

CHEMICAL EDUCATION

## Low-cost experiments on electrolytic conductivity

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A simple and inexpensive system for the measurement of electrolytic conductance is assembled using easily available materials. This system could be used in several student experiments which can facilitate the understanding and learning of such important concepts as ionic mobility, interionic attraction and conductimetric titration.

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**Key Words:** electrolytic conductance; conductivity measurement; low cost experiments.

ELECTRICAL CONDUCTIVITY IS AN IMPORTANT PROPERTY OF ELECTROLYTES. It provides an unequivocal evidence of the existence of ions in electrolyte solutions. This property of electrolyte is usually demonstrated in the laboratory course of university chemistry through a set-up which consists of a pair of electrodes in series with a light bulb and the mains power supply. In spite of the electrical hazard offered by this set-up, it has been widely used because of its simplicity. A recent modification involves the use of a light-emitting diode (LED) and a battery (1-3). The intensity of the light indicates the degree to which the solution is able to conduct electricity.

Quantitative experiments on electrolytic conductance are seldom carried out in the general chemistry laboratory course. The main reason for this has been the limited availability of a conductivity meter due to its high cost. This situation is unfortunate since some important concepts, such as ionic mobility and interionic attraction, can be well understood through the results of these experiments.

A number of low-cost designs for conductivity meters has been previously described in literature (4-6). Most of these meters are black boxes, concealing the complicated electrical circuit and the working principle of the device.

In this paper, an inexpensive and easily constructed system for the measurement of electrolytic conductivity is presented together with some experiments involving the use of this measuring system. A "clear-box" set-up is adopted for the instrumentation in order to foster an appreciation of the principle involved in the measurement.

### Principle of conductance measurement

Electrolytic conductivity arises from the migration of ions under the influence of a potential gradient. The rate of migration depends on the charge and size of the ions, the microscopic viscosity of the medium, and the magnitude of the potential gradient.

The ability of a solution to carry an electric current is expressed in terms of **conductance**,  $G$ , which is defined as the reciprocal of resistance (i.e.  $G = 1/R$ ). The SI unit for this quantity is the siemen, which is equivalent to  $\text{ohm}^{-1}$  or mho. In the presence of various ionic species, the

conductance of a solution is the sum of conductances of the various ions which depend on their number, and their mobilities.

Electrolytic conductance is usually measured by determining the resistance of a solution between two parallel electrodes. Solutions of electrolytes obey Ohm's law. For an applied potential,  $E$ , the current,  $I$ , that flows between the electrodes immersed in electrolyte varies with the resistance,  $R$ , of the electrolyte solution. Since conductance is the reciprocal of resistance,  $G = I/E$ . Thus at a constant potential difference, the intensity of the current flowing between the electrodes is proportional to the conductance of the solution, and may therefore be taken as a measure of the conductance of the solution.

During conductance measurements, the potential drop across the electrodes has to be kept constant. The occurrence of faradic processes or electrolytic reactions in the solution could alter the electrical characteristics of the cell. These changes take place when DC electricity is used. In order to prevent these processes, an alternating potential of a sufficiently high frequency is employed during the measurement.

The measured conductance of a solution is dependent on the dimensions of the electrodes employed. These parameters determine the number of conducting species sensed during the measurement. Consequently, the electrodes used during the measurements should be of a fixed area and kept at a constant distance from each other.

### Construction of the conductance measuring system

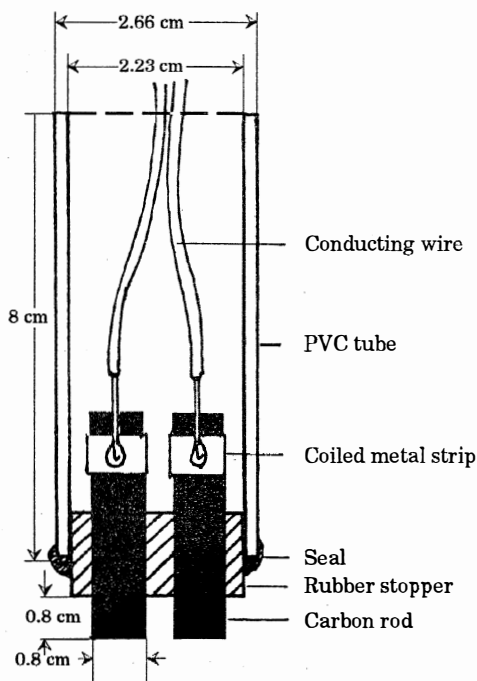
The measuring system consists of a probe containing the electrodes and an electrical circuit for the measurement of the current flowing through the system. The materials needed for the construction of the system are listed in Table 1.

**Procedure.** The conductance probe consists of a pair of carbon rods recovered from a spent Size D dry cell. The carbon rods are cleaned by immersing them in concentrated hydrochloric acid, and washing thoroughly with water.

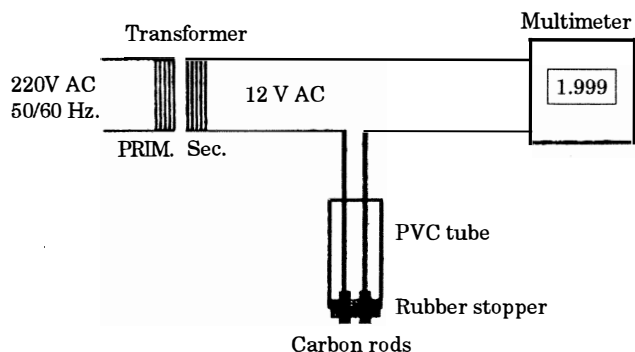
**Table 1.** Materials for the construction of the system for conductance measurement

Probe	Circuit
Carbon rods (from a size D battery)	Transformer (220 V primary/ 12V secondary)
PVC Tube (2 cm o.d.)	AC cord w/ plug
Rubber stopper	AWG #26 wires
Rubber sealant	Digital multimeter
Metal strips (from a paper fastener)	

The dried rods are inserted onto the holes bored about 0.5 cm apart on a rubber stopper which could fit into a PVC tube with an outer diameter of 2 cm. (A piece of the sole of a rubber sandal could be cut to substitute for the stopper.) A metal strip is wound tightly around one end of the rods, and a conducting wire is soldered to the strip. The rubber stopper is then inserted in one end of a 8-cm long PVC tube, with the rods protruding from one end and the connecting wires coming out of the other end of the tube (Fig. 1). The junction of the PVC tube and the rubber stopper is sealed well with a sealant.

**Figure 1.** Diagram of the probe for conductance measurement.

The measuring circuit is diagrammed in Figure 2. The transformer, the probe and the multimeter are connected in series. The connections of the transformer leading to the mains should be wrapped with electrical tape and enclosed in a fabricated plastic or wooden casing. The multimeter is operated in the current-measuring mode.

**Figure 2.** Diagram of the circuit for conductance measurement.

### Experiment 1. Conductance of solutions

Conductance experiments are often employed in the laboratory to demonstrate the difference between an electrolyte and a nonelectrolyte, and between a strong and a weak electrolyte. It can also show the difference in the mobilities of ions in electrolyte solution. All of these can be accomplished by comparing the conductance of the solutions of weak and strong electrolytes, of nonelectrolytes, and of strong electrolytes which differ only in one of the ions.

**Procedure.** Deci-molar (0.1 M) solutions of urea, hydrochloric acid, acetic acid, and sodium chloride are prepared. The conductance of each solution is measured by immersing the conductance probe into the solution, and reading the current displayed on the digital multimeter. The probe should be washed well with distilled water and then dried before immersing it in another solution.

Table 2 shows the multimeter readings, in amperes, obtained when the probe was dipped in 0.10 M solutions of sodium chloride, hydrochloric acid, acetic acid, and urea. The difference between the conductance of hydrochloric acid and acetic acid shows the difference in the extent of ionization of the two acids. The difference in the readings between hydrochloric acid and sodium chloride demonstrates the greater mobility of hydrogen ions in comparison with sodium ions. Among the solutions only that of urea does not conduct, indicating the absence of ions in the solution.

**Table 2.** Effect of nature of solution on conductance.

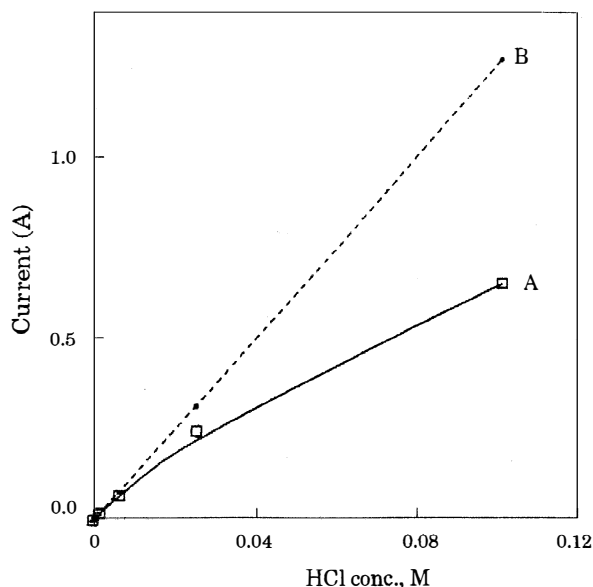
Solution	Current (A)
0.10 M NaCl	0.27
0.10 M HCl	0.67
0.10 M CH <sub>3</sub> COOH	0.02
0.10 M NH <sub>2</sub> CONH <sub>2</sub>	0.00

### Experiment 2. Variation of conductance with concentration

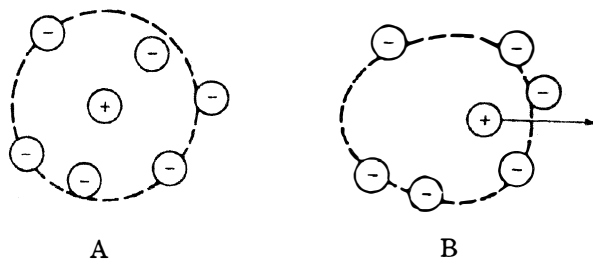
The conductance of an electrolyte solution is dependent on the concentration of the electrolyte. The conducting power of an electrolyte solution is attributed to the movement of the ions in the presence of an electric field. The moving ions carries a portion of the current flowing through the solution. As the number of the ions increases, it is expected that the current flow in the solution should increase correspondingly.

**Procedure.** Several solutions of hydrochloric acid having different concentrations are prepared. The conductance of these solutions are measured in terms of the current flowing in the circuit, as in Experiment 1. The current readings obtained are plotted against the concentration of the solutions.

The plot of the observed current flow against the concentration of the HCl solution is illustrated in Figure 3A. The trend exhibited by the results obtained coincides with the expected behavior. A scrutiny of the data reveals that conductance of a solution of a strong electrolyte does not vary proportionately with its concentration. Figure 3B depicts the plot that would have resulted for a linear relationship between the conductance and the concentration of a strong electrolyte, using the most dilute solution as the reference. The lower conductance values obtained from the concentrated solutions could be explained by the significant interionic attraction occurring in these solutions. In solution, ions are surrounded by an atmosphere of ions of opposite polarity (Fig. 4A). In the presence of an applied potential, ionic migration occurs, with the central ion and



**Figure 3.** (A) Plot of the observed conductance (expressed in terms of current flow) against the concentration of the electrolyte solution. (B) Extrapolated plot of conductance (expressed in terms of current flow) against the concentration of the electrolyte solution, using the most dilute solution as reference and assuming a linear relationship.



**Figure 4.** Diagram illustrating interionic attraction in an electrolyte solution. (A) A positive central ion with an ion atmosphere. (B) A distorted ion atmosphere retarding the mobility of the central ion.

its surrounding ions going in opposite directions. These changes lead to a distortion of the charged atmosphere surrounding the central ion and an unevenly distributed electrostatic attraction (Fig. 4B). This interaction retards the migration of the ions under the influence of an applied voltage and decreases the amount of current conveyed by the ions across the electrodes during the measurement. This behavior is elaborated by the Debye-Huckel theory.

### Experiment 3. Conductimetric acid-base titrations

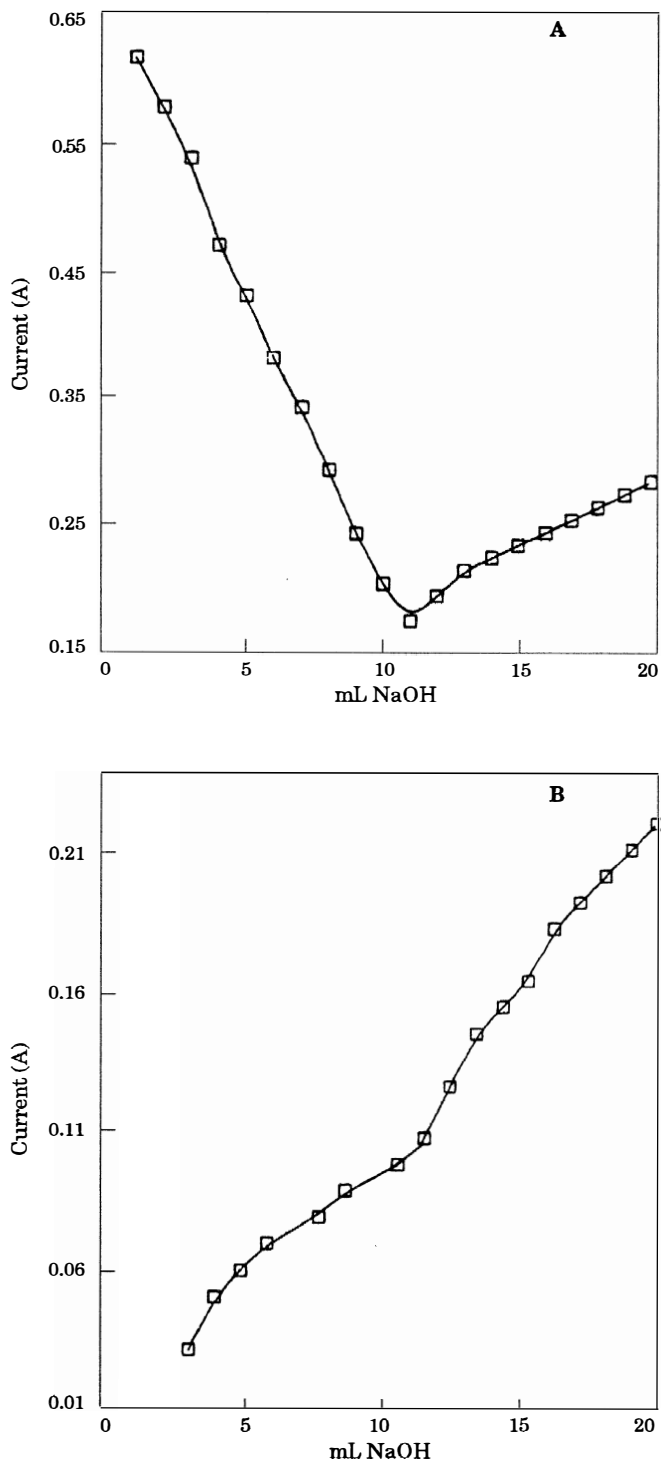
The course of an acid-base reaction can be monitored through the change in conductance caused by the replacement of the highly conducting hydrogen and hydroxide ions by ions with lower conductivity, such as sodium and acetate ions. The progress of an acid-base titration can be monitored through a measurement of the conductance of the titration mixture. The end-point for the titration can be determined graphically. This technique is useful when colored solutions are to be titrated, and dye indicators cannot be used.

**Procedure.** 25 ml of a 0.10 M HCl solution are transferred into a 100 ml beaker. The initial conductance of this solution is measured. The solution is titrated with 1.0 M NaOH, 0.2 ml of the base being added at a time, and the conductance of the mixture is read after each addition. The titration is continued until a total of 3.5 ml of the base has been added. The observed current is plotted against the volume of the titrant added.

The above procedure is repeated for a weak acid such as acetic acid (0.1 M), with 0.1 M NaOH.

Figure 5A shows the curve obtained for the conductimetric titration of a strong acid, such as HCl, and a strong base. During the titration, hydrogen ions are removed from the solution as water which is weakly ionized and is replaced by an equivalent number of the less mobile sodium ions; the conductance of the mixture therefore decreases during the course of the titration. Beyond the equivalence point, further addition of the titrant introduces highly mobile hydroxide ions in the solution. Since no more neutralization can occur, the conductance of the solution increases with the addition of more titrant. The end point is indicated by the sharp break in the titration curve.

The curve obtained for the titration of a weak acid, e.g. acetic acid with a strong base, e.g. sodium hydroxide, is



**Figure 5.** Conductimetric titration curves for (A) a strong acid (HCl) and a strong base (NaOH), and (B) a weak acid (CH<sub>3</sub>COOH) and a strong base (NaOH).

shown in Figure 5B. For this system, the neutralization reaction results in the conversion of the poorly conducting weak acid into a highly ionized salt, leading to an increase in conductivity. Beyond the equivalence point, the conductance changes more steeply because of the addition of highly conducting hydroxide ions into the solution. The break in the curve indicates the endpoint of the titration.

### Conclusion

The conductance of electrolyte solutions can be determined in the general chemistry laboratory through a measuring system which applies Ohm's law. This setup can be easily constructed by the students using easily available materials. The working principle of this measurement can be well understood through this "clear-box" approach to instrumentation.

The measuring system is inexpensive, the total cost of the materials being about P700 (~US\$25). The bulk of this amount is due to the multimeter which costs at least P600 (~US\$22). It should be realized that the investment on a multimeter is worthwhile, since it can be used in several other experiments in the general chemistry laboratory (7,8).

Quantitative experiments on electrolytic conductance enables the study of concepts such as ionic mobility, interionic attraction and conductimetric titrations. Other phenomena which may be investigated are precipitation and dilution effects on ionization. These concepts are simple enough to be taken up in a foundation course in university chemistry.

### References

1. Russo, T. J. *J. Chem. Educ.* 63, 981 (1986).
2. Vitz, E. J. *J. Chem. Educ.* 64, 550 (1987).
3. Gadek, F. J. *J. Chem. Educ.* 64, 628 (1987).
4. Ewing, G. W. *J. Chem. Educ.* 51, A469 (1974).
5. Stork, J. T. *Anal. Chem.* 56, 561A (1984).
6. De la Cuetara, R. A. and Lamba, R. S. *J. Chem. Educ.* 6(1), 33 (1989).
7. Andres, R. T. and Sevilla, F. III. *J. Chem. Educ.*, in press.
8. Sevilla, F. III, Alfonso, R. L. and Andres, R. T. *J. Chem. Educ.*, in press.