Experiment 4: Electrical Conductivity of Aqueous Solutions

Reading: Chapter sections 4.1, 4.3, 4.5 and 4.6 in your textbook and this lab handout.

Ongoing Learning Goals:

- To use a scientific notebook as a primary record of procedures, data, observations, and example calculations
- To make scientific measurements
- To present your formal results through a laboratory report along with proper citations
- To use Excel to tabulate, calculate, analyze, and graph scientific data
- To apply balanced chemical equations and stoichiometric relationships to quantitative measurements

Additional Learning Goals for Experiment 4:

- To perform titrations and evaluate the reaction between strong and weak acids and bases
- To balance and differentiate molecular, complete ionic, and net ionic equations
- To relate measured chemical properties to the reactions of chemical species

Introduction

In this lab you will explore the nature of aqueous solutions by investigating the relationship between conductivity and strong and weak electrolytes. To do this, you will add increasing amounts of either acid or base to several electrolyte solutions. After each addition you will measure the conductivity of the solution. From the conductivity data, you will work to deduce the nature of the chemical reaction that occurred in the experiment. Plotting your data will allow you to follow the chemical reactions that are occurring and the ions that are present.

What is conductivity?

Conductivty is a measure of the concentration of ions in solution. By completing the circuit shown in Figure 1, we can measure the conductivity of the solution in the beaker. The conductivity is proportional to the current that flows between the electrodes. For current to flow, ions must be present in solution to carry the charge from one electrode to another. Increasing the number of ions in solution will increase the amount of charge that can be carried between electrodes and will increase the conductivity. The units microSiemens/cm (μS/cm) and milliSiemens/cm (mS/cm) are most commonly used to describe the conductivity of aqueous solutions.[1]

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[*] The Siemen, was formerly called mho (pronounced "mo") which was derived as a unit of conductivity by reversing the letters in "ohm," the unit of resistance.

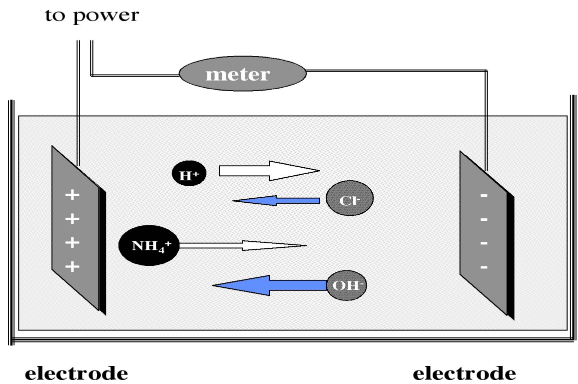

Figure 1. Schematic of a simple conductivity measurement system.

Another factor in conductivity measurements is that not all ions transport charge (conduct electricity) equally well. H^+ and OH^- are unique and move through solution very rapidly and are very good charge carriers. Ions such as NH_4^+ and Cl^- move through solution at a slower rate and therefore do not conduct electricity as well. Uncharged species in solution do not carry any charge. Table 1 is a table of molar conductivities for the ions in this exercise. The molar conductivity is the conductivity of a solution for the ion containing one mole of charge per liter. Note that the molar conductivity of H^+ ions is 5-7 times the conductivity of other small cations. The molar conductivity of OH^- is 3-5 times the conductivity of other small anions. To calculate the conductivity of a solution you simply multiply the concentration of each ion in solution by its molar conductivity and charge then add these values for all ions in solution.

Table 1. Molar Conductivity of Selected Ions

	Ion	Molar Conductivity (S L $mol^{-1}cm^{-1}$)
Cations:	H^+	0.34982
	Na^+	0.05011
	NH_4^+	0.0735
	K^+	0.0735
Anions:	OH^-	0.1986
	Cl^-	0.07635
	CH_3COO^- *	0.0409
	Br^-	0.0781

*CH_3COO^- is the acetate ion, and it is sometimes written as AcO^- or $C_2H_3O_2^-$.

How can we use conductivity to study solutions? Substances that completely dissociate into ions (strong electrolytes) produce solutions with high conductivity. Substances that partially dissociate (weak electrolytes) will produce solutions with low conductivity. Substances that dissolve but do not dissociate produce solutions that have the very low conductivity of pure water. How can we use conductivity to study chemical reactions? Since conductivity is a function of both the concentration and the composition of the solution being measured, we can use conductivity to follow chemical reactions. Consider an acid-base titration of a solution of KOH with HBr. The reaction is represented in the following:

Molecular equation:	$KOH\ (aq) + HBr\ (aq) \rightarrow KBr\ (aq) + H_2O$
Complete ionic equation :	$K^+\ (aq) + OH^-\ (aq) + H^+\ (aq) + Br^-\ (aq) \rightarrow K^+\ (aq) + Br^-\ (aq) + H_2O$
Net ionic equation:	$OH^-\ (aq) + H^+\ (aq) \rightarrow H_2O$

During the course of the titration OH^- is consumed (Figure 2). The equivalence point is reached when the number of moles of H^+ added is equal to the number of moles of OH^- in the starting solution. After the equivalence point, the moles of H^+ increase, because they are no longer consumed by reaction with OH^-. For the spectator ions, the number of moles of K^+ remains constant and the number of moles of Br^- increases as the titrant is added. How will these ion concentrations be reflected in the conductivity of the solution?

Initially the solution conductivity is high, because OH^- has a large molar conductivity. However, as OH- is neutralized the conductivity falls. After all the OH^- is neutralized the H^+ concentration increases and the conductivity again rises sharply. However, remember that the conductivity includes contributions from all ions in solution, not just H^+ and OH^-.

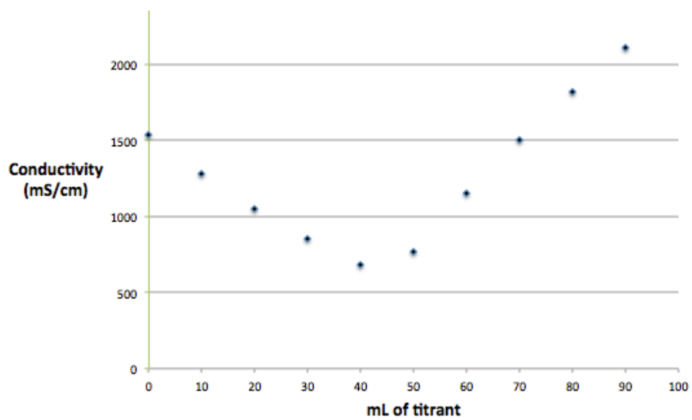

Figure 2. Titration of 150 mL of 0.00667 M KOH with 0.02 M HBr.

Procedure

For this experiment you will work in pairs and perform conductivity titrations on two different solutions assigned from the list in Table 2. You will then acquire data for the remaining three titrations from your classmates. You will use this conductivity data to predict the ionic composition of all five reactions before, during, and after each titration.

Table 2. Solutions to be titrated. ***Note that there are two different concentrations of NaOH present in the lab; be careful to use the correct one for your titrations.***

Starting Solution	Titrant
150 mL of 0.00667 M NaOH	0.02 M HCl
150 mL of 0.00667 M $NaC_2H_3O_2$	0.02 M HCl
150 mL of 0.00667 M $HC_2H_3O_2$	0.02 M NaOH
150 mL of 0.00667 M NH_3	0.02 M HCl
150 mL of 0.00667 M NH_4Cl	0.02 M NaOH

Measuring Conductivity:

You will be using a pH meter in conductivity mode (model: AR20) with a conductivity electrode to analyze your solutions. The meter will be setup for conductivity measurement, so that you will not need to press any buttons on the meter. The electrode will be standardized to 1.400 mS/cm.

When using the electrode, the level of aqueous solution should be in, but not over, the "circular white zone." The "white zone" allows the solution to be accurately measured (the solution must cover the "hole" in the electrode body so that sample can reach the electrode); while keeping the electronics dry. The electrode can be easily raised and lowered to keep the solution in this zone. Read the conductivity on the display when the meter reads "stable." Remember to record the unit of conductivity that the meter is displaying, as the meter will automatically change from uS/cm to mS/cm. When not in use, the electrode is stored in a small beaker of deionized water.

Rinse the electrode with distilled water and dry it gently with a Kim Wipe before each use. This will clean the electrode and minimize contamination between solutions. Use the provided plastic cup to collect the rinse water. This and all other solutions may be poured down the drain.

Performing Titrations:

Place 150 mL of the starting solution in a clean 250-mL beaker. The graduations on your beaker are of sufficient accuracy to measure the solution volume for this experiment. Add a stir bar to the beaker and stir the solution slowly using a magnetic stirrer. Do not allow the stir bar to hit the conductivity electrode. Measure the conductivity of the starting solution.

Titrate the starting solution by adding 10 mL increments of the titrant to the solution (use a graduated cylinder to measure the titrant volume. Adjust the probe height to ensure that the liquid surface is within the probe's white zone. Allow the solution to mix for a few seconds after each addition before measuring the conductivity. Record your results in your lab book. Add a total of 90 mL of titrant to the beaker. Rinse your beaker and electrode well with distilled water before the next titration.

Procedure for data analysis:

To analyze data for all 5 titrations, you will need to share your data with your classmates. First, copy your titration data into the class Excel spreadsheet on the fileserver. Next, copy a data set for each of the other three titration combinations (see Table 2) and add them to your Excel file. You will then make an appropriate plot of the conductivity during each titration (y-axis) versus mL of titrant (x-axis). Plot all five titrations on a single graph. Make sure that you pick a consistent unit for conductivity! Label the five series on the graph with the starting solution and titrant used in each reaction.

What should be in your laboratory notebook (*in addition to title, purpose, procedure, observations, etc.*)

1. Keep track of all measurements and *appropriate units*. Record all data and observations as they occur.
2. A table of your Excel spreadsheet data for all five titrations
3. The names of the other students who's data sets you used for your plot

Laboratory report: Use the **Report Form** for Experiment 4. You will also attach your completed Supplemental Report Sheet (2 sided) to your report.