Aim:
My aim in this investigation is to outline whether the acidity of a bottle of vinegar (ethanoic acid) is labelled correctly.

Prediction:
If the vinegar bottle is labelled correctly at 5% acidity then I predict that the amount of NaOH needed for the titration will be:

Ethanoic acid + Sodium hydroxide → Sodium ethonate + Water

$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow NaCH_3COOH_{(aq)} + H_2O_{(l)}$

5% acidity of vinegar (ethanoic acid) with a volume of $25cm^3$ will be used.
$0.1moldm^{-3}$ of NaOH will be used to neutralise the ethanoic acid.
Therefore:

1. We can work out the cm^3 of the ethanoic acid in $25cm^3$ of vinegar
   $\frac{5\%}{100} \times 25cm^3 = 1.25cm^3$

2. We can then work out the grams of ethanoic acid followed by the moles of ethanoic acid.
   $\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}}$

   Density of ethanoic acid is equal to $1.049g/cm^3$ (reference bibliography 1)
   Therefore $1.25cm^3 \times 1.049 = 1.31125$ this is the mass of the CH_3COOH
   Molar mass of $CH_3COOH = 12+(3X1)+12+16+16+1 = 60$
   Therefore:
   Moles of $CH_3COOH = \frac{1.31125}{60} = 0.02185$ moles
3. Because the ratio of ethanoic acid to sodium hydroxide is 1:1 the moles for both are equal.
   NaOH = 0.02185 moles

4. $\text{Concentration} = \frac{\text{Moles}}{\text{Volume}}$
   $0.1moldm^{-3} = \frac{0.02185}{V}$
   $\text{Volume of NaOH} = \frac{0.02185}{0.1} = 0.21854dm^3$
   $= 218.54cm^3$
5. Because we diluted the vinegar by the factor of 10 we will need to take this in consideration in our NaOH therefore:
   $\frac{218.54}{10} = 0.021854dm^3$
   $= 21.854cm^3$

Therefore I predict that the volume of NaOH needed to titrate the ethanoic acid will be $0.021854dm^3$, which is the same as $21.854cm^3$.

By calculating the mount of NaOH needed for the titration we will be able to foresee whether the titration will be feasible in a school laboratory. Looking at

the volume of NaOH needed from our calculations we can see that such a titration can be carried out.

Justification:

The need to carry out a titration between sodium hydroxide and hydrochloric acid was because we do not have an accurate molarity of sodium hydro xide and by carrying out a titration we will have a more reliable set of results.

A volumetric flask was used because the percentage error of the measurements taken from it if the measurement is taken from the bottom of the meniscus in the appropriate way, the percentage error of using a volumetric flask is 0.08%

A Burette was used because the burette lets out small drops, which have a volume of approximately $0.05cm^3$, this allows you to add small quantities of sodium hydroxide to the vinegar to reach a m ore reasonable end point. An error of 1 drop with volume of $25.00cm^3$ would lead to a percentage error of 0.2% for each reading taken. The percentage error is more reliable if the burette is correctly and efficiently conditioned with the titrant.

A pipette was used because it is a much more accurate way of measuring out solution than that of a measuring cylinder also if it is allowed to drain of all solution attained in it then the percentage error collates to approximately 0.24%. The percentage error is more reliable taken in account that the pipette is properly and adequately conditioned with the vinegar in this case.

Phenolphthalein indicator was used instead of the other selection indicators available to us because, we are titrating primarily vinegar (ethanoic acid) and sodium hydroxide, vinegar is a weak acid whilst sodium hydroxide is a strong base. Therefore we are looking for an indicator which will change colour within the period of pH where the amount of hydrogen ions are the same as the hydroxide ions, this point is called the equivalence point. The graph in figure 1.1 shows pH changes for a titration of weak acid and strong alkali. It also illustrates that the reaction between a weak acid and strong acid will be at the equivalence point from pH7 to about pH10 therefore we need an indicator which changes colour between these pH's. Looking at the table in figure 1.2 where the values of where the indicators change colour show that the most appropriate indicator will be phenolphthalein.

The vinegar is diluted by a factor of 10 because of two reasons; firstly because we need to match the lower concentration of the sodium hydroxide, as a high concentration of sodium hydroxide will react with the burette itself

therefore a lower concentration is used. Then secondly because most vinegar solutions are a dark colour and we will be unable to see the colour change of the indicator, therefore by diluting the vinegar we find that the colour is made more translucent and easier to spot a colour change of the indicator in.

The white tile is placed beneath the conical flask because the white colour will give a better background upon which the slight pink colour change in the indicator can be seen.

Method:

To carry out your titration you must:

METHOD – 1:

Before we can begin our titration we should be aware of the fact that the moles of sodium hydroxide is not accurate but an approximation, therefore to make our results more reliable and accurate we must titrate sodium hydroxide and hydrochloric acid. Therefore method one shows how to titrate the sodium hydroxide with Hydrochloric acid.

1. Before beginning your titration, you must condition your burette with the solution to be titrated in this case 0.100moldm $^{-3}$ Hydrochloric acid (HCl). To condition the glass wear rinse the burette firstly with distilled water and drain then fill the burette with a small portion of 0.100moldm $^{-3}$ HCl. Expel into the sink. Fill again with 0.100moldm $^{-3}$ Hydrochloric acid and expel in the sink. Fill the burette with Hydrochloric acid, check the tip of the burette for an air bubble, remove air bubbles by opening the tap slightly.
2. Take an initial volume reading from the burette, read the bottom of the meniscus. Be sure your eye is at the level of meniscus, not above or below. Reading from an angle, rather than straight on, results in a conceptual error.
3. Condition a pipette the same way in which you conditioned the burette but this time once with distilled water and twice with 0.1moldm $^{-3}$ Sodium Hydroxide (NaOH)
4. Measure out 25cm^3 of 0.1moldm^{-3} NaOH using the pipette into a conical flask.
5. Add two drops of phenolphthalein indicator into the conical flask with the dilute vinegar.
6. Place a white tile beneath the conical flask, so that the change in the indicators colour will be seen more evidently.

7. Slowly release the tap of the burette releasing a slow and steady stream of Hydrochloric acid into the conical flask. Stir the conical flask mixture every so often so as to be aware of the colour change of the indicator.
8. Stop adding Hydrochloric acid at the point when you see a slight shade of pink in the conical flask.
9. Take the reading of the burette, (read the bottom of the meniscus) to the nearest 0.05cm^3 and calculate the volume of the Hydrochloric acid needed to neutralise the Sodium Hydroxide.

Use your first set of results as a Trial and then obtain 3 sets of results, which are within 0.1cm^3 of each other. This will increase the results reliability and give us a better average.

METHOD – 2:

This is the method for titrating Ethanoic acid / vinegar with sodium hydroxide (NaOH).

1. Before beginning your titration, you must condition your burette with the solution to be titrated in this case 1.0$moldm^{-3}$ Sodium Hydroxide (NaOH). To condition the glass wear rinse the burette firstly with distilled water and drain then fill the burette with a small portion of the 0.1$moldm^{-3}$ of sodium hydroxide. Expel into the sink. Fill again with 0.1$moldm^{-3}$ sodium hydroxide and expel in the sink. Fill the burette with the 0.1$moldm^{-3}$ sodium Hydroxide. Check the tip of the burette for an air bubble, remove air bubbles by opening the tap slightly.
2. Take an initial volume reading from the burette, read the bottom of the meniscus. Be sure your eye is at the level of meniscus, not above or below. Reading from an angle, rather than straight on, results in a conceptual error.
3. Condition a pipette the same way in which you conditioned the burette but this time once with distilled water and twice with the 5% acidity vinegar. Measure 25cm^3 of 5% vinegar using a pipette into a 250cm^3 volumetric flask, fill the remaining quantity of the flask with distilled water, this will dilute the vinegar by a factor of 10.
4. Condition a pipette the same way in which you conditioned it the first time but this time with the diluted vinegar. (Once with distilled water and twice with the diluted vinegar)
5. Measure out 25cm^3 of dilute ethanoic acid/vinegar using the pipette into a conical flask.
6. Add two drops of phenolphthalein indicator into the conical flask with the dilute vinegar.
7. Place a white tile beneath the conical flask.
8. Slowly release the tap of the burette releasing a slow and steady stream of 0.1$moldm^{-3}$ of sodium hydroxide into the conical flask. Stir the conical flask mixture every so often so as to be aware of the colour change of the indicator.
9. Stop adding 0.1$moldm^{-3}$ of sodium hydroxide at the point when you see a slight shade of pink in the conical flask.
10. Take the reading of the burette, (read the bottom of the meniscus) to the nearest 0.05cm^3 and calculate the volume of the sodium hydroxide needed to neutralise the vinegar.

The first experiment will be your trial then you will need to obtain 3 sets of results, which are within 0.1cm^3 of each other. This will increase the results reliability and give us a better average.

Risk Assessment:

Sodium Hydroxide can cause severe burns, therefore the bottle of 0.1moles should be labelled IRRITANT as it can cause irritations to the skin, if a spill onto the skin occurs place skin under running water and wash thoroughly if large area is affected or blistering occurs seek medical advice. Goggles must be worn to prevent any of the hydroxide entering the eye area, if solution does splash into the eye then flood the eye with gentle running water for 10 minutes then seek medical advice. If solution is swallowed wash out mouth and give a glass or two of water. Do not induce vomiting and seek medical attention as soon as possible.

Ethanoic Acid is harmful when in contact with the skin, wash the area infected with large quantities of wate r. If the acid splashes into the eye then wash the eye with gentle running water and seek medical attention, for a prevention method goggles should be worn and to protect the skin as well as clothing a lab coat should be worn at all times. If the acid is s wallowed then wash out the mouth and give a glass or two of water and seek medical attention as soon as possible. Do not induce vomiting.

Hydrochloric acid is corrosive. The vapour is very irritating to the respiratory system, therefore if vapour is inh aled taken victim to fresh air to rest if the breathing is affected at all seek medical attention. If the solution is swallowed then the mouth should be washed out and a glass of water should be taken, do not induce vomiting and seek medical attention. If the solution is spilt onto the skin it can be very irritating therefore the area affected should be washed with large quantities of water and if any blistering occurs then seek medical attention, to prevent any eye contact and skin contact with the solutio n both goggles and lab coat should be worn at all times.

Phenolphthalein is a toxic substance when in contact with skin or when swallowed. If however it is swallowed wash out the mouth and give a glass or two of water and seek medical attention as soon as possible. It can cause burns and prolonged contact can irritate the skin resulting in dermatitis. If skin contact does occur add small amounts of water, which may increase absorption. Flood area with water for at least 15 minutes. If available swab repeatedly with polyethylene glycol or glycerol, then with soap and water. Phenol burns are very serious so seek medical attention as soon as possible. If it is splashed into the eye then wash eye out thoroughly with gentle running water for 10minutes and seek m edical attention to prevent this wear a lab coat and goggles.

If either of the solutions are spilt in the lab then they should be mopped up wearing gloves as soon as possible as they can corrode the equipment as well as a wet floor being dangerous for an yone moving around in the lab. Make sure that the person responsible or supervising is aware of any problems that occur and seek their first hand help.

Analysis:

Results for HCl and NaOH titration:

	Trial	1	2	3	4	Average volume of HCl added
Starting Volume Of HCl	$0.30cm^3$	$0.10cm^3$	$0.60cm^3$	$0.35cm^3$	$15.50cm^3$	
Final volume of HCl	$26.30cm^3$	$25.90cm^3$	$26.00cm^3$	$25.70cm^3$	$40.95cm^3$	
Volume of HCl added	$26.00cm^3$	$25.80cm^3$	$25.40cm^3$	$25.35cm^3$	$25.45cm^3$	$25.40cm^3$

When calculating the average amount of Hydrochloric acid added into the NaOH to make it neutral, the results which were concordant were only used therefore the result from trial one was not included in the average as it was not concordant with the other three results.

We can work out the concentration of the NaOH by:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Ratios of moles: 1 : 1 : 1 : 1
Concentration: $0.100moldm^{-3}$ unknown
Volume : $25.40cm^3$ $25.00cm^3$

To find moles of known Hydrochloric acid:
Use the formula: **Concentration = Moles / Volume (in dm^3)**
Rearrange formula to make moles the subject:
Moles = Concentration X Volume (dm^3)
To convert cm^3 into dm^3 divide by 1000. Therefore $25.40cm^3 \div 1000 = 0.0254dm^3$
Substitute values into the formula:
Moles Of HCl = $0.100moldm^{-3} \times 0.0254dm^3$
$= 2.54 \times 10^{-3}mol$
Because of the 1:1 ratio of moles between HCl and NaOH. Moles of NaOH are also equal to $2.54 \times 10^{-3}mol$
Therefore (using formula concentration = Moles $\div$ Volume dm^3)
Concentration of NaOH= $\frac{2.54 \times 10^{-3}mol}{0.025dm^3} = 0.1016moldm^{-3}$
Now we have a more accurate concentration of NaOH which through our calculations is $0.1016moldm^{-3}$.

Results for NaOH and CH_3COOH titration:

	Trial	1	2	3	4	Average volume of HCl added
Starting Volume Of NaOH	$0.10cm^3$	$0.10cm^3$	$0.65cm^3$	$0.75cm^3$	$10.00cm^3$	
Final volume of NaOH	$24.30cm^3$	$22.90cm^3$	$22.35cm^3$	$22.40cm^3$	$31.75cm^3$	
Volume of NaOH added	$24.20cm^3$	$22.80cm^3$	$21.70cm^3$	$21.65cm^3$	$21.75cm^3$	$21.70cm^3$

When calculating the average amount of Sodium Hydroxide added into the NaOH to make it neutral, the results which were concordant were only used

therefore the result from trial one was not included in the average as it was not concordant with the other three results.

We can work out the concentration of ethanoic acid by:

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow NaCH_3COOH_{(aq)} + H_2O_{(l)}$$

Ratio of moles:	1	:	1	:	1	: 1
Concentration:	unknown		$0.1016 moldm^{-3}$			
Volume:	$25cm^3$		$21.70cm^3$			

To find moles of known Sodium Hydroxide:
Use the formula: **Concentration = Moles / Volume (in dm^3)**
Rearrange formula to make moles the subject:
Moles = Concentration X Volume (dm^3) To convert cm^3 into dm^3 divide by 1000.
Substitute values into the formula:
Moles Of NaOH = $0.1016 moldm^{-3}$ X $0.0217dm^3$
= $2.20 \times 10^{-3}mol$
Because of the 1:1 ratio of moles between CH_3COOH and NaOH. Moles of CH_3COOH are also equal to $2.20 \times 10^{-3}mol$
Therefore (using formula concentration = Moles ÷ Volume dm^3)
Concentration of CH_3COOH = $\frac{2.20 \times 10^{-3}mol}{0.025dm^3}$ = $0.0882 moldm^{-3}$
Because we had diluted our vinegar by a factor of ten we have to multiply our result by 10. 0.0882 X 10 = $0.882 moldm^{-3}$
Therefore the concentration of ethanoic acid/vinegar is equal to $0.882 moldm^{-3}$.

Evaluation:

Looking at the results obtained collectively we can see that for both experiments the first set of results after the trial were not concordant with the rest of the results, for this reason I did not include them in my calculated averages, they were disregarded so that they would not influence my calcula tions. A suggestion for why the results were anomalous may be because when carrying out the titration the first few times, we may have gone too far to the end point, and therefore as we gained more experience our results seemed to be more accurate. This is backed up by the fact that the results for our trial were also not concordant with the rest of our results.

We can work out how reliable our results are by calculating the percentage error of the experiment, by comparing the actual result we found and w hat our prediction was, therefore;
I predicted that it would take $0.21854dm^3$ of $0.1 moldm^{-3}$ NaOH to neutralise $25cm^3$.
To find moles of known Sodium Hydroxide:
Moles = Concentration X Volume (dm^3)
Substitute values into the formula:
Moles Of NaOH = $0.1 moldm^{-3}$ X $0.21854dm^3$
= $2.19 \times 10^{-2}mol$
Because of the 1:1 ratio of moles between CH_3COOH and NaOH. Moles of CH_3COOH are also equal to $2.19 \times 10^{-2}mol$
Therefore (using formula concentration = Moles ÷ Volume dm^3)

Concentration of $CH_3COOH = \frac{2.19 \times 10^{-2} mol}{0.025 dm^3} = 0.874 mol dm^{-3}$

Therefore I predicted that the concentration of ethanoic acid/vinegar would be $0.874 mol dm^{-3}$.

Now if we were to calculate the percentage error between the actual result and predicted result we should be able to find out how reliable our results are, therefore using the formula:

Percentage Error = $\frac{\text{Actual result – Predicted Result}}{\text{Actual Result}}$ X 100

Substitute in Values; Percentage error = $\frac{0.882 - 0.874}{0.882} \times 100$ = 0.9% error

This tells us that there is an error of 0.9% in our results, however we are aware of the percentage error associated with each apparatus used within the experiment, therefore we can calculate the procedural error and see if the 0.9% error is due to that:

The pipette was used 4 times within each trial and it had 0.24% error. → 4 X 0.24% = 0.96%
The Burette was used 6 times within each trial and it had 0.2% error. → 6 X 0.2% = 1.2%
The Volumetric flask was used only once in the trial and it had 0.08%error.→ 1 X 0.08%= 0.08%

Total procedural error is equal to 0.96% + 1.2% + 0.08% = 2.24%

Because the procedural error is greater than the percentage error of our results we can say that our results are reliable, and that the vinegar bottle is labelled correctly.

However we can reduce the procedural error and improve the accuracy of our results by using a more accurate method of identifying the end point of the experiment as this was calculated by the naked eye and was not very accurate, we could instead use a calorimeter or a digital pH meter to indicate the end point much more accurately. **Also we could take all readings to 2 decimal places therefore making it more accurate and also maybe even measuring the quantities to large volumes as this would**