The Construction and Characterization of a Conductivity Meter for Use in High School and Undergraduate Science Labs

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Abstract: The focus of this report is the design, construction and evaluation of a low-cost and portable conductivity meter. The instrument has facilitated the introduction and application of conductivity and resistivity to groups of middle school, high school, and undergraduate students. A series of laboratory experiments that demonstrate the utility and capabilities of the conductivity meter have also been developed and tested in multiple contexts. These lab experiments not only help characterize the conductivity meter, but also feature challenging and thought-provoking problems that allow students to realize the limitations of the instrument and place the data in context. The laboratory exercises a) investigate the effect of ion concentration, valence electrons, and temperature on conductivity measurements; b) conductometrically monitor and interpret a titration between a weak acid with a strong base; and c) determine the total dissolved solid (TDS) values for water collected at local creeks and dams on-site.

Introduction

Access to analytical instruments as demonstration and laboratory teaching aids greatly expands the breadth of chemical concepts that can be taught in a classroom and lab environment and allows a more in-depth exposure of those important topics. Unfortunately, the associated costs of maintaining state-of-the-art instructional facilities are increasing while the amount of available funding is shrinking. Such barriers affect curricular programs at all levels from high school to college. However, it is possible to provide rugged, low-cost, low-maintenance, and low-power instruments capable of providing accurate quantitative information for a fraction of the cost of commercial instructional and research grade instrumentation. The basic premise of the Small, Mobile Instruments for Laboratory Enhancement (SMILE) program we have developed is that constructing, tuning, and using instruments, is an invaluable component of understanding science and engineering concepts [1–7]. SMILE brings together aspects of physics, chemistry, mathematics and electronics in a novel way to create an exciting collection of interdisciplinary STEM modules that allow students to build, use, optimize, and adapt small mobile lab instruments. Data that we have collected over the past fourteen years strongly indicates that designing, constructing and testing an analytical instrument facilitates deep conceptual learning in the undergraduate instrumental analysis course [4]. The electronic instruments created in this course are then used in lower-level general and analytical chemistry labs, or donated to high schools and middle schools for use in science class. The latter also brings a sustainable service-learning component to the curriculum that builds connections with pre-college teachers, and helps promote the STEM subjects to their school students, thus improving the educational experience for all [8]. The SMILE program aims to create low-cost, custom-built instruments that facilitate the practical application of chromatography, electrochemistry, and spectroscopy within standard middle school, high school, and undergraduate general and analytical chemistry laboratory courses. The students learn that it is possible, with just a basic understanding of electronics and instrument design theory, to obtain high-quality data from an instrument that they themselves have built and optimized, all within a few hours. The only substantial differences between a commercial instrument and the custom-built instrument are the software, the specialty electronic components, and sensitivity; the basic design and methods for probing chemical systems are identical.

Details regarding construction and characterization of a student-built conductivity meter are reported herein. Electrical conductivity is a measure of how much electrical current a substance can conduct, and is measured using a conductivity cell that determines the electrical resistance. The simplest type of cell consists of two similar electrodes, in which an AC voltage is applied to one of the electrodes causing the ions in solution to migrate towards the electrodes; the more ions, the greater the current that flows between the electrodes. The conductivity meter measures the current and makes use of Ohm's law to calculate the conductance and conductivity of the solution. A number of apparatus for measuring qualitative or semi-quantitative electrical conductivity have been reported in the literature [9–17]. Inspired by some of the positive features in these reports, we have designed a low-cost portable alternative to commercial electrical conductivity meters. The student-built handheld conductivity meter was constructed for less than $45, and includes a case, transformer, plugs and cables, and associated electronic circuitry. The instrument is durable in construction, easy to operate, and ideal for routine

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measurements (Figure 1). In order to fully characterize the instrument, our students developed several lab experiments, and compared their results with data obtained from a commercially available conductivity probe. The lab constructed conductivity meter has provided instrumental analysis students with a practical understanding of electronics, and the labs have helped students realize the usefulness and the critical limitations of analytical instruments.

The device is relatively simple in design and construction, such that high school students, participating in a six-week summer research program sponsored by Summer Experience in the Eberly College of Science (SEECoS) in collaboration with Upward Bound Math and Science, have built and used the device as part of a water quality study. The SEECoS program promotes educational opportunities for low-income students by helping them overcome class, social, and cultural barriers to higher education. The program specifically targets low B and C level students, a group of students who typically do not consider college as an option upon graduation from high school. At the end of the six-week summer enrichment program, each team presents the results of their research in an oral-format symposium. Several teams that we have mentored in the SEECoS program over the years have received first-place in their division and, in some cases first-place overall. Evaluations are based on the quality of the student group's presentations, in-depth knowledge of the project and material, and the ability to answer questions about the research conducted and any issues they may have encountered.

Recently, students in the instrumental analysis course donated fifteen of the conductivity meters to an NSF-funded international environmental awareness program (NSF CNH:0909447). The ReBUild initiative was a cooperative study between American and Ghanaian researchers to identify correlations between disease outbreaks in Ghana and changes in the local environment and ecology: The project primarily focused on linkages between land disturbance through mining and logging, severe rainfall events, and outbreaks of an aggressive tropical skin disease called Buruli ulcer. As part of the outreach component of that grant, researchers in collaboration with the Center for Science and the Schools at Penn State facilitated professional development workshops that enabled Ghanaian school teachers in Subin, Pokukrom, Dunkwa, and Tarkwa districts to support their students in classroom research projects.

Experimental

Instrument Design and Construction. The detailed circuit diagram for the conductivity meter is provided in the Supplementary Information. The circuit board was created and etched in-house, and the electrical components were soldered to the board by students. The meter consists of an amplitude adjustable triangle wave generator, a variable gain trans-impedance amplifier, a comparator synched to the generator (to achieve a DC output), an offset null, and a final low-pass filter stage before sending the signal that is displayed on an LCD. All measurements in this study employed a 150 mV p-p, 2 kHz triangle wave. The cell probe is constructed with two 0.1 mm-thick, 1 × 4 cm stainless steel strips separated by a Teflon block. The exposed electrode surface area is 1 × 1 cm and the electrodes are spaced 1 cm apart. Two leads of stranded wire are soldered to the electrode surfaces and directly connected to the meter via alligator clips. The electrode assembly, except for the active surface area, is wrapped in black insulation tape. A complete list of components is provided in the Supplementary Information. The student-built conductivity meter takes about 45 mins to construct from kit form, an exercise that school and college students can readily and safely perform under supervision.

General Procedures. Three experiments of varying complexity were developed for the student-built conductivity meter. Each of the experiments can be completed within a standard three-hour lab period. Data generated by our students are presented below. Overall, results were found to be quite comparable to the commercial Vernier probe (Vernier, CON-BTA), both in reproducibility and accuracy.

Results and Discussion

Experiment 1: The Effect of Ion Concentration, Valence Electrons, and Temperature on Conductivity. The conductivity of solutions can give important insights into the nature of the solutions and the particles dissolved: Five solutions each of NaCl, $CaCl_2$, and $AlCl_3$ were prepared in Nanopure water (0.01, 0.005, 0.0025, 0.00125, and 0.000625M), and the conductivity was measured using the student-built meter and also a Vernier conductivity probe. Measurements were obtained by pouring 40 ml of each solution into a 50 ml beaker and submerging the probe to a depth of one centimeter (i.e. the active surface area). This consistency in method was maintained for each test sample. Results are illustrated in Figure 2 for the student-built and Figure 3 for the Vernier commercial instrument. There is clearly a direct relationship between salt concentration and conductivity for all three salt solutions. Additionally, the slopes for each of the external calibration curves are proportional to the stoichiometric number of ions generated upon dissolution of the salts. The ratio of the slopes should theoretically be 4:3:2 ($AlCl_3$:$CaCl_2$:NaCl); our meter measured a ratio of 5.1:3.5:2, while the Vernier probe measured the ratio to be 6:3.6:1.

In a related experiment, the effect of temperature on solution conductivity was explored. A solution of $AlCl_3$ was prepared at 0.00125M and the conductivity was measured at 24, 35, 45, 55, 65, 75, and 85°C. Figure 4 illustrates the linear relationship between temperature and conductivity; conductivity measurements are temperature-dependent, as the temperature increases, conductivity increases.

Experiment 2a: Conductometric titration of a strong acid with a strong base. In a conductometric titration the conductivity of the reaction mixture is continuously monitored as one reactant is added to the other. In Figure 5, a 10 mL 0.06 M HNO_3 solution is titrated with 0.0881 M NaOH. The NaOH solution was standardized using a potassium hydrogen phthalate primary standard. Initially, the conductance is high because the strong acid completely dissociates. As NaOH is continually added, water is formed by the reaction $H^+ + OH^- \leftrightarrow H_2O$, which decreases the conductivity. The conductivity reaches a minimum at the equivalence point (6.8 mL of NaOH), and then increases as excess OH^- ions are added upon further addition of NaOH titrant.

Experiment 2b: Conductometric titration of a weak acid with a strong base. In another experiment, as show in Figure 6, a 15 mL ~0.1 M acetic acid solution is titrated with the 0.0881M NaOH. The solution conductivity at the start of the titration is low, because of the weak dissociation of acetic acid.

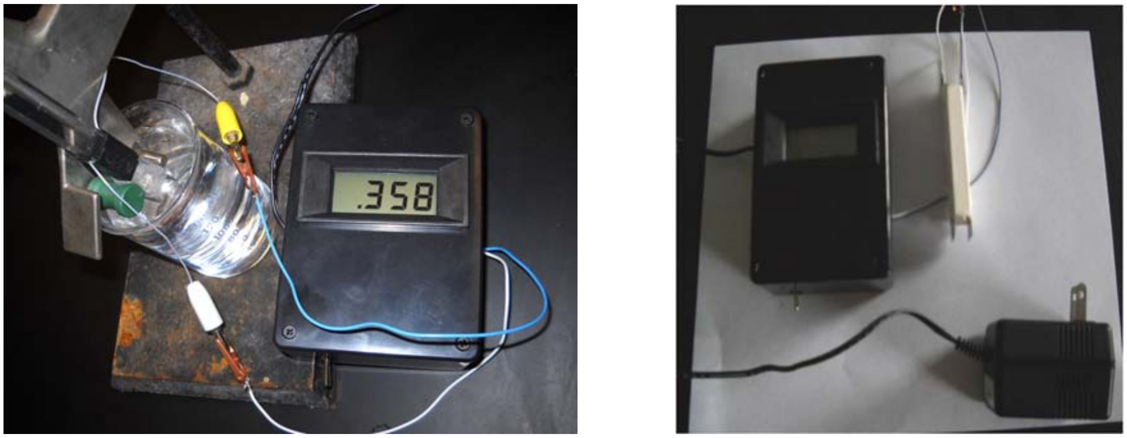

Figure 1. Student constructed conductivity meter.

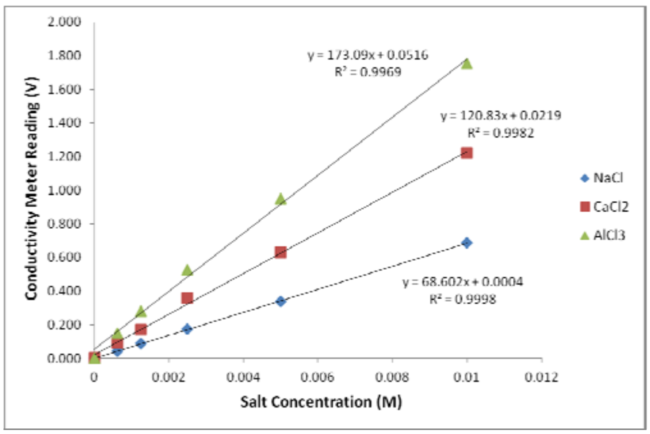

Figure 2. Conductivity measurement of salts using the student-built EC meter. The theoretical ratio of slopes was 4:3:2; the values reported above are 5.1:3.5:2. Do note that these readings are in V whereas the commercial instrument output is in μS/cm.

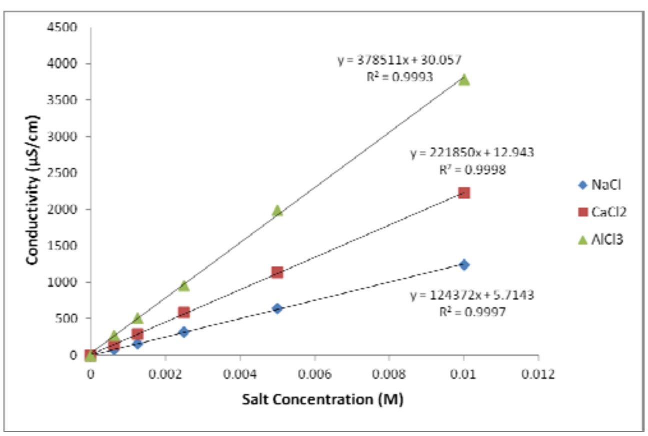

Figure 3. Conductivity measurement of salts using the commercial Vernier instrument. The theoretical ratio of slopes was 4:3:2; the values reported above are 6:3.6:1.

As the titration proceeds, the conductance increases almost linearly to the equivalence point. Beyond the equivalence point, the conductance increases rapidly with excess titrant ions. The slope of the titration curve changes only slightly at the equivalence point making it difficult to read directly, however the plots from the student-built and the commercial instrument are very similar. The data clearly demonstrates that the student-built instrument can detect small changes in conductivity, producing accurate titration curves comparable to the commercial conductivity probe.

Experiment #3: Conductivity as a Measure of Water Quality The concentration of the total dissolved solids (TDS) in a water sample is one measure of its potability and quality. TDS accounts for the total amount of mobile charged ions in water, including minerals, organic matter and inorganic salts. The principal constituents are usually the cations calcium, magnesium, sodium and potassium, and the anions carbonate, bicarbonate, chloride, sulfate and, particularly in groundwater, nitrate. TDS is expressed in units of mg/L or ppm. The U.S. EPA and WHO classify TDS as a secondary contaminant and has set the maximum level for drinking water at 500 ppm. At or above this level, hardness, mineral deposition, corrosion, colored water or even a salty taste to the water become noticeable. Although not harmful, the unusual smell, taste and color may be undesirable.

The standard method in determining TDS involves evaporating a measured sample of water to dryness at 180°C and then carefully weighing the amount of dry solids that remain [18]. Conductivity measurements offer a quicker and easier way of determining TDS, by simply measuring the conductivity, and then using a conversion factor to obtain a TDS value. The student-built conductivity meter was used to generate TDS values for various local bodies of water, as collated in Table 1. One site examined was Spring Creek, a limestone stream in central Pennsylvania that is popular for its wild trout. Water samples were collected and measured for their conductivity on-site by plugging the conductivity meter into a 12 V DC/AC adapter in a car. The samples were then transported in sealed vials and re-measured in the lab to identify any differences associated with the temperature change.

The LCD output from the conductivity meter displays the signal in volts. The reading can be converted to μS/cm and a TDS value through a series of calculations based on the circuit specifications: First, the meter measures the current between the electrodes and converts this cell current into a voltage based on the selected gain setting (i.e. Ohm's Law). For example, we obtained a reading of 1.957 V at low gain for a 0.01 M KCl standard solution. The corresponding current, $I = V_{LCD}/(\text{Gain Setting}, \Omega) = (1.957\ \text{V})/(10{,}200\ \Omega) = 1.92 \times 10^{-4}$ A. The cell current is related to the solution impedance, or resistance, and the voltage (150 mV p-p) of the applied 2 kHz triangle wave, $R = (0.150\ \text{V})/(1.92 \times 10^{-4}\text{A}) = 782\ \Omega$.

Table 1. EC meter readings and corresponding TDS values for several local bodies of water

	Field Measurement		Lab Measurement	
	EC (V)	TDS (ppm)	EC (V)	TDS (ppm)
Shavers Creek 1	0.042	27	0.042	27
Shavers Creek 2	0.081	53	0.078	51
Whipple Dam Upper	0.019	12	0.022	14
Whipple Dam Lower	0.020	13	0.023	15
Spring Creek	0.263	172	0.264	173

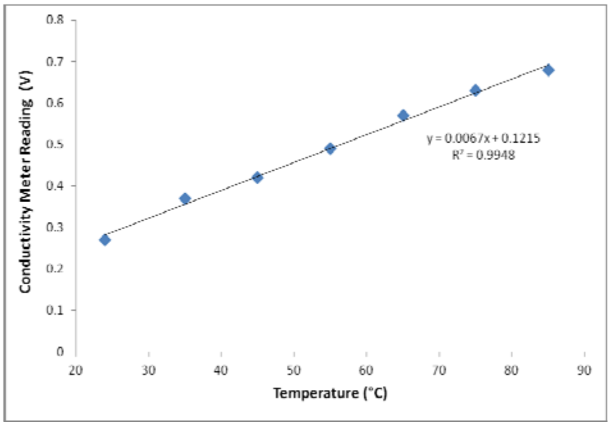

Figure 4. $AlCl_3$ solution was prepared at 0.00125 M and measured at seven different temperatures. Data was obtained using the student-built EC meter.

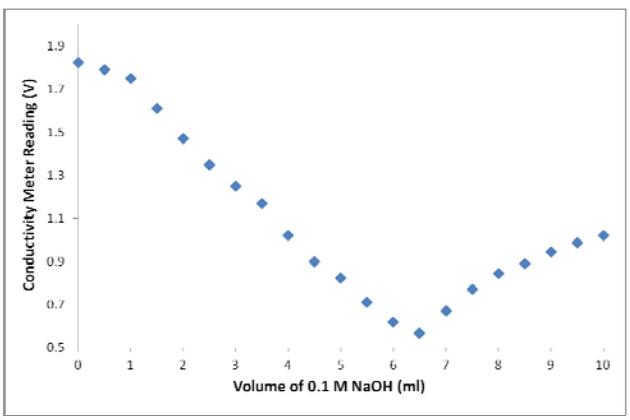

Figure 5. Conductometric titration of 0.06 M HNO_3 with 0.0881 M NaOH.

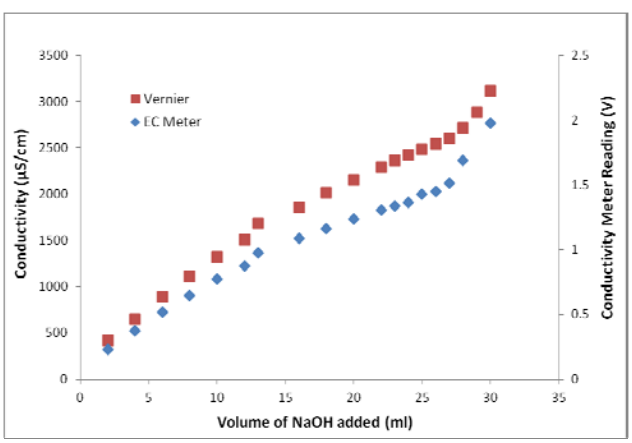

Figure 6. Conductometric titration curves for an approximate 0.1 M acetic acid solution with 0.0881 M NaOH, measured using the student-built EC meter and the commercial Vernier instrument.

Conductance, G, is equal to 1/R; in this case G = (1/782 Ω) = 1.279×10^{-3} Siemens (or 1279 µS). Specific conductivity, C = G*(the cell constant), where the cell constant is equal to the distance between electrodes divided by the surface area of the electrodes. The cell constant for the student-built conductivity probe is 1 cm^{-1}; therefore C = (1.279×10^{-3} S)*(1 cm^{-1}) = 1.279×10^{-3} S/cm or 1279 µS/cm. The stated specific conductivity for the 0.01 M KCl solution is 1413 µS/cm at 25°C.

Converting specific conductance to TDS depends on the nature and quantity of dissolved salts: TDS values derived from conductivity is not recommended for critical quantitative purposes, because there is no relationship between conductivity and TDS that is suitable across different locations and different dissolved material. For natural samples [19], an average conversion factor of 0.640 is used for specific conductance values below 5000 µS/cm, and an average conversion factor of 0.800 for µS/cm values above 5000 in these studies. The actual multiplier is dependent on the activity on the relative amounts of each species, the total concentration of dissolve solids in the sample, and temperature; relationships which can be non-linear. However, if measurements will always be made at the same location, then using a method that measures TDS of samples gravimetrically [20] can be correlated against the measured specific conductance of the samples in order to determine a more precise conversion factor.

For calibration and verification purposes, a plot of the specific conductance values from the student-built meter versus a series of fresh standard solutions of known conductance was conducted and is presented in Figure 7. Ideally, such a plot should have a slope of 1, and the data obtained for the student-built instrument highlights a very strong correlation. Deviations from 1:1 can arise for a number of reasons: (1) the value of the gain setting is taken as the stated value of the resistor but the resistor has a ±1% tolerance (Note: choosing a different nominal value would not affect the slope), (2) inconsistent immersion depth of the conductivity probe, (3) non-ideal solution behavior at higher concentration, and (4) dilution errors, among others.

Conclusion

We have successfully designed and constructed a low-cost alternative to commercial electric conductivity meters. The device is relatively simple to fabricate and has been successfully assembled, tested, and used by middle school, high school and college undergraduate students, thus allowing for a much deeper understanding of the electronic gadgetry and inner workings of such an analytical instrument that are often hidden in 'black-boxes' and misunderstood by students. The conductivity meter described is stable and accurate enough for quantitative experiments, and its performance was validated in

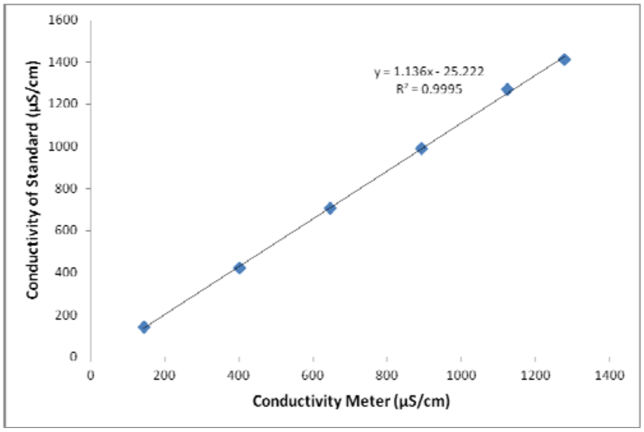

Figure 7. A plot illustrating the specific conductance values calculated from the student-built EC meter reading *versus* a series of standard solutions of known specific conductance. Standard solutions were prepared by successive dilution of a traceable standard of 0.01 M KCl (C = 1413 µS/cm ±1% at 25°C; Hach Company, cat no. C20C270).

comparison to a commercially available probe and standard solutions. Together with the experimental labs, the conductivity meter makes it easier to effectively teach important chemistry, physics, and environmental concepts.

The three labs described are low in cost, engaging, and highly rewarding experiments for students; each of which can be tailored according to student ability level and course objectives and outcomes. Over the past few years we have modified lesson plans to reach a wide range of student abilities, from middle school general science students all the way through to upper-level undergraduate instrumental analysis courses. Our analytical chemistry courses offers students research opportunities, which not only attracts and retains STEM majors, but also improves classroom performance. The SMILE program was intentionally designed to provide undergraduate academic research challenges and opportunities, but we have found that it also serves as a tried and tested mechanism for student engagement and retention.

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