Conductivity on the cheap

The conductivity of a solution is, not surprisingly, a measure of its ability to conduct electricity. The measurement of conductivity is a routine technique in many industrial and environmental applications as a fast, cheap and reliable way of measuring the ionic content of a solution. It is commonly used, for instance, as a method of assessing the effectiveness of water purification systems.

Figure 1 - A traditional conductivity flask.

In the classroom, conductivity is commonly used as a method of distinguishing between ionic and covalent compounds. It is also used as a means of following the course of chemical reactions, especially neutralisation reactions.

The measurement of conductivity of a solution is most commonly done using a conductivity meter. These are fairly expensive and quite fragile.

Another method is to use a conductivity flask. This is a modified conical flask which has an additional hole for a bung in its side. A 2-hole bung is placed in this and electrodes passed through the holes. You then pass an AC current through the circuit (An AC supply is used in order to avoid electrolysis) and measure the resistance.

This is much more robust but the flasks are quite expensive.
We have, however, come up with a simple way you can make your own conductivity apparatus and thus allow the whole class to carry out investigations into conductivity.

Apparatus
- Plastic drinks bottle
- Scissors/modelling knife
- 2-hole bung (to fit the bottle)
- 2 x carbon rods
- Crocodile clips and leads*
- AC power supply*
- Multimeter (capable of reading AC current)*

* You will probably be able to borrow these from the physics department.

Assembling the flask
1) Cut round the drinks bottle about 10 cm from the top and dispose of the bottom half.
2) Push the 2 carbon rods through the holes in the bung so that 2-3 cm or so are protruding. (You may need to use a little petroleum jelly to ease them through).
3) Push the bung into the neck of the bottle. Now is a good time to check it is waterproof by holding it upside down and filling the vessel with water.
4) Use a clamp to grip the neck of the bottle holding your flask upside down.
5) Connect up the circuit with the multimeter and conductivity flask in series.
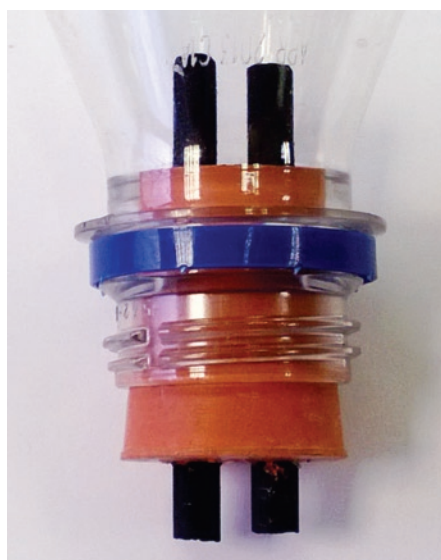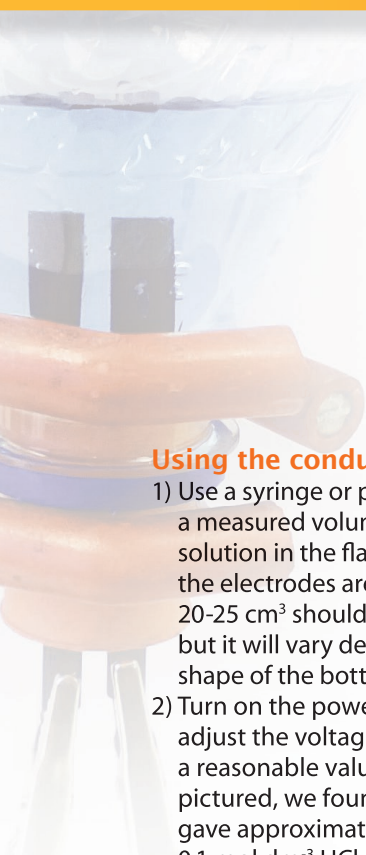6) It is now ready to use.

Figure 2 - Carbon rods pushed through the bung.

Figure 3 - An assembled 'flask'.

Using the conductivity flask

1) Use a syringe or pipette to put a measured volume of the first solution in the flask. (Ensure the electrodes are just covered - 20-25 cm³ should be enough but it will vary depending on the shape of the bottle you are using).
2) Turn on the power supply and adjust the voltage until you get a reasonable value. (For the flask pictured, we found that 4 V gave approximately 0.6 A with 0.1 mol dm⁻³ HCl and 0.3 A with 0.1 mol dm⁻³ NaOH).
3) Record the value before any addition.
4) If you are carrying out a neutralisation reaction, it is interesting, though not essential, to add an indicator.
5) Add the second solution a bit at a time, recording the current after each addition. You can either use a burette or a graduated pipette. The graph (Figure 4) was obtained using a Pasteur pipette.

Results

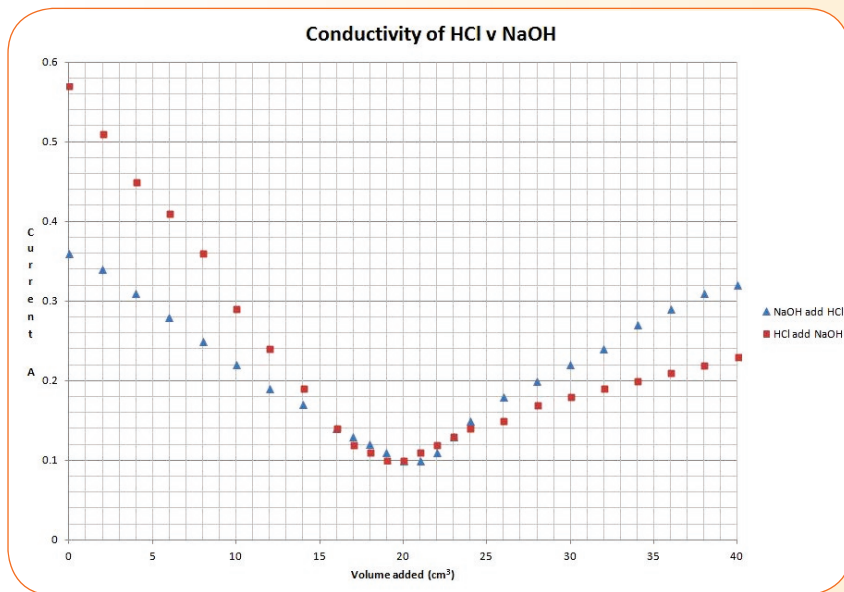The following graphs show sample results obtained with this apparatus.

Figure 4 *- Graph of conductivity in a reaction between a strong acid and a strong base.*

Discussion

Strong acid v strong base

Figure 4 shows the results for a strong acid and base (both 0.1 mol dm⁻³). The graph shows the results both ways round, adding acid to base and base to acid. From the graph it is clear that the acid has significantly higher conductivity than the base. This is due to the small size and hence high mobility of H⁺ ions compared to OH⁻. (See Table 1). So what is happening here? To understand the shape of the graph, you need to know what reaction is taking place. In this case, it can be represented by this equation:

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

At the start of the reaction, you have H⁺ and Cl⁻ ions to conduct electricity. As NaOH is added, the H⁺ ions combine with OH⁻ to form water, leaving Na⁺ and Cl⁻ ions. As Na⁺ is larger and less mobile than H⁺, the conductivity of the solution drops. After the end point, further addition of NaOH results in more Na⁺ and particularly OH⁻ ions being present. OH⁻ ions are less mobile than H⁺ but more than other ions and so the conductivity of the solution increases, but not to the level it was before.
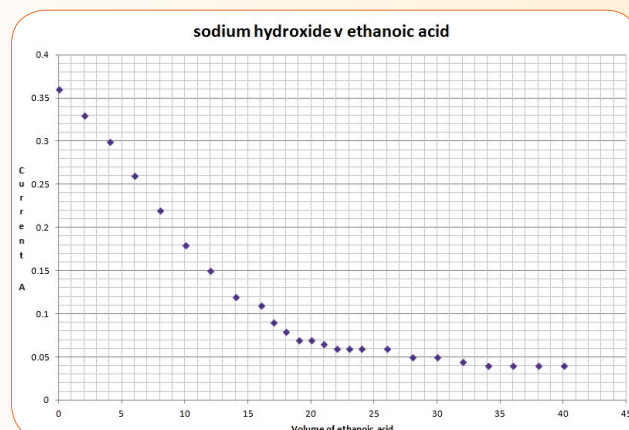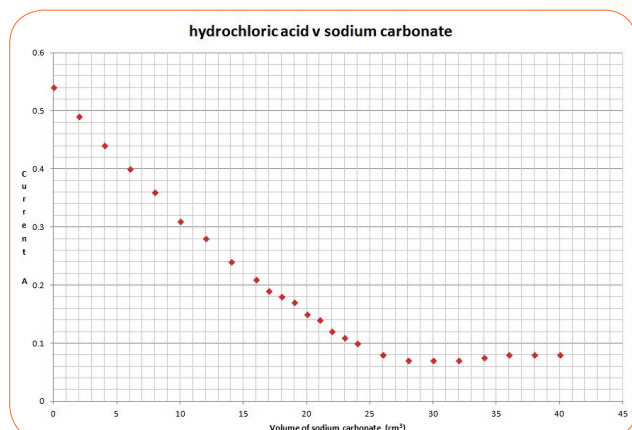

Figure 5 *- Graphs of conductivity in reactions between strong and weak acids and bases.*

If you look at the blue graph (HCl being added to NaOH) you will note that the conductivity at the end is much less than might be expected for that amount of HCl. This is simply the effect of dilution. The conductivity of 20 cm³ of hydrochloric acid is about 0.58 A. The blue plot starts with 20 cm³ of sodium hydroxide and hydrochloric acid is added, ending up with 40 cm³. There is, indeed 20 cm³ of 0.1 mol dm⁻³ hydrochloric acid in the flask now but it is in a total volume of 60 cm³ and so the conductivity is less.

Strong v weak acids and bases
These graphs are less clear and the experiment works best if you start with the strong acid/base and add the weak one rather than the other way round. Because the weak acid/base does not conduct well, the conductivity of the mixture does not rise again after the end point of the reaction but there is still a clear end point to be seen.

Trying this with a weak acid and a weak base gave no worthwhile results.

Precipitation reactions
Any reaction can be followed as long as there is, at some point, a significant change in the conductivity of the solution.
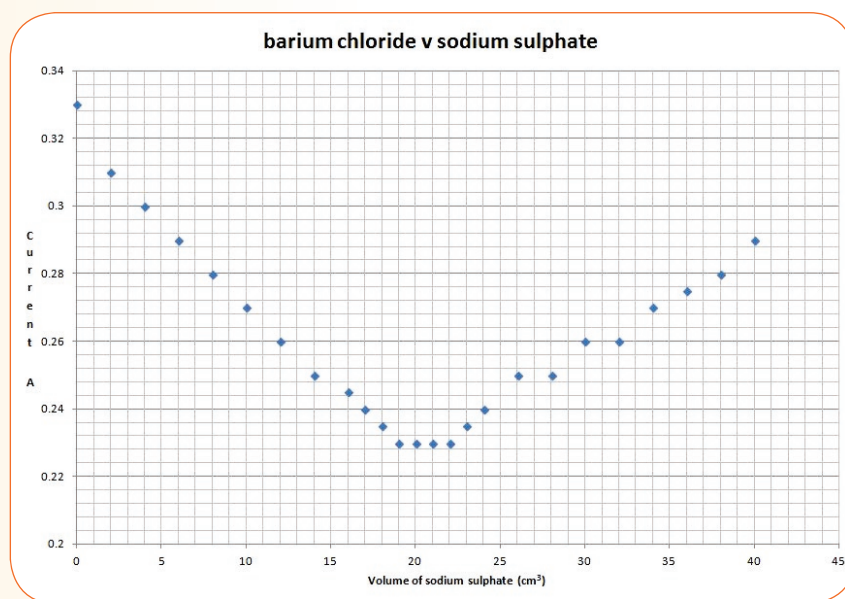

Figure 6 - *Graph of conductivity in a precipitation reaction.*

As well as neutralisation, a good example is precipitation. In this case, one of the products, being insoluble, (barium sulphate in this instance) is removed from the mixture thus reducing the conductivity to a minimum. As excess of the second reactant is added (sodium sulphate here) the conductivity increases again.

Where does it fit in the curriculum? There is no specific mention of conductivity of solutions at any point but, nevertheless, this

apparatus can easily be used to demonstrate the different conductivities of ionic and covalent compounds in solution and also to show how the conductivity depends on concentration (N4 & N5 Chemical Changes and Structure: atomic structure and bonding related to properties of materials).

Using conductivity to follow a reaction as described above could fit into the Advanced Higher Physical Chemistry unit which covers acids, bases and indicators.

The method might find application for the N4/5 assignments and in the Higher/Advanced Higher Researching Chemistry units.

Safety considerations
All the solutions used are of low hazard. The electrical supply should be kept well away from the liquids and should have been PAT tested to ensure its safety. ◄

Ion	Conductivity	Ion	Conductivity
H^+	35.0	OH^-	19.8
K^+	7.36	Br^-	7.84
$NH4^+$	7.34	I^-	7.68
Ag^+	6.38	Cl^-	7.63
Na^+	5.01	$NO3^-$	6.92
Li^+	3.87	$HCO3^-$	4.45

Table 1 - *Molar conductivities of selected ions (10^3 Ω m^2 mol^{-1}).*