Experiment 5

Conductivity and Chemical Reactions

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PURPOSE

A conductivity probe will be used to identify different types of electrolytes. The relationship between conductivity and ionic concentration will also be explored.

ELECTROLYTES

An **electrolyte** is a species that exists as ions (charged particles) when dissolved in solution. Since electricity is simply the flow of charged particles, an aqueous solution containing ions is able to conduct electricity since the ions are free to move.

STRONG ELECTROLYTES

A species that completely dissociates into ions in solution is called a **strong electrolyte**. When *any ionic compound* dissolves in water, it will completely dissociate into ions. For example:

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq) \qquad (1)$$

$$K_2SO_4(s) \xrightarrow{H_2O} 2\,K^+(aq) + SO_4^{2-}(aq) \qquad (2)$$

Notice that when one mole of NaCl dissolves it produces two moles of ions. But, when one mole of K_2SO_4 dissolves it produces three moles of ions. Thus, if equal moles of each compound are dissolved, fewer moles of ions are produced by the NaCl solution and so we would expect its conductivity to be smaller.

Note that when the ionic compound dissolves, the ions separate from each ion, but the polyatomic ions do not decompose. For example, in reaction (2) we see that the sulfate ion did not separate into sulfur and oxide ions. So, it is important that you recognize the polyatomic ions (and do not decompose them) when writing the dissociation reactions.

Almost all common cations are monatomic (Na^+, Ca^{2+}, $K^+, \ldots$), except for the ammonium ion (NH_4^+). There are many common polyatomic anions encountered in the laboratory including carbonates (CO_3^{2-}), acetates ($C_2H_3O_2^-$), nitrates (NO_3^-), and nitrites (NO_2^-), sulfates (SO_4^{2-}) and sulfites (SO_3^{2-}), hydroxides (OH^-), and phosphates (PO_4^{3-}).

When dissolved in water, some acids (called *strong* acids) completely dissociate into ions:

$$HBr(g) \xrightarrow{H_2O} H^+(aq) + Br^-(aq) \qquad (3)$$

The chemical formulas for acids are commonly written with an H in front ($HCl,\ HBr,\ HNO_3, \ldots$).

Similarly, strong bases will also completely dissociate in water to produce hydroxide ions.

$$LiOH(s) \xrightarrow{H_2O} Li^+(aq) + OH^-(aq) \qquad (4)$$

Strong bases are typically metal hydroxides.

WEAK ELECTROLYTES

Other acids and bases only partially ionize in solution. They are called *weak* acids and bases and are examples of **weak electrolytes**:

$$HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq) \qquad (5)$$

$$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3{}^+(aq) + OH^-(aq) \qquad (6)$$

Notice how these reactions are written with a double arrow (instead of a single arrow). This indicates that at the end of the reaction (at equilibrium) species on both sides of the equation still exist in solution.

Weak bases are typically nitrogen-containing compounds where the nitrogen can pull an H^+ off of water, leaving an OH^- behind.

A weak-electrolyte solution has a much smaller conductivity than a strong-electrolyte solution of equal concentration. *Under our experimental conditions, the conductivity will change by a factor of ten (approximately). 0.050 M strong electrolytes will have conductivities in the 1000's (of μS/cm), whereas 0.050 M weak electrolytes will be in the 100's.*

NONELECTROLYTES

Not all species produce ions when dissolved in solution. Most covalently-bonded (molecular) compounds are **nonelectrolytes**, meaning they do not form any ions in solution. For example:

$$CO(NH_2)_2(s) \xrightarrow{H_2O} CO(NH_2)_2(aq) \qquad (7)$$

The (aq) phase indicates that the $CO(NH_2)_2$ molecules have dissolved and are now surrounded by water molecules (instead of being surrounded by other $CO(NH_2)_2$ molecules as in the solid phase). However, since it is a covalently-bonded compound, the $CO(NH_2)_2$ molecules remain intact and do not break apart into ions in solution.

However, even in a solution containing only nonelectrolytes, there will always be some conductivity since a small fraction of the water molecules will dissociate to form ions (although sometimes it is too small to easily measure)

$$2\,H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \qquad (8)$$

Furthermore, even distilled water can readily absorb impurities from the pipes it travels through or from the surrounding air. For example, water will absorb carbon dioxide from air to produce a weak acid, carbonic acid.

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq) \qquad (9)$$

Carbonic acid produces ions in solution:

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \qquad (10)$$

Unlike distilled water, tap water contains many ionic impurities that significantly increase its conductivity. The exact conductivity value depends upon the composition of the tap water.

THE CONDUCTIVITY OF MIXTURES

If equal volumes of two solutions are mixed, we expect the conductivity of the mixture to be approximately the average of the conductivities of the two solutions (assuming no chemical reaction occurs). Specifically, if the two solutions initially had equal conductivities then the mixture should have about the same conductivity.

However, if a chemical reaction occurs when the two solutions are mixed, this changes the total concentration of ions and the resulting conductivity of the mixture. Most commonly, chemical reactions reduce the number of ions in solution. One common way this occurs is through solid formation (precipitation reactions):

$$KCl(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgCl(s) \qquad (11)$$

Another common way is through the formation of a nonelectrolyte (*e.g.*, H_2O), such as in a strong acid-strong base neutralization reaction:

$$HBr(aq) + LiOH(aq) \rightarrow LiBr(aq) + H_2O(l) \qquad (12)$$

In either of the above reactions, ions are being removed from solution resulting in a decrease in the mixture's conductivity.

THE MEASUREMENT OF CONDUCTIVITY (This section is rated P – and should only be read by those with an interest in Physics)

We will use a conductivity probe to measure the concentration of ions in solution. The probe consists of two electrodes separated by an empty space where solution can flow. (When you look at a conductivity probe you will see a single tube with an elongated hole in the end. The two electrodes are on either side of the hole and the entire elongated hole must be immersed in solution). A voltage (potential difference) is applied across the two electrodes causing positive ions to move towards the more negative electrode and negative ions to move towards the more positive electrode.

Using the applied voltage *V* and measured current *I*, the resistance R can be calculated (using Ohm's Law: $V = IR$). The conductance is equal to the reciprocal of the resistance and is measured in **siemens** (formerly called *mhos*). The conductivity equals the conductance multiplied by the separation between the two electrodes (in cm) divided by the surface area of the electrodes (in cm^2). We are using 1 cm^2 electrodes, 1 cm apart, so conductance and conductivity have the same numerical value for our probes. The units of conductivity are siemens/cm, but since we are using dilute solutions, we will be taking our measurements in microsiemens per cm (μS/cm).

IN THIS EXPERIMENT

The conductivity of a variety of different solutions will be measured. The solutes will be identified as strong, weak, or non-electrolytes. The relationship between concentration and conductivity will be investigated. Finally, several pairs of solutions will be mixed and the conductivity of the resulting mixture will be used to determine whether a chemical reaction occurred in each case.

PRE-LABORATORY PREPARATION

1. **Read the procedure and data analysis sections of this experiment.**
2. **Complete the PRELAB assignment in Canvas.**

EXPERIMENTAL SECTION

REAGENTS PROVIDED

Solid sodium chloride, NaCl(s)
0.050 M aqueous solutions of the following:

- **calcium chloride, $CaCl_2(aq)$**
- **hydrochloric acid, HCl(aq)**
- **nitric acid, $HNO_3(aq)$**
- **acetic acid, $HC_2H_3O_2(aq)$**
- **sodium hydroxide, NaOH(aq)**
- **ammonia, $NH_3(aq)$**
- **methanol, $CH_3OH(aq)$**
- **sucrose, $C_{12}H_{22}O_{11}(aq)$**
- **potassium nitrate, $KNO_3(aq)$**
- **sodium carbonate, $Na_2CO_3(aq)$**

Dispensing stock reagents. Take only a <u>reasonable</u> amount as specified in the directions. Some reagents are quite expensive, and all cause great inconvenience to others if they are prematurely exhausted.

A Golden Rule of chemistry laboratory work is that **solutions are never returned to stock bottles**. We tolerate zero risk of some klutz with dirty glass-ware contaminating stock solutions, and thereby wrecking numerous subsequent analyses. So, if you take too much, it is of necessity wasted.

HAZARDOUS CHEMICALS

Handle any acidic or basic solutions with care as they are often toxic and/or corrosive. Methanol is flammable and toxic.

WASTE DISPOSAL

All of the chemicals used in this experiment may be safely disposed of by washing down the sink.

PROCEDURE

Unless told otherwise, you will work with a partner.

CONDUCTIVITY PROBE SETUP

1. Clean a 400-mL beaker and fill it half full with distilled water. Place a conductivity probe in the water to soak while completing the preparation steps.

This will remove any residue left on the probe from previous labs. You will use this solution to rinse your probe between measurements.

2. Turn on the LabQuest2 by pressing the red power button on the top.

It takes a minute to warm up. It then should start in Time Based Data Collection mode.

3. Once warmup is complete, connect the Conductivity Probe amplifier box to CH 1 on the left end of the LabQuest2. Set the range switch on the amplifier box to the 0-20000 μS setting (the up position).

Most of the conductivity probes will immediately be recognized and a box will appear containing the conductivity reading.

<u>If the meter box does not appear</u>, you have an older probe with an analog-to-digital adapter and it needs to be set up manually, as follows:

a. Tap ***Sensors*** from the top menu line.
b. Tap ***Sensor Setup…*** .
c. Tap in the large box next to CH 1 that reads "No Sensor".
d. Scroll down and tap on ***Conductivity***.
e. Tap on ***Conductivity 20000MICS***.
f. Tap ***OK***.
g. Tap ***OK*** to return to the meter screen, where you should now see the conductivity reading.

5. Calibrate the conductivity probe:

a. Tap ***Sensors*** (from the top menu line).
b. Tap ***Calibrate*** and then ***CH1: Conductivity***
c. Tap ***Calibrate Now***.
d. Dry the conductivity probe thoroughly (even inside of the elongated hole at the bottom).
e. Hold the dry conductivity probe in the air and observe the voltage reading displayed in the middle of the Sensor Settings window. It should be 0.00 V (± 0.01). If it is fluctuating, wait 10-15 seconds until it stabilizes, then <u>enter 0 in the box for **Value 1**</u> and tap ***Keep***.
f. Immerse the conductivity probe in the 10,000 μS NaCl(aq) standard solution provided. (Leave this solution in the bottle provided and do NOT pour it into a beaker.) Swirl the probe to remove any trapped air bubbles. Wait until the voltage stabilizes (in about a minute) between 0.9 and 1.5 V, <u>then enter 10000 in the box for **Known value 2**</u> and tap ***Keep*.**
g. Tap OK to complete the calibration.
h. The conductivity value should be fluctuating around 10,000 μS/cm. If not, something is wrong. Try re-calibrating or get help.

CONCENTRATION STUDY

6. Add 150 mL of distilled water to a clean 250 mL beaker.

Ordinarily, we do not really trust the calibration lines on beakers, but in this experiment we do not need to be very precise (or accurate) so this will work fine.

7. On separate pieces of glassine paper, weigh out three samples of NaCl(s). Each sample should be between 0.14 and 0.15 g. Record these values on your data sheet.

We will use these samples to test the effect of concentration on conductivity. Each sample contains about 0.0025 moles of NaCl. So when added to 150 mL (0.150 L) of water, they each increase the concentration by 0.0025 mol / 0.150 L = 0.017 mol/L. Further, when all three samples are added, the concentration will be about 0.050 M (the same concentration as the other solutions we will use later).

Even though we will not be performing any calculations using these masses, it is still a good idea to record the values. If we get surprising results, we will be able to determine whether any error was due to the mass of NaCl used.

8. Rinse the conductivity probe (by swirling it in your 400-mL rinse beaker), dry it with a paper towel, and then place it in the 250-mL beaker (containing 150 mL of distilled water) and measure the conductivity of the distilled water. Record the result on your data sheet.

The readings will normally fluctuate quite a bit. *Typically, wait about 15 seconds and then record the average reading to* <u>*one or two significant figures*</u> *– whatever you think you can accurately obtain*.

For distilled water, you should see a value close to zero. It may even be slightly negative. Theoretically, the conductivity should have a very tiny positive value. But, the random fluctuations in the measurements (and calibration) are larger than the value itself.

9. Lift the probe out of the beaker and add one of the three solid NaCl samples, weighed out earlier, to the distilled water. Stir thoroughly (with a glass stirring rod) until all of the salt has dissolved.

Some good habits: (1) Try to hold the probe above the solution so that if it drips, the drops stay in the beaker. (2) Leave the stirring rod in the beaker (since removing it will remove solution and influence your later results).

10. Put the probe back into the solution and swirl briefly to remove any air bubbles. Record the conductivity on your data sheet.

This is your 0.017 M NaCl(aq) solution. You may need to wait 10-15 seconds for the reading to stabilize. Again, do not expect the reading to completely stop fluctuating (or, sometimes, even systematically drifting). However, the magnitude of the change is small compared to the value itself and so it is not worth the time it would take to wait for the value to totally stop fluctuating. <u>*Record the reading to the nearest 100 (or 10 if it is sufficiently stable).*</u>

It is not necessary to clean the conductivity probe between measurements here, because you are putting the probe back into the same solution from which it was removed.

11. Add a second solid NaCl sample, weighed out earlier, to the same solution. Stir thoroughly (with a glass stirring rod) until all of the salt has dissolved. Then measure and record the conductivity.

This is your 0.033 M NaCl(aq) solution.

12. Add your final solid NaCl sample to the same solution. Stir thoroughly until all of the salt has dissolved. Then measure and record the conductivity.

This is your 0.050 M NaCl(aq) solution. *Save this solution since you will use it again later*.

IONIC – MOLECULAR COMPOUNDS STUDY

13. Clean and dry six small (50 or 100 mL) beakers. Add the following six solutions to the six beakers (one solution per beaker) until the beaker is about half full:

- **(a) tap water**
- **(b) 0.050 M KNO_3(aq)**
- **(c) 0.050 M CH_3OH(aq)**
- **(d) 0.050 M $C_{12}H_{22}O_{11}$(aq)**
- **(e) 0.050 M $CaCl_2$(aq)**
- **(f) 0.050 M Na_2CO_3(aq)**

Add enough liquid so that the entire elongated hole at the bottom of the conductivity probe will be immersed. (so, 50-mL beakers may need to be a little more than ½ full.) Do not use larger beakers because it wastes too much solution.

Be sure to label the beakers (with tape) or place the beakers on paper (or paper towels) with labels so that you do not confuse what is in the beakers. *Solutions (b), (e), and (f) will be used again later so do not discard them after recording their conductivity*.

14. Rinse (in your 400-mL beaker) and dry the conductivity probe and then measure the conductivity of the six liquids listed above. Record your results on your data sheet.

Always swirl the probe in the solution to remove any air bubbles. As above, only wait 10-15 seconds for the reading to somewhat stabilize before taking your reading. *Try to read the conductivity to the nearest 10, but for conductivities greater than 1000 you may not be able to do any better than the nearest 100.*

<u>*Be sure to thoroughly rinse the probe between measurements.*</u> After removing it from each solution, immerse it in your 400-mL beaker of distilled water. Then dry the probe using a paper towel (including inside of the elongated hole).

15. Pick up the 250-mL beaker containing 0.050 M NaCl that you prepared earlier. Carefully pour out solution (into a clean beaker or flask – in case you pour out too much) until only about 50 mL are left in the 250-mL beaker. Then add an equal volume (50 mL) of solution (b) 0.050 M KNO_3(aq) to the 250-mL beaker. Read and record the conductivity of this solution (Mixture I).

This solution and the excess NaCl solution can be discarded down the drain once the conductivity is measured.

16. Clean and dry your 250-mL beaker and then mix equal amounts of solutions (e) 0.050 M $CaCl_2$(aq) and (f) 0.050 M Na_2CO_3(aq) together in the beaker. (You should have at least 40 mL of each solution.) Stir thoroughly (with a glass stirring rod) and then measure and record the conductivity of this solution (Mixture II).

You should see a white precipitate form when the two solutions are mixed.

17. Discard all of the solutions down the drain (even the last one – the solid will dissolve later when the acid goes down the drain). Clean and dry your beakers.

18. Clean and <u>dry</u> the conductivity probe. Next obtain the beaker of solid NaCl from the front bench. Immerse the probe in the solid salt and measure and record its conductivity.

Return the beaker of solid salt to the front bench when you are finished (so other groups can use it).

ACID – BASE CONDUCTIVITY STUDY

19. Now add the following liquids to the clean, dry, small beakers (one solution per beaker):

- **(a) 0.050 M HCl(aq)**
- **(b) 0.050 M HNO_3(aq)**
- **(c) 0.050 M $HC_2H_3O_2$(aq) or CH_3COOH(aq)**
- **(d) 0.050 M NH_3(aq) or NH_4OH(aq)**
- **(e) 0.050 M NaOH(aq)**

Again, fill the beakers about halfway, so there is enough liquid to completely immerse the entire elongated hole at the bottom of the conductivity probe. Be sure to label the beakers. *Solutions (a) and (e) will be used later so do not discard them after measuring their conductivities.*

When ammonia is dissolved in water, some of it reacts with H_2O to form NH_4OH(aq), so this is an alternate way of writing NH_3(aq).

The formula for acetic acid, $HC_2H_3O_2$, is frequently written as CH_3COOH, which better shows how the atoms are bonded together in the molecule.

20. Measure and record the conductivity of each of these five solutions.

Be sure to thoroughly rinse and dry the probe in between measurements.

21. Clean and dry your 250-mL beaker and then mix equal amounts of solutions (a) 0.050 M HCl(aq) and (e) 0.050 M NaOH(aq) together in the beaker. (You should have at least 40 mL of each solution.) Stir thoroughly (with a glass stirring rod) and then measure and record the conductivity of this solution (Mixture III).

22. Discard all of the solutions down the drain. Clean and dry your glassware.

Return all of the beakers to the drawers from whence they came.

23. Shut down the LabQuest2 by first tapping *File – Quit*. Next, tap on the *System* folder and then *Shut Down* and, finally, *OK*.

RETURN EVERYTHING TO WHERE IT WAS AT THE START OF LAB. HAVE AN INSTRUCTOR CHECK YOUR STATION BEFORE LEAVING.

Once your station is clean, wash your hands.

Name ______________________ Station Used ____________ Instructor/Day/Time ______________________

Partner ______________________ Station Checked & Approved ______________________

DATA SHEET

Record all data values to the appropriate number of significant figures and with units (if not given below).

Concentration Study

Mass NaCl ____________ ____________ ____________

Sample 1 Sample 2 Sample 3

Conductivity (μS/cm)			
Distilled Water	0.017 M NaCl	0.033 M NaCl	0.050 M NaCl

Ionic – Molecular Compounds Study

Conductivity (μS/cm)			
Tap Water	0.050 M KNO_3	0.050 M CH_3OH	0.050 M $C_{12}H_{22}O_{11}$

Conductivity (μS/cm)				
0.050 M $CaCl_2$	0.050 M Na_2CO_3	NaCl & KNO_3 Mixture I	$CaCl_2$ & Na_2CO_3 Mixture II	Solid NaCl

Acid/Base Conductivity Study

Conductivity (μS/cm)					
0.050 M HCl	0.050 M HNO_3	0.050 M $HC_2H_3O_2$	0.050 M NH_3 (or NH_4OH)	0.050 M NaOH	HCl & NaOH Mixture III

DATA ANALYSIS

1. Based upon your measured conductivities, classify each of the 11 <u>solutes</u> as strong, weak, or non-electrolytes:

NaCl(aq), KNO_3, CH_3OH, $C_{12}H_{22}O_{11}$, $CaCl_2$, Na_2CO_3, HCl, HNO_3, $HC_2H_3O_2$, NH_3, NaOH

Strong:

Weak:

Non:

2. Complete and balance each reaction indicating what happens when the solute dissolves in H_2O. For examples, see reactions (1) – (7) in the background section on Electrolytes. For any weak electrolytes change the arrow to the proper type (→ changes to ⇄). Be sure to give the proper phase for each product (aq).

$$NaCl(s) \xrightarrow{H_2O}$$

$$KNO_3(s) \xrightarrow{H_2O}$$

$$CH_3OH(l) \xrightarrow{H_2O}$$

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O}$$

$$CaCl_2(s) \xrightarrow{H_2O}$$

$$Na_2CO_3(s) \xrightarrow{H_2O}$$

$$HCl(g) \xrightarrow{H_2O}$$

$$HNO_3(l) \xrightarrow{H_2O}$$

$$HC_2H_3O_2(aq) \xrightarrow{H_2O}$$

$$NH_3(aq) + H_2O(l) \longrightarrow$$

$$NaOH(s) \xrightarrow{H_2O}$$

3. Based upon your conductivity vs. concentration data for the three NaCl(aq) solutions, it appears that (choose the best choice):

A. Conductivity is independent of concentration.
B. Conductivity is proportional to concentration.
C. Conductivity is inversely proportional to concentration.

Based upon your experimental values and your conclusion circled above, roughly predict the conductivity of 0.10 M NaCl(aq). Briefly explain how you arrived at your predicted value.

4. Consider the four 0.050 M salt solutions that were used: NaCl, KNO_3, $CaCl_2$, and Na_2CO_3

(a) Look at the balanced reactions you completed in question 2. How many moles of ions are produced when one mole of each salt dissociates:

NaCl __________ **KNO_3** ___________ **$CaCl_2$** ___________ **Na_2CO_3** ___________

(b) Copy the conductivities (in μS/cm) of the four separate salt solutions from your data sheet.

NaCl __________ **KNO_3** ___________ **$CaCl_2$** ___________ **Na_2CO_3** ___________

(c) Is there a relationship between the moles of ions formed when the salt dissociates and the measured conductivities? If so, what is it?

5. Look at the measured conductivities for distilled water and solid NaCl. Now look at the conductivities for the solutions of dissolved NaCl in distilled water. Briefly explain why the solution's conductivity is so much different than that of the pure solute and pure solvent. (In other words, what can happen in the solution that cannot happen in the pure solute or solvent?)

6. Based upon your measured conductivities, classify each of the 3 acids (HCl, HNO_3, $HC_2H_3O_2$) **and 2 bases** (NH_3, NaOH) **as strong or weak:**

Strong Acids:

Weak Acids:

Strong Bases:

Weak Bases:

7. Use the conductivity readings for the strong and weak acids to determine the approximate percentage of weak acid molecules that ionize (form hydronium ions) in water.

Here we will assume that the strong acid molecules quantitatively (100%) form hydronium ions and that the different anions contribute equivalently to the conductivity. Based upon the relationship between ion concentration and conductivity, we can estimate the percentage of weak acid molecules that ionize as

$$\frac{\text{Weak acid conductivity}}{\text{Strong acid conductivity}} \times 100\% = \%\ \text{ionization}$$

If you found more than one weak or strong acid, use the value for the one that gave the highest conductivity in your calculation. *Round your answer to 2 significant figures*.

Show your calculation below:

8. Based upon the conductivities of the three mixtures prepared, relative to the conductivities of the reactants, determine whether or not a chemical reaction occurred, as follows:

(a) Copy the relevant conductivity values from your data sheet and complete the table below:

Conductivities (μS/cm)				
	First Reactant	Second Reactant	Average of the Reactants	Measured
Mixture I	NaCl	KNO_3		
Mixture II	$CaCl_2$	Na_2CO_3		
Mixture III	HCl	NaOH		

(b) By comparing the average and measured conductivities, decide whether or not a reaction occurred in each mixture (see the background section of the experiment on the conductivity of mixtures, if necessary). If a reaction occurred, complete and balance the reaction equation below [including the appropriate phase (aq), (l), or (s) for each product]. Write N.R. if no reaction occurred. [For examples, see reactions (11)-(12) of the background section.]

Hint: if needed, the solubility guidelines (given in section 4.5 of your lecture textbook) can help you determine the identity of the precipitate (solid) that formed in one of the reactions.

I. $NaCl(aq) + KNO_3(aq) \longrightarrow$

II. $CaCl_2(aq) + Na_2CO_3(aq) \longrightarrow$

III. $HCl(aq) + NaOH(aq) \longrightarrow$