

## Chem 60 Expt 9 \* Ionization and Electrical Conductivity



**Introduction:** Ionic substances “ionize” when dissolved in water, as shown by S. Arrhenius in the late 19<sup>th</sup> century. He showed that ions could carry only specific charges, measured in coulombs, i.e. that cations had a +1 or +2 or +3 relative amount charge (for ex. Al could carry three times as much positive charge as potassium) so Al was given the symbol  $\text{Al}^{3+}$  and K was shown as  $\text{K}^+$ . Similar work was done with anions

An ionization equation is written thus:  $\text{CuSO}_4(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

PS Arrhenius also did much work on chemical kinetics (rates of reactions) and proposed the “greenhouse effect” – as it relates to global warming. He was an early Nobel Prize winner.

You will observe the conductivity of several pure substances and of aqueous solutions of a number of substances. You will decide, based on the observed conductivity, if a substance is a strong electrolyte, a weak electrolyte or a nonelectrolyte. An electrolyte is a substance which conducts electric current in aqueous solution. A nonelectrolyte is a substance which does not ionize. Some substances are “weak” electrolytes and ionize to a limited extent. You will make a correlation between the concentration (molarity) of an electrolyte and the number of ions or the charges on the ions. You will use observed changes in conductivity as a way to monitor the progress of a chemical reaction and to predict the products of the reaction. The conductance probe you will use will measure the specific conductance (in microSiemens,  $\mu\text{S}$ ). This is the actual unit used. The term “conductivity” is a broader general term.

Conductance as it happens, will depend on a variety of factors, not only the charge of the ions, but the total number of cations and anions and also more complicated factors such as the sizes of the ions and the extent to which they hydrate in solution. We can limit our investigation to just examining the charges of the cations vs conductance and make a conclusion about whether this gives a good correlation. Also, the concentration (“Molarity”) of the solute should have an effect on the conductance.

It is important to rinse the probe prior to starting the measurements, and between measurements, in deionized water. Since you will need a number of beakers, and due to equipment limitations, you may need to work with a partner.

Part 1. Preparation of and measurements on 0.01 M solutions. ( M = moles per liter )

**Procedure.** Label five 125 mL Erlenmeyer flasks: 0.010 M NaCl, 0.010 M  $\text{MgCl}_2$ , 0.010 M  $\text{AlCl}_3$ , 0.010 M sucrose, 0.10 M acetic acid (diluted vinegar). Vinegar is about 5% or 0.80 M so you will dilute this by 1/80 to get 0.010 M acetic acid.

*Rinse your cylinder after each use, with some de-ionized water.*

- Obtain 10.0 mL of 0.10 M NaCl. Add deionized water to the 50 mL mark in a grad cylinder. Pour the mixture into the Erlenmeyer labeled 0.010 M NaCl. Fill the cylinder again with de-I water and add it to the flask.
- Similarly, prepare 0.010 M solutions of  $\text{MgCl}_2$  and  $\text{AlCl}_3$  in separate flasks
- Weigh 0.90 g sucrose into an Erlenmeyer flask. Add 100 mL of deionized water
- Measure 1.2 mL (26 drops) vinegar into a grad cylinder. Add deionized water to the 50 mL mark and empty the solution into an Erlenmeyer flask. Add 50.0 more mL de-I water to the flask.

Prepare a 250 mL beaker with about 200 mL of deionized water. Plug in the probe to the meter and the meter to an AC outlet. Push only the ON button at the upper left of the meter. Set the switch on the conductance probe black box to 0-20000. Actually the probe will accept readings up to about 30000 but such a reading should be viewed as unreliable. Measure the conductance of the de-I water. Some meters have a default minimum reading of about 100 microS. You can call this the zero conductivity baseline. If the conductance is more than 100 microSiemens, rinse the beaker with more de-i water and try again. The meter may fluctuate slightly at the low level of conductivity. Also read the conductance of tap water.

Read the conductances of all the 0.010 M solutions you have prepared. (You can insert the probe into the flasks but be sure to rinse the probe between each measurement.) Readings may fluctuate. 3 sig figs ok.

### Part 2. Effect of dilution of solutions on conductance

- a) Prepare diluted solutions of NaCl as follows.
  - i) Pour 25 mL of the 0.010 M NaCl into the 50 mL grad cylinder and dilute with de-i water to 50 mL.  
Pour this into a flask or beaker and label 0.0050 M NaCl
  - ii) Pour 10 mL of the 0.010 M NaCl into the grad cylinder and add de-i- water to the 50 mL mark.  
Pour this into a flask or beaker and label it 0.0020 M NaCl
  - iii) Pour 5 mL of the 0.010 M NaCl into the grad cylinder and add de-i water to the 50 mL mark.  
Pour this into a flask or beaker and label it 0.0010 M NaCl

Measure the specific conductance of each of the NaCl solutions, starting with the most diluted solution (iii) and finally the original 0.0100 M solution. (This will obviate the need for rinsing the probe between each run, as the effect of a drop or two of the dilute solution will have little effect on the conductivity of a more concentrated solution).

- b) Prepare three solutions of  $\text{MgCl}_2$  in exactly the same manner as in part a) with the NaCl solutions and measure their specific conductances. (don't forget to record the conductivity of the 0.0100 M solution)
- c) Prepare three diluted solutions of the  $\text{AlCl}_3$  just as in parts a and b You will graph the NaCl and  $\text{MgCl}_2$  and  $\text{AlCl}_3$  conductances vs M on the same graph – be sure to label which set of points belongs to which chemical ! Use different colors of ink. Do record the conductivity of the 0.0100 M solution.

### Part 3. The effect on conductivity by precipitating out some of the ions.

- a) Weigh out 0.60 g MgO into a watch glass
- b) Weigh out 1.26 g oxalic acid dihydrate (molar mass 126)  $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$  into a 250 mL beaker  
Add 100 mL  $\text{H}_2\text{O}$  and measure the conductance.
- c) Using the tip of your spatula, add small portions of the MgO to the oxalic acid solution, each time stirring with a stirring rod. It takes a minute for each portion of MgO to dissolve and react with the acid. Measure the conductances after each portion has been added (about 5 portions in all), after waiting about a minute each time. After adding the last portion, stir the mixture for about 2 minutes and measure the conductance. If the value doesn't become constant, wait an additional 5 minutes and call it done – take the final reading. (you will have added all the MgO in about 5 or 6 portions)

Optional. If time permits, your instructor may assign some additional samples for you to measure with the conductance probe, such as various kinds of bottled drinking water. Or a demonstration using a light bulb can be done to demonstrate general conductivity. Write these observations in your notebook and comment on them in your conclusion. For example – compare your drinking water with other students in the class. Set up the data in your notebook. Your instructor will collect the copy sheet. (some extra credit)

### Analysis

Graph the data obtained in part 2. On the same piece of graph paper, make three graphs showing the conductance ( $\mu\text{S}$ ) on the y axis, and the concentrations (ranging from 0.0010 to 0.0100 M) on the x axis. Do this by hand. Plot the points and draw the best fit straight line you can through each set of points. Find the slope for each line. Set up the slope calculations neatly on the graph paper.

In your questions and conclusion, what can you say about a) the conductance vs the charges of the ions and b) the conductance vs the number of ions in solution.

**Report. Expt 9**

Name: \_\_\_\_\_

**Ionization & Conductivity**

Section \_\_\_\_\_ Lab partner \_\_\_\_\_ Date of experiment \_\_\_\_\_

Due date \_\_\_\_\_

## Data &amp; Observations

Part 1. Initial reading on deionized water: \_\_\_\_\_  $\mu\text{S}$  . regular tap water \_\_\_\_\_  $\mu\text{S}$ 

0.0100 M solution	Conductance ( $\mu\text{S}$ )	strong, weak, or non-electrolyte
a) NaCl		
b) $\text{MgCl}_2$		
c) $\text{AlCl}_3$		
d) sucrose		
e) $\text{HC}_2\text{H}_3\text{O}_2$		

Part 2 a) conductance

Part 2 b).

Part 2 c)

NaCl(aq)	( $\mu\text{S}$ )
i. 0.0050 M	
ii. 0.0020 M	
iii. 0.0010 M	

$\text{MgCl}_2$ (aq)	( $\mu\text{S}$ )
i. 0.0050 M	
ii. 0.0020 M	
iii. 0.0010 M	

$\text{AlCl}_3$ (aq)	( $\mu\text{S}$ )
i. 0.0050 M	
ii. 0.0020 M	
iii. 0.0010 M	

Part 3

solution	( $\mu\text{S}$ )
0.10 M $\text{H}_2\text{C}_2\text{O}_4$	
Addition of MgO ( 1 <sup>st</sup> portion )	
Additional portions	

Describe the appearance of the mixture as the MgO is added.

Analysis, questions, conclusions.

1. Compare the relative conductances of the three ionic substances. calculating ratios

slope of NaCl graph _____	Ratios	cation charge
slope of MgCl <sub>2</sub> graph _____	slope of MgCl <sub>2</sub> / slope of NaCl _____	$\frac{+1}{+2}$ .
Slope of AlCl <sub>3</sub> graph _____	slope of AlCl <sub>3</sub> / slope of NaCl _____	$\frac{+1}{+3}$

Is there a general correlation between the charges on the cations and the slope ratios ? Explain

2. Is sucrose an ionic substance ? \_\_\_\_\_ Find the formula somewhere in your text or another reference. Explain your observation of the conductance reading.

3. A “strong” acid such as HCl is ionized like NaCl . We don’t abuse the sensitive probes with this substance, but we can test acetic acid. What kind of acid is acetic acid – based on your conductance measurement ? Is it a strong or weak acid? Why ?

4. Tap water contains Ca<sup>2+</sup> and Mg<sup>2+</sup> and smaller amounts of other cations. Assuming that the Ca<sup>2+</sup> and Mg<sup>2+</sup> behave much the same, you can read the conductance of tap water and Read the approximate total molarity of the Mg<sup>2+</sup> and Ca<sup>2+</sup> on the graph you have constructed

tap water : \_\_\_\_\_ μS      \_\_\_\_\_ M

5. A commonly used unit of concentration for tap water is called “ppm” or parts per million of CaCO<sub>3</sub>. If we assume all the dissolved solids are calculated as ppm of CaCO<sub>3</sub> , then a 0.01M solution is equivalent to 1000 ppm. ( also read as mg CaCO<sub>3</sub> / L ) How many ppm of dissolved solids (assuming CaCO<sub>3</sub> ) were present in the LACC tap water.

M CaCO<sub>3</sub> \_\_\_\_\_ off the graph = \_\_\_\_\_ ppm .

6. What did you see happening in the beaker as you added the MgO to the H<sub>2</sub>C<sub>4</sub>O<sub>4</sub> in part 3. Write an equation for the reaction that occurs. Designate the states (s) = solid (aq) = solution

Observation:



7. What happened to most of the ions in the MgO – oxalic acid mixture ? Why can you draw this conclusion ?

8. A small excess of MgO was added to the oxalic acid. Excess MgO is virtually insoluble. Still there is a residual conductivity, presumably from some small amount of magnesium oxalate which is more soluble than MgO. Compare the conductivity of the mixture before and after the addition of MgO. What percent of the magnesium oxalate precipitated out? (Assume all the  $H^+$  reacts with oxide to form water, and does not contribute to conductivity)).

General conclusions: What did you learn. Summarize your results briefly, including any extra discovery work you may have done on bottled water samples. Which brands are good for you? (i.e. contain sufficient  $Ca^{2+}$ ). If you did a lab where the  $Ca^{2+}$  was determined by titration – compare the results of the ppm of  $Ca^{2+}$  and  $Mg^{2+}$  with the results from the titration experiment using EDTA. If your handwriting is lousy – type! Use the back of this paper or a separate sheet if needed with your neatly written or typed conclusions. **WRITE NEATLY. SLOPPY HANDWRITING IS NOT ACCEPTABLE.**