Electrochemistry: Electrolysis of Solutions
Student Activity

Introduction and Historical Context

We know from the work of Svante August Arrhenius in the late 1 800s that aqueous solutions of ionic compounds are electrolytes: these solutions carry an electric current. The positive and negative ions that are formed when the ionic compounds break apart in water are able to move freely about in solution carrying an electric current. In this lab exercise you will be using the process of electrolysis: an electric current in an ionic solution that causes some of the ions to change chemically through the gain or loss of electrons. This electrolysis process is powered by an electric source, in this case a 9-volt battery to which are attached two electrodes—a positive electrode, or anode, and a negative electrode, or cathode. When the circuit is completed ("turned on"), current flows from one terminal of the battery through the conducting solution (ions) back to the other terminal of the battery. During this current flow, electron changes take place at the electrodes in the solution, producing new species. In other words, a non-spontaneous chemical reaction takes place. It is being driven by an energy source: the battery.

Purpose

In this lab exercise students will cause an electrolytic reaction and observe the results through chemical testing of the products in order to explain what is happening at each electrode of the electrolysis apparatus.

Safety

- Goggles must be worn throughout this activity.
- After working with the solutions and cleaning up, students need to wash their hands.

Materials and Apparatus

Consumables

Solutions (0.1 M) of the following:

- Sodium sulfate (Na_2SO_4)
- Sodium chloride (NaCl)
- Copper (II) sulfate ($CuSO_4$)
- Lead nitrate ($Pb[NO_3]_2$)
- Potassium iodide (KI)
- Potassium bromide (KBr)
- Starch solution
- Universal indicator
- Distilled water
- Cloth or filter paper strips

Non-consumables

- Nine-volt battery (or conductivity apparatus with nine-volt battery)
- Battery clip with leads and attached alligator clips
- Pencil "lead" (carbon)

- Microplate (12 wells)
- Small-scale pipets

Pre-Lab Questions

1. Does “pure water” (distilled) conduct electricity? Test this.
2. Does a solution formed from the dissolving of sugar in water conduct electricity? Test the solution and explain your results in terms of Arrenhius’s ideas.
3. Does a solution formed from the dissolving of table salt in water conduct electricity? Test the solution and explain your results in terms of Arrenhius’s ideas.
4. Using equations, show why there is a difference in conductivity between the solution with the dissolved dissociated sugar and that with the dissolved table salt.
5. If an ion loses electrons, is it oxidized or reduced?
6. If an ion gains electrons, is it oxidized or reduced?

Procedure

1. For each of the reactions you will be conducting in your microplate, identify three adjoining “wells” that will serve as a control and a pair of reacting cells. Using a dropping pipet, add 6 to 10 drops of any one solution to each of the three wells. Connect each pair of reacting wells with either a small piece of paper towel or a cotton cloth strip. Add drops of the reacting solution to the paper or cloth connector that completes the “circuit.”
2. Add a drop of universal indicator to each of the three wells. Note the starting color in each well and what pH it represents.
3. Attach the two leads from your 9-volt battery to the two pencil “lead” carbon electrodes. Note which electrode is connected to the positive lead of the battery and which is connected to the negative lead. When reactions take place in the wells, note which well is connected to which lead (electrode).
4. Place one of the carbon electrodes in each of two wells. Observe and record all activity over a 5-minute interval before removing the electrodes. Note any changes on the electrodes themselves.
5. For starters, practice using the setup described in step #1, but use distilled water. Moisten the connecting strip with distilled water. Add a drop of universal indicator to all three wells, then add the electrodes to two of the three wells and run the electrolysis for 5 minutes. Observe all changes and record the type of activity at each electrode (positive and negative). Is there evidence of change? Is a reaction taking place at either electrode? How would you explain the “results”?
6. Repeat step # 5 but add a drop of sodium sulfate (Na_2SO_4) to each of the three wells containing distilled water. Place the paper or cloth connector strip in one of the wells to absorb some of the solution. Again, use the strip to physically connect two of the wells. Now place the electrodes into two of the three wells and observe any and all changes for 5 minutes. How would you explain these results compared with what happened in step #5?
7. If the word *electrolysis* means to “cut” using electricity, what would result from the cutting of water (H_2O)? What evidence from your observations do you have for the products of “cutting”? How do you account for the color changes in any of the cells? (Consult a chart of redox [oxidation or reduction] potentials.)
8. Prepare another three wells for a different electrolytic reaction. Connect two of the reacting wells with a new paper or cloth connector soaked with the reacting solution,

in this case, potassium iodide (KI). Add universal indicator to each of the three wells. With the two electrodes in the two reacting wells, make the electrical connections and run the system for 5 minutes, observing all changes for each of the electrodes (positive and negative).

9. After disconnecting the battery and removing the electrodes, add a few drops of starch solution to each of the wells and observe any changes. What is the starch solution testing for? At which electrode(s) is there a positive test?
10. Prepare the next electrolysis setup. Add the sodium chloride (NaCl) solution to each of the three wells, and use a connector soaked in the same solution. Add a drop of universal indicator to each well. With the electrodes in place, run the electrolytic process for 5 minutes, noting any changes at each electrode.
11. Repeat the electrolysis process using the potassium bromide (KBr) solution. Record your observations at each electrode.
12. Repeat the electrolysis process using the copper sulfate ($CuSO_4$) solution. Record your observations at each electrode.
13. Repeat the electrolysis process using the lead nitrate ($Pb[NO_3]_2$) solution. Record your observations at each electrode.

Post-Lab Questions

1. How do you explain the results in your first electrolysis using distilled water? What does this tell you about the presence or absence of ions in “pure” water?
2. What happens when you add drops of sodium sulfate (Na_2SO_4) to the distilled water? What changes occur in the water through the addition of sodium sulfate?
3. Observation of bubbles would indicate the presence of a gas. What gases could be produced from the breakup (“lysis”) of water, assuming sodium sulfate in solution is not changed?
4. What pH category (acidic, basic) is shown by color changes for the universal indicator in each of the reacting wells containing water? Knowing that water contains hydrogen (H^+) and hydroxide (OH^-) ions in equal numbers when neutral (pH = 7), what must have happened in each of the reacting wells to have produced the particular pH changes noted through the color changes of the universal indicator? Identify each pH change (acidic, basic) with each type of electrode (positive, negative).
5. Consult a redox chart (oxidation or reduction potential) to find the half-reaction equations associated with the redox (by electrolysis) of water. Relate each of the half-reaction products to your observations in #3 above.
6. Using the information from #5, relate the products of each half reaction to the pH changes noted in #4.
7. Explain your observations for the electrolysis of the potassium iodide (KI) solution. Note that based on redox potentials for K^+, I^-, and H_2O, potassium ions are not reduced. That leaves iodide ions and water to react. Write a half-reaction equation for the oxidation of iodide ions. Does the product match your observations? At which electrode does the oxidation take place?
8. Using the information from question #5 above for the half reactions of water electrolysis, how do you account for the pH change at the negative electrode (cathode)? Write the appropriate half reaction and circle the product in the equation responsible for the pH change.
9. What does the starch react with in the KI electrolysis? Write the chemical formula for the substance that starch reacts with.
10. 10.If the chloride ion of NaCl reacts in the same manner as the iodide ion of KI

because they are from the same family, write a half-reaction equation for the change that occurs to the chloride ion (Cl^-) during electrolysis. At which electrode did this change take place? What is your evidence from your observations?

11. 11.As with the KI electrolysis, water electrolysis takes place in the NaCl solution rather than sodium electrolysis. Write the appropriate half reaction for water electrolysis, based on the information from question #5. Does this correlate with the pH change observed at the negative electrode (cathode)?
12. 12.As with KI and NaCl electrolysis, the electrolysis of potassium bromide (KBr) produces similar results. Write the half-reaction equations for the electrolysis of Br^- and H_2O. What is your evidence for the formation of Br_2 and the electrolysis of water?
13. What changes did you observe when you did electrolysis of the copper sulfate ($CuSO_4$) solution? Relate each change to a particular electrode (positive, negative).
14. Write half-reaction equations for each observation in #13. Relate each change to a specific electrode. Remember to relate any pH changes to your observations. Assume electrolysis of water rather than that of the sulfate (SO_4^{2-}) ion took place at one of the electrodes.