Objectives

- To observe electrical conductivity of substances in various aqueous solutions
- To determine if the solution is a strong or weak electrolyte
- To interpret a chemical reaction by observing aqueous solution conductivity.

Electrical conductivity is based on the flow of electrons. Metals are good conductors of electricity because they allow electrons to flow through the entire piece of material. Thus, electrons flow like a “sea of electrons” through metals. In comparison, distilled water is a very poor conductor of electricity since very little electricity flows through water. Highly ionized substances are **strong electrolytes**. Strong acids and salts are strong electrolytes because they completely ionize (dissociate or separate) in solution. The ions carry the electric charge through the solution thus creating an electric current. The current, if sufficient enough, will light one or both LEDs on a conductivity meter, shown at right.

Slightly ionized substances are **weak electrolytes**. Weak acids and bases would be categorized as weak electrolytes because they do not completely dissociate in solution.

Substances that do not conduct an electric current are called **non-electrolytes**. Non-electrolytes do not ionize; they do not contain moveable ions. The LEDs of a conductivity meter will not light because there are no ions to carry the electric current. The table below lists examples of strong, weak and non-electrolytes.

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**Strong Electrolytes**

**Strong Acids**

- Hydrochloric acid
- Hydrobromic acid
- Hydroiodic acid
- Nitric acid
- Sulfuric acid
- Perchloric acid
- Chloric acid

**Strong Bases**
Sodium hydroxide
Potassium hydroxide
Calcium hydroxide
Barium hydroxide

**Soluble Salts**

Sodium chloride
Potassium carbonate
Copper(II) sulfate

**Weak Electrolytes**

**Weak Acids**

Acetic acid
Carbonic acid
Citric acid
Phosphoric acid

**Weak Bases**

Ammonia
Ammonium hydroxide
Magnesium hydroxide

Most other bases

**Slightly Soluble Salts**

Silver chloride

Calcium carbonate

Barium sulfate

**Non-Electrolytes**

Distilled water

Methanol

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**Procedure**

**Materials and Equipment**

conductivity meter, wash bottle with distilled water, large beaker for rinsing/waste, small beakers, Kimwipes, solid sodium chloride, solid calcium carbonate

*Solutions*: acetic acid, aluminum nitrate, ammonium hydroxide, calcium hydroxide, citric acid, ethanol, hydrochloric acid, magnesium hydroxide, magnesium sulfate, nitric acid, potassium iodide, sodium chloride, sodium hydroxide, sucrose

**Safety**

Be cautious with hydrochloric acid, nitric acid, sulfuric acid and concentrated acetic acid. Although low in concentration, some individuals may have extreme skin sensitivities. If you experience any tingling sensations or skin discolorations, rinse immediately with large amounts of water for 15 minutes. Inform your instructor ASAP.

*Personal Protective Equipment (PPE)* required: lab coat, safety goggles, closed-toe shoes
Conductivity Testing – Evidence for Ions in Aqueous Solution

1. The meter has a 9V battery, and two parallel copper electrodes. Use a wash bottle with distilled water and a large beaker labeled “waste” to rinse the copper electrodes. Dry using a Kimwipe tissue. When switched on, the lights should not be lit any color. If they are, repeat the rinsing and drying.

Note

DO NOT EXPOSE THE CIRCUIT BOARD TO WATER. Only the copper electrodes should be rinsed with water.

2. Place the meter so that the circuit board is facing up (the battery will be below). Always place the meter in this way so that the circuit board will not get wet. On this side, there is a guide to the possible conductivity measurements:

<table>
<thead>
<tr>
<th>Scale</th>
<th>Red LED</th>
<th>Green LED</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>off</td>
<td>off</td>
<td>low or none</td>
</tr>
<tr>
<td>1</td>
<td>dim</td>
<td>off</td>
<td>low</td>
</tr>
<tr>
<td>2</td>
<td>medium</td>
<td>off</td>
<td>medium</td>
</tr>
<tr>
<td>3</td>
<td>bright</td>
<td>dim</td>
<td>high</td>
</tr>
<tr>
<td>4</td>
<td>very bright</td>
<td>medium</td>
<td>very high</td>
</tr>
</tbody>
</table>

Note

Switch the meter on and dip the copper electrodes to test conductivity. Thoroughly rinse with distilled water after each test, and dry with Kimwipes. Switch the meter off between uses.

3. Place 5 mL of **distilled** water into a small, clean beaker. Test and record your results.

4. Place 5 mL of **tap water** into a small, clean beaker. Test and record your results.

5. Place about 0.2 g of solid sodium chloride (\(\text{NaCl}\)) into a small, clean beaker and test the conductivity. Add 5 mL distilled water to the sodium chloride; test the conductivity of the solution. Dispose of this solution in the sink and rinse the beaker.

6. Place about 0.2 g of solid calcium carbonate (\(\text{CaCO}_3\)) into a small, clean beaker and test the conductivity. Add 5 mL distilled water to the calcium carbonate; test the conductivity of the solution. Dispose of this solution in the sink and rinse the beaker.
7. Use 5 mL of each of the following in 100-mL beaker to test the conductivities.

**Be sure to rinse and dry the electrodes between tests, using your wash bottle with waste beaker, and Kimwipes.**

Dispose the solution and rinse the beaker in the sink between tests. Dispose the waste beaker solution in non-hazardous waste in the hood.

- acetic acid, 0.1 M \( \text{\ce{HC2H3O2}} \)
- aluminum nitrate, 0.1 M \( \text{\ce{Al(NO3)3}} \)
- ammonium hydroxide, 0.1 M \( \text{\ce{NH4OH}} \) (aq)
- calcium hydroxide, saturated \( \text{\ce{Ca(OH)2}} \)
- citric acid, 0.1 M \( \text{\ce{C6H8O7}} \)
- ethanol, \( \text{\ce{CH3CH2OH}} \)
- hydrochloric acid, 0.1 M \( \text{\ce{HCl}} \)
- magnesium hydroxide, saturated \( \text{\ce{Mg(OH)2}} \)
- magnesium sulfate, 0.1 M \( \text{\ce{MgSO4}} \)
- nitric acid, 0.1 M \( \text{\ce{HNO3}} \)
- potassium iodide, 0.1 M \( \text{\ce{KI}} \)
- sodium chloride, 0.1 M \( \text{\ce{NaCl}} \)
- sodium hydroxide, 0.1 M \( \text{\ce{NaOH}} \)
- sucrose, 0.1 M \( \text{\ce{C12H22O11}} \)
# Lab Report: Electrical Conductivity of Aqueous Solutions

## Conductivity Testing - Evidence for Ions in Aqueous Solution

<table>
<thead>
<tr>
<th>Solution</th>
<th>Observations: red LED</th>
<th>green LED</th>
<th>Conductivity</th>
<th>Ionized, Partially ionized, or Non-ionized</th>
</tr>
</thead>
<tbody>
<tr>
<td>examples: (\ce{LiOH}) (aq), (\ce{HNO2}) (aq), methanol (l)</td>
<td>red bright, green dim red dim, green off red off, green off</td>
<td>high/low/none</td>
<td>strong electrolyte/weak electrolyte/non-electrolyte</td>
<td>ionized/partially ionized/non-ionized</td>
</tr>
<tr>
<td>distilled water, (\ce{H2O}) (l)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>tap water, (\ce{H2O}) (l)</td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>solid sodium chloride, (\ce{NaCl}) (s)</td>
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<td></td>
</tr>
<tr>
<td>sodium chloride solution, (\ce{NaCl}) (aq)</td>
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<tr>
<td>solid calcium carbonate, (\ce{CaCO3}) (s)</td>
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<tr>
<td>calcium carbonate solution, (\ce{CaCO3}) (aq)</td>
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<td></td>
</tr>
<tr>
<td>Solution</td>
<td>Observations: red LED</td>
<td>green LED</td>
<td>Conductivity</td>
<td>Electrolyte Type</td>
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<td>----------</td>
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<tr>
<td>acetic acid, (\ce{HC2H3O2}) (aq)</td>
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<tr>
<td>aluminum nitrate, (\ce{Al(NO3)3}) (aq)</td>
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<td>ammonium hydroxide, (\ce{NH4OH}) (aq)</td>
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<tr>
<td>calcium hydroxide, (\ce{Ca(OH)2}) (aq)</td>
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<tr>
<td>carbonic acid, (\ce{H2CO3}) (aq)</td>
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<tr>
<td>ethanol, (\ce{CH3CH2OH})</td>
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<tr>
<td>hydrochloric acid, (\ce{HCl}) (aq)</td>
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</tr>
<tr>
<td>magnesium hydroxide, (\ce{Mg(OH)2}) (aq)</td>
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<tr>
<td>magnesium sulfate, (\ce{MgSO4}) (aq)</td>
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<tr>
<td>nitric acid, (\ce{HNO3}) (aq)</td>
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<tr>
<td>potassium iodide, (\ce{KI}) (aq)</td>
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<td></td>
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<tr>
<td>sodium chloride, (\ce{NaCl}) (aq)</td>
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</tr>
</tbody>
</table>
sodium hydroxide, \(\text{NaOH}\) (aq)

sucrose, \(\text{C12H22O11}\) (aq)

**Questions**

1. Why must the electrodes on the conductivity apparatus, as well as all the beakers, be rinsed with distilled water after each conductivity test?

2. Why is distilled water a weaker conductor than tap water?

3. Why does solid sodium chloride act as a non-electrolyte while an aqueous while an aqueous \(\text{NaCl}\) solution acts as a strong electrolyte?

4. Classify each of the following as **non-ionized**, **partially-ionized**, or **ionized**.

   Write each compound as it exists in aqueous solution e.g. \(\text{NaCl} (aq) \rightarrow \text{Na}^+ (aq) + \text{Cl}^- (aq)\)

   - \(\text{HCl} (aq)\) – a strong acid
   - \(\text{Ca(OH)2} (aq)\) – a strong base
   - \(\text{HC2H3O2} (aq)\) – a weak acid
   - \(\text{Ba(OH)2} (aq)\) – a weak base

5. For the chemical reaction

   \[\text{H2SO4 (aq) + 2 NaOH (aq) \rightarrow Na2SO4 (aq) + 2 H2O (l)}\]

   Write the complete ionic equation:

   Write the net ionic equation:

6. For the chemical reaction

   \[\text{KNO3 (aq) + NaCl (aq) \rightarrow NaNO3 (aq) + KCl (aq)}\]

   Write the complete ionic equation:

   Write the net ionic equation: