperature – it will still be moving but at a much reduced rate which can be considered negligible. This means that titrating with the sodium hydroxide will still neutralise the acid but because the reaction is moving so slowly it will not have time to move towards producing more acid and so the concentrations will be the same as those of the equilibrium at room temperature giving a more reliable value for K_c