Determination of the purity of Sodium Carbonate

Aim:
To determine the percentage purity of a sample of sodium carbonate.

Equation:

Na_2CO_3	+	2HCl	→	2NaCl	+	H_20	+	CO_2
Sodium Carbonate		Hydrochloric Acid		Sodium Chloride		Water		Carbon Dioxide

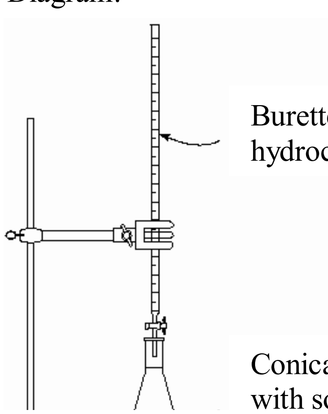
Chemicals:
- 4.00 g dm^{-3} of impure Sodium Carbonate in solution
- 1.00 mol dm^{-3} Hydrochloric acid

Apparatus:
- $50cm^3$ burette — burette clamp and stand
- $25cm^3$ bulb pipette — pipette pump
- Teat pipette
- Conical flask — white tile
- $250cm^3$ graduated flask
- 1 beaker for waste — funnel
- Methyl Orange indicator
- Distilled water

Method:
To determine the percentage purity of a sample of impure sodium carbonate we must first find out how much, in volume, hydrochloric acid it takes to neutralise the solution (this is an acid and base titration and will need an indicator). To do this I will need to adhere to the following method;

1. Clamp the burette carefully. Fill the burette with the hydrochloric acid using the funnel. Remove the air space below the tap and use the wastage beaker to catch any of the acid. Remove the funnel
2. Use the pipette to take $25cm^3$ of the sodium carbonate solution and place it in the conical flask. Add 1-2 drops of methyl orange indicator. Place the conical flask directly under the burette.
3. Take note of the initial reading of the acid in the burette and titrate until the indicator changes from a yellow colour to a pink colour. Take note of the final reading of the acid in the burette.
4. Repeat a minimum of three times, and for each, calculate the volume of acid used by subtracting the initial reading from the final reading.

Diagram:



Burette filled with hydrochloric acid.

Conical flask filled with sodium carbonate

Before I start any titration, I must first dilute the hydrochloric acid to a suitable level

To calculate the volume of hydrochloric acid I will add to the $250cm^3$ graduated flask, I must first calculate the molarity of the sodium carbonate.

For sodium carbonate (Na_2CO_3):

$$\text{Moles} = \frac{\text{Mass}}{\text{RMM}}$$

$$= \frac{4.0}{106}$$

$$= 0.0377 \text{ mol dm}^{-3} \quad \text{(3s.f.)}$$
$$= 0.4 \quad \text{(1s.f.)}$$

From the formula equation I can see that 1 mole of sodium carbonate reacts with 2 moles of hydrochloric acid, therefore the molarity of the acid will be double that of the sodium carbonate.

Concentration of hydrochloric acid = 0.04 x 2 = 0.08 mol dm^{-3}

To make up a solution of $250cm^3$ of 0.08 mol dm^{-3} hydrochloric acid using 1.00 mol dm^{-3} hydrochloric acid I must find out how much of the acid, and therefore how much distilled water, I should add to the graduated flask for a $250cm^3$ solution.

Volume needed using 1.00 mol dm^{-3} moles

$$= \frac{\text{volume x concentration}}{1000} = \frac{\text{V x C}}{1000}$$

$$= \frac{250 \text{ x } 0.08}{1000}$$

$$= 20cm^3$$

According to the above equation, I will need to add $20cm^3$ of hydrochloric acid to a graduated flask that should also contain $230cm^3$ of distilled water to make up the rest of the $250cm^3$ solution of the acid. I should use a burette to add an accurate volume of hydrochloric acid to the graduated flask, and when filling up the flask a teat pipette should be used when making up the final few cm^3 of the solution.

Risk analysis:

There is little risk involved in this experiment as all the solutions used are very dilute and are unlikely to harm anyone if spilled. However, for safety measures, a lab-coat and goggles should be worn and extra care should be taken when using the glassware.

Error Analysis:
The following table shows the error that can occur when reading off the measuring device. I must take these into consideration for my experiment as it could affect my results.

	Volume / cm^3	**Error / cm^3**	**Error per cent / %**
Graduated flask	250	± 0.20	0.08
Bulb pipette	25	± 0.06	0.24
Burette	20	± 0.10	0.50
Pipette	2.5	± 0.04	0.16

The error taken for the burette is as such because I am taking a reading from it twice (initial and final reading), but I am reading to the nearest $0.05cm^3$ to give a greater degree of accuracy. I will also need to take into account the degree of accuracy the titration was, but first I will have to see the volume of hydrochloric acid that will be used in the titration.

To ensure accuracy in this experiment I must:
- Wash all equipment with distilled water and then whatever may be put in it. i.e. wash the burette out with a bit of acid.
- Make sure there are no air bubbles in jet in the burette
- Make sure the jet is below the conical flask
- Make sure that there is no funnel in the burette when titrating
- Read from the bottom of the meniscus
- Record all readings to the nearest $0.05cm^3$
- Make sure there are no air bubbles in the pipette
- Release solution and touch pipette onto surface
- Swirl the conical flask constantly
- Use single drops toward the end of the titration
- Record at least two concordant results after one rough titration

Results:

In the table below, I have tabulated my results for the titrations that I completed. The "**Initial reading**" and "**Final reading**" are indicating the level of hydrochloric acid used in the burette.

	Initial reading / cm^3	**Final reading / cm^3**	**Hydrochloric Acid used / cm^3**
Rough	0.90	17.25	16.85
1	1.15	17.35	16.70
2	0.30	16.55	16.65

3	0.85	17.10	16.60
4	0.55	16.85	16.40

The average of this titration shall be taken from the most similar results. In this case they are: 16.70, 16.65 and 16.60.

$$\text{Average titration} = \frac{16.70 + 16.65 + 16.60}{3}$$

$$= 16.65\text{cm}^3$$

From the average titration I can work out the error percentage of the titration. This works out at 0.60%.

From the average titration volume I will be able to calculate the percentage purity of the sodium carbonate solution.

$$\text{Moles HCl} = \frac{\text{Average HCl used x concentration HCl}}{1000}$$

$$= \frac{16.65 \text{ x } 0.08}{1000}$$

$$= 0.00132$$

$$\text{Moles } Na_2CO_3 = \frac{\text{moles HCl}}{2}$$

$$= 0.00066$$

$$\text{Moles } Na_2CO_3 = \frac{\text{Volume x Concentration}}{1000}$$

$$\text{Concentration } Na_2CO_3 = \frac{\text{Moles } Na_2CO_3 \text{ x}1000}{\text{Volume}}$$

$$= \frac{0.00066 \text{ x } 1000}{250}$$

$$\text{Percentage purity of } Na_2CO_3 = \text{Concentration } Na_2CO_3 \text{ x RFM}$$

$$= 0.0264\text{x } 106$$

$$= 0.27984$$

$$\text{Moles HCl} = \frac{16.65\text{x } 0.08}{1000}$$

$$\therefore \text{Moles } Na_2CO_3 \text{ in } 25\text{cm}^3 = \frac{16.65 \text{ x } 0.08}{1000 \text{ x } 2}$$

$\therefore$ Moles in 250 cm^3 $= \dfrac{16.65 \times 0.08}{1000 \times 2} \times 10$

$\therefore$ Mass Na_2CO_3 in 250 cm^3 $= \dfrac{16.65 \times 0.08 \times 10}{1000 \times 2} \times 106$

$= 0.70596$

$\therefore$ Percentage purity of Na_2CO_3 $= 70.596\%$

I now need to take into account the errors that could have been made.

Overall percentage error $= 0.50 + 0.24 + 0.08 + 0.16 + 0.60 = 1.58\%$

So there is a possible error in my percentage purity of 1.58%,

Possible percentage purity of Na_2CO_3 with error $= 70.596 \pm \dfrac{1.58 \times 100}{70.596}$

$= 73.2\%$ to 68.4%

Therefore the percentage purity of the sodium carbonate solution could range from the upper bound of 73.2%, to the lower bound of 68.4%.