
LABORATORY REPORT

GROUP NUMBER: W3

EXPERIMENT NUMBER: 2

TITLE: Conductivity of Electrolyte Solutions

DATE SUBMITTED: Friday, February 4, 2000

OVERALL REPORT GRADE 7.5

ROLE ASSIGNMENTS

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ABSTRACT

Equivalent conductances of K_2SO_4 were determined at concentrations ranging between $7.010E-3 \pm 4.000E-5$ M and $1.001E-4 \pm 5.000E-7$ M using an electrode with 0.1 nominal constant. A graph of equivalent conductance versus the square root of concentration was plotted. Regression analysis showed that at concentrations from $1.001E-5$ M to $2.00E-3$ M, the plot is linear ($R^2 = 0.999$). The y-intercept of the regression curve, which corresponds to the equivalent conductance at infinite dilutions (λ_0), is found to be $144.42 \text{ S}\cdot\text{cm}^2/\text{mol}$. WHAT ARE ITS CONFIDENCE LIMITS? This value is significantly different from the literature value of $153.5 \text{ S}\cdot\text{cm}^2/\text{mol}$ at the 95% confidence interval. Such result shows that data obtained from this lab are not significant. THIS LAST SENTENCE TELLS NOTHING. SAY NOTHING OR BETTER TELL WHY YOUR VALUE DIFFERS FROM THE LITERATURE.

GRADE FOR SUMMARY PAGE 2.0

RESULTS

Results for the lab conducted are listed and explained below.

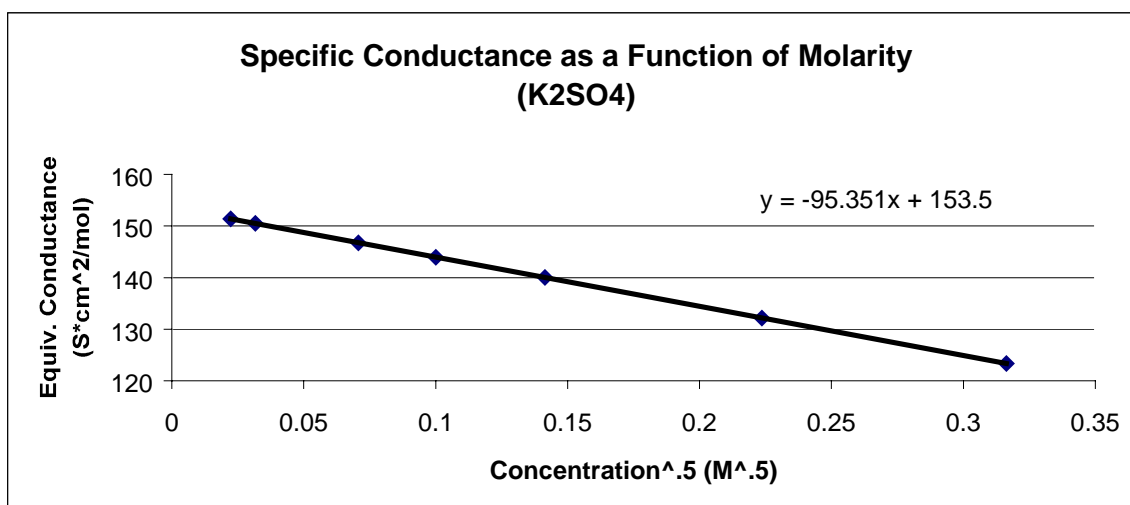
Table 1. Measured Values of Cell Constant (K) for 2 electrodes

Electrode (nominal constant)	Measured K value
1 ±5%	1.000 ±0.003
0.1 ±5%	0.1022 ±0.0003

Calibration of the 1.0 and 0.1 electrodes revealed that the measured values differed from the nominal values by 0.04% and 2.2%, respectively. This was well within the ±5% deviation acceptable for this apparatus.

GOOD!

Figure 1. Linear Kohlrausch plot (based on literature values): Λ vs. $C^{1/2}$ of K_2SO_4



This graph was generated from literature data¹ for ideal values WHAT DOES THAT MEAN? to construct a linear relationship between equivalent conductance and the square root of concentration. This relationship is modeled by the equation $y = -95.351x + 153.5$. The y-intercept value for this graph gives the literature value of Λ_0 , 153.5 S*cm²/mol. WHY DID YOU NOT JUST GET THE VALUE FOR Λ_0 FROM THE LITERATURE?

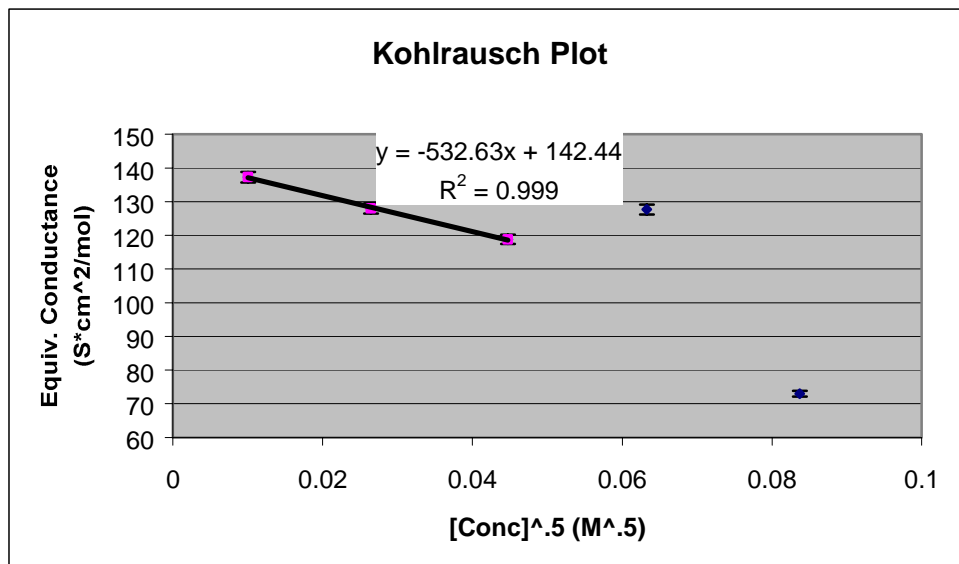
Table 2. Equivalent Conductance of K_2SO_4 with Varying Concentration.

Concentration (M)	Equivalent Conductance λ (S*CM ² /mol)
0.0001001 ± 0.0000005	137 ± 2
0.000701 ± 0.000004	128 ± 2
0.00200 ± 0.00001	119 ± 1
0.00400 ± 0.00002	128 ± 1
0.00701 ± 0.00004	73 ± 2

¹ "Equivalent Conductance of Some Electrolytes in Aqueous Solution at 25° C." *CRC Handbook of Chemistry and Physics*. 63d Edition, 1982-1983. p 5-89

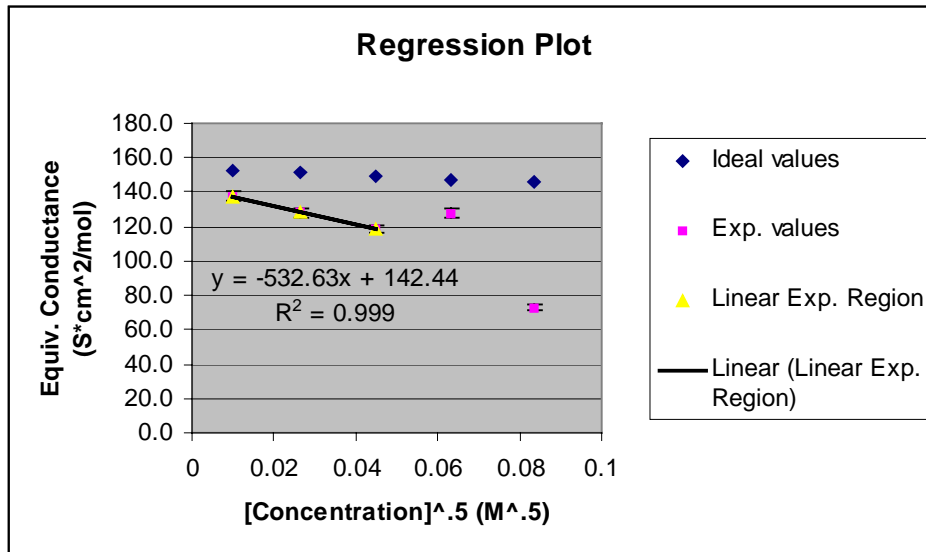
Experimental Equivalent Conductance was calculated for each concentration using the equation $\Lambda = K / (CRq)$, where K = cell constant, C equals the concentration in molarity, R equals resistance (Ω), and q equals the charge carried by the compound, which is 2 for K_2SO_4 . The plot for this data is shown below (Figure 2).

Figure 2. Kohlrausch plot based on experimental values (K_2SO_4)



Equivalent conductance decreases linearly ($R^2 = 0.999$) with the square root of concentration in the region from 0.01 to 0.0447 $M^{1/2}$. Kohlrausch's law is applicable over our first three data points, which had concentrations of $0.0001001 \pm 0.0000005M$, $0.000701 \pm 0.000004M$, and $0.00200 \pm 0.00001M$, respectively. When concentration increases above 0.0002M, Kohlrausch's law is no longer applicable for K_2SO_4 . The uncertainty in the concentration is $\pm 0.5\%$ and the uncertainty in the equivalent conductance is $\pm 1.17\%$. YOU CANNOT MAKE SUCH A CONCLUSION WITH THE DATA SHOWN. DID YOU USE JUST 3 POINTS? WHERE ARE THE REST OF YOUR EXPERIMENTAL POINTS? YOU CANNOT DECIDE ON SUCH THINGS BY LOOKING AT A GRAPH AND FINDING 3 POINTS THAT LINE UP. THERE WILL ALSO BE SCATTER AND IT IS JUST AS UNLIKELY THAT POINTS LINE UP EXACTLY AS THAT THEY DON'T. YOU MUST PLOT ALL YOURS POINTS AND THEN USE REGRESSION ANALYSIS TO DECIDE IF ANY ARE SO FAR OFF THAT THEY MUST BE REPEATED. THIS SHOULD HAVE BEEN DONE IN THE LAB.

Figure 3. Regression Plot – Ideal vs. Experimental Results for K_2SO_4



THERE IS NO IDEAL- IT IS LITERATURE EXPERIMENTAL DATA. SINCE THE LITERATURE DATA INDICATE THAT YOUR DATA ARE ALL IN THE LINEAR RANGE, YOU MUST USE ALL YOUR DATA- IN THIS CASE YOU SHOULD HAVE REPEATED THE POINT AROUND .06 SINCE IT SEEMS TO BE TOO HIGH.

Theoretically, WHAT THEORY STATES THIS? the experimental values of the equivalent conductance should not differ significantly from the ideal values - they should fall within the acceptable BY WHAT CRITERION? uncertainty range of +1.17%. As seen through the graph this is not the case as all the experimental values are less than that of the ideal. See Table 3 for a quantitative treatment of the error. Sources that lead to this uncertainty are listed in Table 4a and 4b.

According to the experimental regression line, the value of Λ_0 obtained is 142.44 $S \cdot cm^2/mol$. The upper bound of the 95% confidence interval is 149.06 and the lower bound is 135.83. Therefore, the literature value, $\Lambda_0 = 153.53 S \cdot cm^2/mol$, is not contained in the confidence interval, and the experimental value is significantly different. The experimental value of Λ_0 is 7.8% lower than the literature value. YOU FOCUS ONLY ON THE LIMITING VALUE BUT LOOKING AT YOUR GRAPH SHOWS THAT ALL THE POINTS ARE DIFFERENT FROM THE LITERATURE. BY THINKING THIS WAY YOU MAY HAVE A CHANCE TO EXPLAIN WHAT IS INVOLVED.

Table 3. Percentage Errors of Experimental Values

Concentration (M)	Percentage Deviation from Ideal (%)
0.0001001 ± 0.0000005	-9.9
0.000701 ± 0.000004	-15.2
0.00200 ± 0.00001	-20.4
0.00400 ± 0.00002	-13.4
0.00701 ± 0.00004	-49.9

This table is a quantitative analysis of the percentage deviation from ideal of the experimental data. Deviation was derived by comparing the experimental values to the ideal values in Figure 3. NOTE THAT ALL ARE NEGATIVE, BUT ALSO, SINCE THE SLOPES ARE VERY DIFFERENT IT INDICATES THAT THIS IS NOT A RANDOM SITUATION, IN WHICH CASE SOME WOULD HAVE BEEN POSITIVE.

Table 4a. Uncertainty Values for Concentration

Apparatus Used	Uncertainty
Mettler H2 mass balance	± 0.008 %
100 mL volumetric flask	± 0.008 %
10 mL pipette	± 0.5%

Practical Uncertainty Value for Concentration: ± 0.5%

Table 4b. Uncertainty Values for Equivalent Conductance

Apparatus Used	Uncertainty
Potentiometer (resistance box)	± 1 %
Resistor	± 0.1 %
Uncertainty for Concentration	± 0.5%
Electrolytic Cell	± 0.34%

Practical Uncertainty Value for Equivalent Conductance: ± 1.17%

Tables 4a and 4b tabulates the uncertainties inherent in the apparatuses that were used to conduct this experiment.

YOUR RESULTS ARE PRESENTED WELL . BUT- WHERE ARE THE CALIBRATION VALUES FOR THE CELL? WHERE ARE VALUES AT HIGHER CONCENTRATIONS WHICH WOULD GIVE SOME IDEA OF THE SOURCE OF THE NON RANDOM ERRORS THAT OBVIOUSLY ARE INVOLVED?

GRADE FOR RESULTS SECTION 3.5

ANALYSIS

The experimental values for equivalent conductance listed in Table 2 were analyzed to determine whether or not they were significantly different than the values predicted by the ideal linear relationship for infinite dilution. This was done by graphing the results on the same graph and using error bars for the experimental values to

determine whether or not the points fell on the line for ideal values (Figure 3). The first three points, having concentrations of $0.0001001 \pm 0.0000005\text{M}$, $0.000701 \pm 0.000004\text{M}$, and $0.00200 \pm 0.00001\text{M}$, respectively, are expected to be not significantly different from the ideal line. This is so because since the points form a linear curve and are thus part of the range over which Kohlrausch's law is applicable. THIS DOES NOT FOLLOW AT ALL. Results show that the most concentrated solution deviates the most from the ideal (-WHAT IS THIS IDEAL?49.9% at $0.00701 \pm 0.00004\text{M}$ as opposed to -9.9% at a concentration of $0.0001001 \pm 0.0000005\text{M}$). This is expected because the most concentrated dilution is the farthest from infinite dilution and thus Kohlrausch's law does not apply. AT SUCH CONCENTRATIONS K'SLAW ISNOT A FACTOR. YOUR VALUE IS EXPECTED TO BE JUST AS GOOD AS THE LITERAURE- IF IT ISNT, LINEARITY HAS NOTHING TO DO WITH IT,AND YOU SHOULD ENDEAVOR TO DISCOVER WHY THEY DIFFER. Percentage errors for the equivalent conductances compared to the actual are listed in Table 3.

As shown in the data, all the experimental values obtained for equivalent conductance at different concentrations are below that of the ideal. There are many reasons as to why the results obtained do not exactly reflect the literature values. There are multiple sources where the deviation from the literature values may have originated. YOUSAIID THAT First, the electrodes were difficult to clean when moving them from solution to solution. Because of their unique design, it is difficult to completely rinse off the ions of one solution before submerging the electrode into another. DID YOU TEST THIS IDEA, SAY BY MOVING THE ELECTRODE BETWEEN SOLUTIONS AND MEASURING FOR DIFFERENT DEGREES OF RINSING? IF NOT, STATEMENTS LIKE THAT THAT YOU DREAM UP INYOUR HEADARE NOT ADMISSIBLE. Also, the distilled water in which the electrodes were initially stored may ???????be contaminated if the electrode is not thoroughly cleaned before it is re-submerged, which can result in ionization of the distilled water. To avoid such sources of error, the electrode should each be separately stored in individual beakers of DI water.DID YOU TRY IT? In addition, it should be rinsed with DI both before and after each use. Another source of error is the method by which the data was taken. The Wheatstone Bridge, which was used to quiet the otherwise noisy resistance reading of a given solution, was

connected to a resistance box. It is difficult to duplicate the resistance in the solution with the resistance box because of the error associated with finding the minimum voltage output of the circuit. Using the electrode with the smaller cell constant should give more accurate resistance readings at very small concentrations. However, in both solutions tested, the 1.0 electrode gave resistance readings that were closer to the ideal than the readings obtained using the 0.1 electrode. For instance, in the 0.01 KCl solution, the 1.0 electrode gave an error of only 0.04%, whereas the 0.1 electrode had an error of 5.54% from the ideal (see Appendix for data). Thus, despite the theoretical advantage of the electrode with the lower cell constant, it was limited by error associated with the resistance box ($\pm 1\%$). ?????? Therefore, the decision of the group to use the 0.1 electrode based on its theoretical advantage resulted in increased error due to the practical limitations of the apparatus. One way to eliminate such error is to use the 1.0 electrode when dealing with a resistance box at these levels of resistances. DID YOU USE THAT ELECTRODE WITH YOUR SOLUTIONS TO SEE?

COMMENTS SUCH AS THOSE YOU GIVE ABOVE ARE COUNTERPRODUCTIVE, NOT SCIENTIFIC, AND PURE BULL. THEY MUST BE RESISTED AT ALL COSTS. NO COMMENT OTHER THAN WHAT YOU OR OTHERS HAVE SHOWN IN THE LABORATORY WITH THE SETUP IN QUESTION WILL BE ACCEPTED. IN THIS CASE YOU MUST LEARN WHAT NOT TO DO.

A third source of error is the creation and dilution of the K_2SO_4 solution. A method of improving the dilution of the solution would be to minimize the serial dilution error by utilizing a mass balance to create each diluted solution. NOT ACCEPTABLE UNLESS YOU CAN SHOW HOW THE UNCERTAINTY IN CONCENTRATION COULD BE IMPROVED TO MAKE A DIFFERENCE IN THE CONDUCTIVITY VALUES.

In conclusion, in this experiment, equivalent conductivity for the compound K_2SO_4 was determined at different concentrations of K_2SO_4 . A Kohlrausch graph was then constructed, plotting equivalent conductance versus the square root of concentration. A regression line was then constructed based on the linear portion of the curve and the equivalent conductance at infinite dilution (Λ_0) was found. Experimental data for (Λ_0) is

144.42 S*cm²/mol, which is significantly less than the literature value of 153.5 S*cm²/mol.

The experimental and educational objectives were both met. Group members measured the equivalent conductance of a soluble salt as a function of concentration and at infinite dilution. In addition, they were also introduced to the Wheatstone bridge and conductivity measurements.

GRADE FOR ANALYSIS SECTION 2.0

REFERENCES

1. "Equivalent Conductance of Some Electrolytes in Aqueous Solution at 25° C." *CRC Handbook of Chemistry and Physics*. 63d Edition, 1982-1983. p 5-89
2. Litt, Mitchell. Conductivity of Electrolyte Solutions. *Bioengineering Laboratory Manual*. pp. 1-10, 2000

APPENDIX

Figure 1. Determination of the Cell Constant K using KCl solution

Electrode (nominal constant)	Concentration of KCl (M)	Resistance Measured (Ω)	Resistance (Ω) (Literature Value)	Cell Constant K determined
1.0	0.01	708	707.714	1.000404
1.0	0.001	6850	6847.261	1.000404
0.1	0.01	110	707.714	0.1272
0.1	0.001	3700	6847.261	0.10223

At a smaller dilution, the smaller K is more accurate for reading. Thus the 0.1 nominal constant electrode was picked to assay the conductance of K_2SO_4 .

Figure 2. Resistance Measured at Varying Concentration of K_2SO_4

Concentration (M)	Resistance Measured (Ω)
0.0001001 ± 0.0000005	3720
0.000701 ± 0.000004	570
0.00200 ± 0.00001	215
0.00400 ± 0.00002	100
0.00701 ± 0.00004	100