INTRODUCTION

Electrolytes are substances that consist of charged particles called ions. When electrolytes are dissolved in water (or other polar solvents) they ionize into positive (cation) and negative (anion) ions. In this experiment, you will explore what types of compounds can become electrolytes, what determines electrolyte strength, and how electrolytes are involved in the conduction of electricity.

Discussion

Before we can discuss the properties that make a compound an electrolyte, we must first understand something about the properties of electricity.

Electricity results from the movement of charged particles through a conductor. The charged particles can be either electrons or ions (positive or negative). In some cases, both types of particles can be involved. When the movement of electricity is through a metal, the electrons move from one metal atom to another which serves as the means for carrying the charge in the electrical circuit. If a liquid is included as part of the electric circuit, something must carry the charge through this solution otherwise no electrical current will flow. When a non-electrolyte is added, no current flows. Therefore, if a light bulb is also included in an electrical circuit containing a polar liquid, it is possible to tell whether the compound being added to the liquid is an electrolyte or a non-electrolyte by whether or not the light bulb lights up. Note the set-up in the fume hood. The bulb assembly clamps to a ring stand.

Determining the type of bonds present (Ionic, Polar Covalent or Nonpolar Covalent)

The percent ionic character of a bond is dependent on the differences in electronegativities (see sec. 9.6, Tro, 1st ed. and Figure 9.10) of the atoms present, and the arrangement of these atoms in the molecule. When electronegativity differences are very large (>2.0) a bond is considered ionic; the
electrons are completely transferred from one atom to another. When the electronegativity difference is intermediate \((2.0 > x > 0.5)\) the bond is considered *polar covalent*, and regions of partial positive and partial negative charges are formed. When the electronegativity difference is small \((< 0.5)\) then the molecule is considered *nonpolar covalent*. Although we have not covered these concepts in class yet, these trends can be estimated by the position of the two atoms on the Periodic Table. When a compound is formed from elements that are on opposite sides of the Periodic Table, such as from the s-block (metal) and the p-block (nonmetal), that compound is typically ionic. When a compound is formed from elements all within the p-block (all nonmetals), that compound is typically covalent. Hydrogen is often an exception – it is often considered a nonmetal in bonding. Use the flowchart below to help you decide how to classify individual molecules as ionic, polar covalent, or nonpolar covalent.

### Determining/Predicting Electrolyte Behavior

There are three types of electrolytes: strong, weak and non-electrolytes. A *polar solvent* is necessary for an electrolyte to function:

1. Ionic compounds that are *soluble* are *strong electrolytes*. (Review solubility rules). Ionic compounds that are *insoluble* are *non-electrolytes*.
2. Polar covalent compounds can be strong, weak or non-electrolytes. Most acids have weak polar covalent bonds to the hydrogen, allowing water to pull the \(H^+\) off of the molecule – thus acids will be an electrolyte – either strong or weak. *(strong acids are strong electrolytes, weak acids are weak electrolytes)*
3. Non-polar covalent compounds form *non-electrolytes*.

For example, the ionic compound NaCl, when dissolved in a polar solvent, dissociates (splits apart) completely to form sodium ions and chloride ions. These ions form because the electrostatic forces present in the polar solvent help to stabilize the ions.

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\[
\text{NaCl} \xrightarrow{\text{polar solvent}} \text{Na}^+ + \text{Cl}^-
\]

**Ionic compounds** are strong electrolytes in polar solvents if they are soluble; they form large numbers of ions that can carry a large electrical current.
**Polar covalent compounds** may or may not form ions when dissolved in a polar solvent, depending on the compound. **Strong acids (polar covalent),** (e.g. HNO₃, H₂SO₄, HClO₄, HCl, HBr, HI), when dissolved in polar solvents, will form large numbers of ions and therefore form **strong electrolytes** in water solution.

\[ \text{H}_2\text{SO}_4\text{ (aq)} \rightarrow \text{H}^+\text{ (aq)} + \text{HSO}_4^-\text{ (aq)}\rightarrow 2 \text{H}^+\text{ (aq)} + \text{SO}_4^{2-}\text{ (aq)} \]

100% ionized

**Weak acids and weak bases** (i.e., ammonia, carbonic acid) will only partially ionize (partially dissociate or react to form ions) in a polar solvent. This partial dissociation is denoted by “equilibrium” arrows: ⇌. Most of the compound remains in solution as non-ionized molecules. Since there are very few ions in solution, only a small amount of the electrical current can flow. The light bulb will light up, but only dimly, and is very dependent on the concentration of the compounds. These compounds are **weak electrolytes**.

\[ \text{H}_2\text{CO}_3\text{ (aq)} \rightleftharpoons \text{HCO}_3^-\text{ (aq)} + \text{H}^+\text{ (aq)} \]

99% molecules 1% ions

All **nonpolar covalent** and many **weakly polar covalent** compounds stay as molecules in solution and do not form ions. They are unable to carry a charge and therefore are **non-electrolytes**.

**If the solvent is nonpolar**, it does not have the electrostatic forces (ion/dipole interactions) to stabilize ions and so none form. Therefore, in a non-polar solvent, all compounds, which are added to it, are non-electrolytes.

Also, a solvent may be very polar, but if there is no compound added to it to provide ions or support the formation of ions, the solvent will not conduct electricity.

**Summary of Bond Types**

1. Ionic: Contains both metals and nonmetals, or the ammonium (NH₄⁺) ion. Elements are greatly separated on the Periodic Table.

2. Covalent: Contains only nonmetals (including hydrogen).
   a. Polar Covalent: typically contains H-O or H-N bonds, strong acids. Elements have intermediate separation on the periodic table (separated by two or three other elements).
   b. Nonpolar Covalent: typically contains lots of H-C bonds, or all atoms are the same element. Elements are right next to each other on the Periodic Table.

**Summary of Electrolytes in Water**

1. Strong electrolytes: a. soluble ionic compounds that **dissociate** (almost) **completely** into ions.
   b. strong bases (NaOH, LiOH, KOH, RbOH, CsOH, Ca(OH)₂, Ba(OH)₂, and Sr(OH)₂ **dissociate** almost **completely** into ions.
   c. strong acids (polar covalent molecules) that **dissociate** almost **completely** into ions.

2. Weak electrolytes: a. polar molecules that **dissociate partially** into ions, eg., weak acids, weak bases.
   b. slightly soluble salts that **dissociate partially** into ions.

3. Non-electrolytes: a. nonpolar (covalent) compounds, that **do not dissociate** into ions.
   b. very weakly polar compounds, that **do not dissociate** into ions.
In Part 4 of the experiment, you will observe what happens when the electrolyte, Ba(OH)$_2$, is reacted with a second electrolyte, H$_2$SO$_4$. The reaction between these two compounds is an acid/base reaction and a precipitation reaction that produces water and the insoluble salt, barium sulfate.

**word equation:**

barium hydroxide + sulfuric acid $\rightarrow$ barium sulfate + water

**molecular equation:**

Ba(OH)$_2$ (aq) + H$_2$SO$_4$ (aq) $\rightarrow$ BaSO$_4$ (s) + 2 H$_2$O (l)

**strong base**  strong acid  **insoluble salt**  water

strong electrolyte  strong electrolyte  non-electrolyte  non electrolyte

**ionic equation:**

Ba$^{+2}$ (aq) + 2 OH$^{-}$ (aq) + 2 H$^{+}$ (aq) + SO$_4^{2-}$ (aq) $\rightarrow$ BaSO$_4$ (s) + 2 H$_2$O (l)

**net ionic equation:**

Ba$^{+2}$ (aq) + 2 OH$^{-}$ (aq) + 2 H$^{+}$ (aq) + SO$_4^{2-}$ (aq) $\rightarrow$ BaSO$_4$ (s) + 2 H$_2$O (l)

Water is polar covalent and forms very few ions in solution. Also, since BaSO$_4$ is insoluble in water, it is a non-electrolyte. Therefore, there is a point in the titration when equal stoichiometric amounts of H$_2$SO$_4$ and Ba(OH)$_2$ will be added and there are no ions to carry any charge. At this point the light will go out. Addition of more H$_2$SO$_4$ after this point will supply more ions. Since there is no more Ba(OH)$_2$ available to react with these ions, the H$^+$ and HSO$_4^{-}$ ions are available to carry the current.

**Experimental**

**CAUTION:**

1) It is extremely important to use safety goggles for this experiment and to use caution in the handling of acids and bases.

2) It is also important to **unplug the conductivity tester when cleaning the electrodes or when not using the tester** so that your body does not succeed in completing the circuit which will result in a nice jolt of electricity coursing through your body!

**Procedure**

**Technique tip:** Rinse the electrodes between solutions by unplugging the conductivity tester and using a water wash bottle to rinse the electrodes into a waste beaker.

1. Obtain a set of electrolyte solution wide-mouth jars. You will use these solutions as-is, there is no need to transfer them to another container. Do not throw out the solutions when you are finished, return them to the supply bench. **Please share these solutions with other groups! Do not cross-contaminate these solutions!** Obtain a water wash bottle and a medium size beaker large enough to fit under the conductivity tester, to use as a ‘Waste Beaker’. Set up the conductivity tester at your bench (shown in Figure 1), to determine if the compound behaves as an electrolyte or a non-electrolyte in that solution. Make sure to **completely** submerge the electrodes into the solutions before plugging in the conductivity tester. Also, using your observations and rules in Appendix A attached and/or the Laboratory Handbook, Appendix B, Tables 3 and 4, determine whether each compound is ionic, polar covalent or non-polar covalent. The structures of ammonia, toluene, acetic acid and methanol are:

- **Ammonia**
- **Toluene**
- **Acetic Acid**
- **Methanol**
Data:

<table>
<thead>
<tr>
<th>Solution Tested</th>
<th>Conductivity</th>
<th>Type of Bonds in Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>(name)</td>
<td>Formula</td>
<td>(strong, weak or non)</td>
</tr>
<tr>
<td>0.1 M sodium chloride</td>
<td></td>
<td>(ionic, polar covalent, or nonpolar covalent)</td>
</tr>
<tr>
<td>0.1 M methanol (CH₃OH)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.1 M sodium hydroxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.1 M barium hydroxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.1 M ammonia (NH₃ or NH₄OH)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.1 M hydrochloric acid</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.2 M acetic acid (CH₃COOH)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.1 M sulfuric acid</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

For each solution tested, write a chemical equation representing the dissociation behavior of each compound in water solution. If no dissociation occurs, write NR for No Reaction. Include charges and phase labels. The first two are started for you as an example of how to structure your equations.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium chloride</td>
<td>NaClₐq) →</td>
</tr>
<tr>
<td>Methanol</td>
<td>CH₃OHₐq) →</td>
</tr>
<tr>
<td>sodium hydroxide</td>
<td></td>
</tr>
<tr>
<td>barium hydroxide</td>
<td></td>
</tr>
<tr>
<td>ammonia (a weak base)</td>
<td>NH₄OHₐq) ⇌</td>
</tr>
<tr>
<td>hydrochloric acid</td>
<td></td>
</tr>
<tr>
<td>acetic acid</td>
<td>CH₃COOHₐq) ⇌</td>
</tr>
<tr>
<td>sulfuric acid</td>
<td></td>
</tr>
</tbody>
</table>
2. In the fume hood, ~20 mL of glacial acetic acid (100% hydrogen acetate) is already in a dry 100 mL beaker. Using this solution, and the accompanying lightbulb set-up, perform the conductivity test. **Caution:** Glacial acetic acid has harmful fumes. Avoid breathing vapors and avoid contact with skin. Obtain a your own 100-mL beaker and fill it with ~ 25 mL distilled water. Now pour about 10 mL of the glacial acetic acid into that beaker and perform the conductivity test. Using a squirt bottle add more distilled water to the beaker while testing the conductivity. Note what happens as you add more water. Dispose of in the appropriate electrolytes waste container.

**Questions:**

a. Is glacial acetic acid (i.e., 100% acetic acid) a good conductor?    Yes          No

b. Does the diluted acetic acid give a greater conductivity than the pure compound?

Greater conductivity    Less conductivity    About the same

c. Explain what happens to the glacial acetic acid when water is added.

3. The solution is already made up and is in a labeled container. In the fume hood, test the conductivity of hydrogen chloride dissolved in toluene. **Do not dump out or contaminate this solution.** Leave the solution for the next group.

a. Was the HCl in toluene an electrolyte?    Yes          No

b. Was the HCl in water (part 1 of the lab) an electrolyte?    Yes          No

b. Therefore, what happens to the HCl when it is dissolved in water?

_________________________________________________________________

_________________________________________________________________

_________________________________________________________________

_________________________________________________________________

c. In toluene, does the HCl exist as ions or as molecules – how do you know?

_________________________________________________________________

_________________________________________________________________

_________________________________________________________________
4. Obtain a magnetic stirrer (under the hoods), magnetic stirbar (the stirbar looks like a giant white “pill” and is in a beaker on the stock bench – if you can’t find any there is a stir bar “lifeguard” which looks like a long white wand hanging by the fume hood. Obtain the “lifeguard”, stick it in the titration waste container, swirl it around, and you will rescue a stir bar from drowning! – rinse it off into the waste container with distilled water), buret, buret clamp, conductivity tester and a 250 mL beaker to perform this part of the experiment. Set up the apparatus with the beaker on the magnetic stirrer, with the stir bar inside (Figure 2). You will be using reagents from large stock bottles located on the lab bench. Do NOT use the same solutions that were used for the conductivity tests. Place in the beaker about 30 mL of ~ 0.1 M barium hydroxide solution (from the large stock bottle, record the actual amount below) and dilute it with 50 mL of distilled water. Fill the buret full of dilute, ~0.1 M, sulfuric acid (again, the sulfuric acid is in a large stock bottle. Be sure to rinse the buret with water and 2 small portions of the sulfuric acid solutions before titrating). Remember to fill the tip of the buret. Suspend the buret using the buret clamp and ringstand so that you can titrate (slowly add) the sulfuric into the beaker of barium hydroxide. Start the apparatus so that a mild swirling action occurs. Immerse the electrodes of the conductivity tester into the barium hydroxide solution. You will be titrating the acid into the base once the experiment begins (answer questions 1 and 2 below before beginning!). You will be testing the conductivity continuously while titrating. **Record your observations about the solution and conductivity in each section below.**

**Figure 2**

**Questions:**
Describe in words and write net ionic reactions to demonstrate the electrolyte behavior, include all ions and molecules present. Show any dissociation reactions. **Be sure to include phases!!**

Volume of Ba(OH)₂ _______ mL in beaker.

1. Write a chemical **equation** representing the dissociation behavior of Ba(OH)₂ in water

Would you expect Ba(OH)₂ to conduct electricity – why or why not?:

2. Write a chemical **equation** representing the dissociation behavior of H₂SO₄ in water

Would you expect H₂SO₄ to conduct electricity- why or why not?:

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2. Add approximately 15 – 18 mL of H₂SO₄ to the solution in the beaker. Describe what happens to the solution, and to the light bulb, after 15-18 mL of H₂SO₄ is titrated into the beaker. Write equation(s) to describe any reaction that is taking place. In the description section, address the following: What happened to the light bulb? Is the light bulb still on? Is it as bright? Why is the light bulb still on? What species are present in solution? Do you expect them to be electrolytes, if yes, how does that correlate with the light bulb being on?

Equation(s): 

Description:

3. Add additional sulfuric acid until the light bulb goes out. This is the “endpoint” of the titration. In the description section, be sure to address the following: Why did the light bulb go out? What species are actually present in solution? Do you expect them to be electrolytes, if yes, how does that correlate with the light bulb being on/off. If no, how does that correlate with the light bulb being on/off?

Volume of H₂SO₄ when the light goes out: ______________

Equation(s): 

Description:

4. Add more sulfuric acid until the light bulb comes back on brightly. In your description, be sure to address the following: what species caused the light bulb to turn back on? What species are actually present in the beaker? Do you expect any of them to be electrolytes?

Volume H₂SO₄ when light is bright: ______________

Equation(s): 

Description:
Homework Problems – these can be done at home – they do not need to be completed to get a final stamp!

1. A solution of acetic acid is made by dissolving 0.100 mole of CH₃COOH, in enough water to make one liter of solution. Conductivity measurements show that the acetic acid is 1.34% ionized. Remember: Avogadro's number is 6.022 x 10²³ molecules (or ions, or atoms, or students, or anything)/mole. Write the dissociation equation for acetic acid in water and then calculate \(a\) and \(b\).

**Equation:**

a. If 1.34% of the acetic acid is ionized, what % was not ionized and remains as molecules?

answer. _______________

b. Calculate, using stoichiometry and your knowledge of %, the number of acetate ions, CH₃COO⁻ \((\text{aq})\)

answer. _______________

c. Using your answer in part b, calculate the number of hydrogen acetate molecules, CH₃COOH, in the solution.

answer. _______________

2. How many mL of 0.200 M H₂SO₄ solution must be added in order to react completely with 25.0 mL of 0.100 M NaOH?

**Balanced Equation:**

**Calculations:**

answer. ______________ mL
1. Name the following compounds. Classify the following compounds as containing polar covalent, nonpolar covalent, or ionic bonds. What is the conductivity in solution (strong, weak, or nonelectrolyte)? See the flowchart on p. 2 of the lab for help!

<table>
<thead>
<tr>
<th>Name</th>
<th>Type of bond</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaBr</td>
<td></td>
<td></td>
</tr>
<tr>
<td>LiOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HBr (aq)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HF (aq)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>N₂</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. Name the following ionic compounds. Classify them as soluble or insoluble in water using your solubility guidelines. What is the conductivity in solution (strong or non-electrolyte)?

<table>
<thead>
<tr>
<th>Name</th>
<th>Soluble/Insoluble</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaBr</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ba(OH)₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca₃(PO₄)₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgCl₂</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

3. Consider the following titration reaction: \( \text{HCl (aq)} + \text{AgNO₃ (aq)} \rightarrow \text{AgCl (s)} + \text{HNO₃ (aq)} \)

   a. Write the ionic equation for the reaction.

   b. Write the net-ionic equation for the reaction.

   c. At the equivalence point, when exactly 0.050 mol of HCl (aq) has reacted with 0.050 mol of AgNO₃ (aq), would you expect the solution to conduct electricity? Explain your choice.

4. Why is it important to unplug the conductivity tester while cleaning the electrodes?