AP Chemistry By Satellite
Laboratory Manual
Instructor's Edition

EXPERIMENT 6: CONDUCTIVITY

Contents Page

Reagents

CHEMICAL	Quantity	CHEMICAL	Quantity
acetic acid, 0.1 M	15 mL	methanol	15 mL
acetic acid, 1.0 M	50 mL	nickel chloride hexahydrate, 0.1 M	15 mL
ammonium acetate, 0.1 M	15 mL	nickel nitrate hexahydrate, 0.1 M	15 mL
ammonium chloride,0.1 M	15 mL	nitric acid, 0.1 M	15 mL
ammonium hydroxide, 0.1 M	15 mL	potassium chloride, 0.1 M	15 mL
ammonium sulfate, 0.1 M	15 mL	potassium iodide	1 g
ammonium thiocyanate	1 g	potassium iodide, 0.1 M	15 mL
barium hydroxide octahydrate, 0.1 M	15 mL	potassium nitrate, 0.1 M	15 mL
calcium carbonate (ppt. chalk)	9 g	silver nitrate, 0.1 M	15 mL
calcium hydroxide, sat'd	15 mL	sodium acetate trihydrate, 0.1 M	15 mL
cobalt chloride hexahydrate, 0.1 M	15 mL	sodium carbonate	1 g
cupric chloride dihydrate, 0.1 M	15 mL	sodium chloride	1 g
ethanol	15 mL	sodium chloride, 0.1 M	15 mL
ferric nitrate nonohydrate	1 g	sodium hydroxide, 0.1 M	15 mL
hydrochloric acid, 1.0 M	50 mL	sucrose	1 g
hydrochloric acid, 0.1 M	5 mL	sulfuric acid, 0.5 M	50 mL
lead (II) nitrate	1 g	sulfuric acid, 0.1 M	5 mL
lead (II) nitrate, 0.1 M	15 mL		

Solution Preparation

Use the table below to prepare the solutions of acids and bases. For general instructions on how to prepare solutions of dilute acids (or bases) from more concentrated solutions, see Appendix III. Some of the volumes of concentrated acid may be too small to measure accurately. It is recommended larger amounts of the dilute solutions be prepared. These reagents can be used in subsequent experiments.

Acid or base	Conc. of dilute acid or base	Vol. of dilute acid or base	Conc. of concentrated acid or base	Vol. of concentrated acid or base
NH_3 (ammonium hydroxide)	0.1 M	15 mL	6 M	0.25 mL
HCl (hydrochloric acid)	1.0 M	50 mL	12 M	4.2 mL
HCl (hydrochloric acid)	1.0 M	50 mL	6 M	8.3 mL
HCl (hydrochloric acid)	0.1 M	5 mL	1 M	0.5 mL
H_2SO_4 (sulfuric acid)	0.5 M	50 mL	18 M	1.4 mL
H_2SO_4 (sulfuric acid)	0.5 M	50 mL	3 M	8.3 mL
H_2SO_4 (sulfuric acid)	0.1 M	5 mL	0.5 M	1 mL
HNO_3 (nitric acid)	0.1 M	15 mL	16 M	0.1 mL
HNO_3 (nitric acid)	0.1 M	15 mL	6 M	0.25 mL
$HC_2H_3O_2$ (acetic acid)	1.0 M	50 mL	17.3 M	2.89 mL
$HC_2H_3O_2$ (acetic acid)	0.1 M	15 mL	1 M	1.5 mL

Use values from the table below to prepare the solutions of the salts needed in this experiment. (See Appendix I & III. for more detail.) To facilitate dispensing these solutions, each should be placed in dropper bottles. Since the experiment is performed at the microscale level students will only require a few drops of each solution. **15 mL is an estimate of the amount of each solution needed for every 5 students.**

Salt	Conc. of salt	Weight of salt	Vol. of solution
$NH_4C_2H_3O_2$ (ammonium acetate)	0.1 M	0.116 g	15 mL
NH_4C1 (ammonium chloride)	0.1 M	0.080 g	15 mL
$(NH_4)_2SO_4$ (ammonium sulfate)	0.1 M	0.198 g	15 mL
$Ba(OH)_2 \cdot 8H_2O$ (barium hydroxide octahydrate)	0.1 M	0.473 g	15 mL
$Ca(OH)_2$ (calcium hydroxide)	0.1 M	0.111 g	15 mL
$CoCl_2 \cdot 6H_2O$ (cobalt (II) chloride hexahydrate)	0.1 M	0.357 g	15 mL
$CuCl_2 \cdot 2H_2O$ (cupric chloride dihydrate)	0.1 M	0.256 g	15 mL
$Pb(NO_3)_2$ (lead nitrate)	0.1 M	0.497 g	15 mL
$NiCl_2 \cdot 6H_2O$ (nickel (II) chloride hexahydrate)	0.1 M	0.357 g	15 mL
$Ni(NO_3)_2 \cdot 6H_2O$ (nickel (II) nitrate hexahydrate)	0.1 M	0.436 g	15 mL
KCl (potassium chloride)	0.1 M	0.112 g	15 mL
KI (potassium iodide)	0.1 M	0.249 g	15 mL
KNO_3 (potassium nitrate)	0.1 M	0.152 g	15 mL
$AgNO_3$ (silver nitrate)	0.1 M	0.255 g	15 mL
$NaC_2H_3O_2 \cdot 3H_2O$ (sodium acetate trihydrate)	0.1 M	0.204 g	15 mL
NaCl (sodium chloride)	0.1 M	0.088 g	15 mL
NaOH (sodium hydroxide)	0.1 M	0.060 g	15 mL

Since the amounts of reagents used are so small, disposal problems are minimized.

The test solution for the conductivity apparatus should be 1 M HCl. Place about 15 mL in a dropper bottle labeled 'Test Solution.' Students will use this solution at the beginning of the experiment to check that their conductivity apparatus is working. If the conductivity apparatus is not working, you should replace the battery.

Equipment

PART I: Introduction to Conductivity

EQUIPMENT	Quantity
beaker, 50 mL	
conductivity apparatus	1
dropper	1
paper towels	several
petri dish	1
plate, 96-well	1
spatula or microspatula	2
watch glass	2

PART II: Strong and Weak Electrolytes

EQUIPMENT	Quantity
beaker, 50 mL	1
conductivity apparatus	1
dropper	1
paper towels	several
plate, 96-well	1

PART III: Electrolyte Strength and Reaction Rate

EQUIPMENT	Quantity
laboratory balance	1
beaker, 50 mL	2
capillary tube	1
conductivity apparatus	1
dropper	1
flask, Erlenmeyer, 250 mL	1
glass bend	1
glass tubing	1
graduated cylinder, 50 mL	1
paper towels	3
plate, 96-well	1
ring stand	2
spatula	1
stopper, 1-hole #6	1
trough or pan	1
tubing,1' rubber	1
utility clamps	2
weighing paper	3

PART IV: Chemical Reactions

EQUIPMENT	Quantity
beaker, 50 mL	1
conductivity apparatus	1
dropper	1
paper towels	several
plate, 96-well	2
spatula	1

For information on building the conductivity apparatus, see Appendix IV.

For information on obtaining well plates, see Appendix VI.

Useful Experimental Comments

Experiment Scheduling

This experiment will require approximately 5 hours to complete.

The experiment is organized to easily stop and start at almost any point. A discussion is needed at **Expl. #1** in the experiment.

It is intended the **Expl.** (explanations) portions of the experiment be completed while students are working in the laboratory. Depending on how far students progress each day, students could be assigned to complete some **Expl.** sections outside of class. However, students should not be encouraged to proceed too far beyond the uncompleted **Expl.** sections while working in the laboratory. The **Expl.** sections are intended to make the students address and formulate ideas based on their experimental observations. The experiment is long and will require a great deal of preparation time. You may decide to omit some portions of the experiment or to have students do different parts and then report their data to the rest of the class. **Part III** can be easily omitted.

College Board Recommendations: This experiment is not of the type specifically suggested by the College Board, but it does require careful recording and analysis of experimental observation. The descriptive chemistry included in the experiment and the practice in writing chemical equations should prove to be very useful to those students taking the AP Examination. The experiment is intended to familiarize students with ideas such as conductivity and net ionic equations which will be included on the examination.

The basic approach students are expected to follow in the laboratory is to do the experiments and then immediately answer the questions. If students have trouble answering questions, they should discuss the ideas with you or their classmates. Many of the important ideas will not be clear to students if they delay answering these questions.

PART I: Introduction to Conductivity

The test solution for the conductivity apparatus should be 1 M HCl. Place about 15 mL in a dropper bottle labeled 'Test Solution.' Students will use this solution at the beginning of the experiment to check that their conductivity apparatus is working. If the conductivity apparatus is not working, you should replace the battery. Remind students to clean the electrodes between tests. If the potentiometer is set too low on the conductivity apparatus, some electrolysis may occur in some of the solutions. To prevent this from happening adjust the potentiometer or use stainless steel wire for electrodes. If copper wire is used, cover the end of the wire with solder.

The first nine observations are intended to familiarize the students with the conductivity apparatus and to introduce the kind of observations they will make in the laboratory. The students should briefly summarize their observations following each experiment. Remind the students they will need to return to these observations later in the experiment.

The point of the petri dish experiment is to introduce the idea that species can migrate in water and that the migration is slow. You need to take some time to discuss the observations with your students. Careful discussion at this time will help students understand the rest of the experiment. The article by Vos and Verdonk from the *Journal of Chemical Education 62(8), 649* includes an excellent model for the needed discussion. In the experiment under Expl. #1 you will find suggested questions you may ask the students and likely responses, to guide the discussion.

To answer Explanation 3 – 6 the students must apply the ideas developed in the petri dish experiment. The students may have difficulty with the difference in conductivity of NaCl solid and NaCl in aqueous solution.

PART II: Strong and Weak Electrolytes

This series of observations and tables exposes the student to a variety of solution types. They are to measure the conductivity of each solution and then suggest the types of particles present. The concept of strong and weak electrolytes is introduced. The students are to write equations (a sample is provided) to summarize their observations. They are to then summarize the conductivity data. The students are asked to organize their conductivity data according to their observations. This exercise causes the students to look carefully at all the compounds and to develop a classification scheme for each group. In the Post–laboratory exercises they will be asked to use their scheme to predict the conductivity of several other compounds. The quality of the predictions will depend on how carefully they have worked out Explanation #6.

PART III: Electrolyte Strength and Reaction Rate

This experiment verifies **Obs #12** using a different approach. The strong acids HCl and H_2SO_4 react rapidly with $CaCO_3$ to produce CO_2, while the weak electrolyte $HC_2H_3O_2$ reacts more slowly. The rate of evolution of CO_2 depends on the concentration of hydrogen ion in solution. Because the rates of reactions are very rapid it may be difficult to judge the difference in reaction rates with the two strong acids. The weak acid reaction will be only slightly slower. The conductivity data and the rate data verify the amount of hydrogen ion in solution.

Blowing CO_2 (exhaling) into a saturated solution of $Ca(OH)_2$ is a demonstration of the removal of ions from solution by precipitation. $Ca(OH)_2$ is a strong conductor, but bubbling CO_2 into the solution precipitates $CaCO_3$ which reduces the concentration of ions in solution over time. It is helpful to keep a sample of the original $Ca(OH)_2$ in order to see the change clearly.

PART IV: Chemical Reactions

This part (**Obs. #19**) of the experiment is designed to introduce the concept of an ionic equation and a net ionic equation. The students begin by completing the table to establish the connection between particular cations and anions and their color in aqueous solution.. The order of assigning colors (according to the order in the table) to the salt solutions is important. Beginning with colorless KCl solution, students should recognize a colorless solution must mean both ions are colorless. The next solution, nickel (II) chloride, is green. Knowing that Cl^- is colorless, the students should conclude the Ni^{2+} is green. The rest of the table should be completed in a similar manner.

Next the students mix $AgNO_3$ and $NiCl_2$ and observe the change. They have already determined the ions in each of the these solutions so they should be able to write the reactant side of the equation. Based on the fact that the precipitate is white and the solution's color is pale green they should conclude that Ni^{2+} ion is in solution and the precipitate must contain Ag^+. The counterion in the precipitate must be chloride, since the nitrate salt of silver is soluble. They repeat the experiment with Co^{2+} to reinforce the ideas of color.

EXPERIMENT 6: CONDUCTIVITY

Pre-lab Questions:

Answer these questions before coming to class. They introduce you to several important ideas that you will use in this experiment. You must turn-in this exercise before you will be allowed to begin the experiment. Be sure to bring a calculator and paper to laboratory.

1. Write the formulas for the following acids:

   (a) phosphoric **H_3PO_4** ;

   (b) perchloric **$HClO_4$** ;

   (c) nitric **HNO_3** ;

   (d) sulfuric **H_2SO_4** ;

   (e) hydrochloric **HCl** ;

   (f) acetic **$HC_2H_3O_2$** .

2. Write the formulas for the following bases:

   (a) calcium hydroxide **$Ca(OH)_2$**;

   (b) potassium hydroxide **KOH** ;

   (c) sodium hydroxide **$NaOH$** ;

   (d) ammonia **NH_3**.

3. Write the formulas for the following salts:

   (a) potassium chromate **K_2CrO_4** ;

   (b) potassium sulfate **K_2SO_4** ;

   (c) copper (II) nitrate **$Cu(NO_3)_2$** ;

   (d) calcium carbonate **$CaCO_3$** ;

   (e) potassium iodide **KI** ;

   (f) lead (II) nitrate **$Pb(NO_3)_2$**

EXPERIMENT 6: CONDUCTIVITY

In this experiment (and again in Experiment 9) you will use a piece of equipment called a *well plate* shown in Figure I. This piece of equipment is made of clear plastic and contains wells used to hold solutions. Each well can be identified using a combination of a letter and a number. In the figure, well B3 (2nd row–3rd column) is labeled. A particular well will be indicated by a letter (row) and a number (column). In this experiment you will use a 96–well plate. The wells are small and only a few drops of reagent will be needed. Doing experiments on a 'microscale' is very economical and considerably safer than large scale experiments.

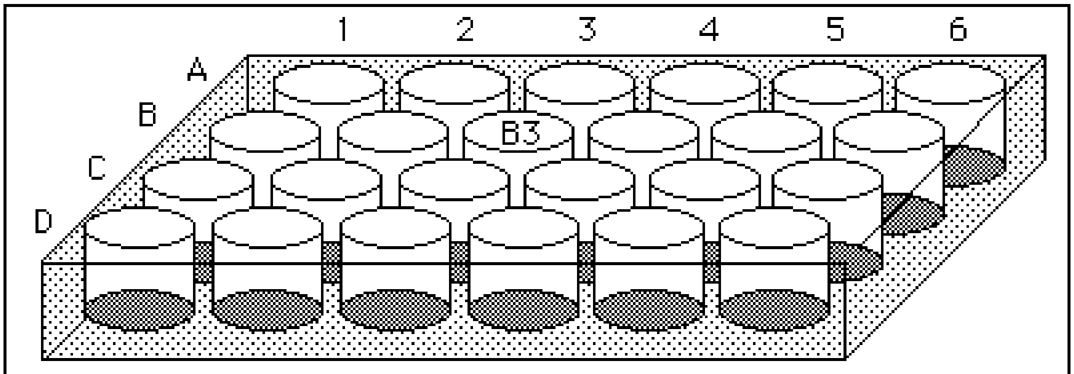


Figure I.

The conductivity apparatus used in this experiment consists of a 9–volt battery inside a 35 mm plastic film container. A light–emitting–diode (LED) has been wired to the battery. Two wires (electrodes) are attached to the LED so that if the electrodes are placed in a solution which conducts electricity the LED will glow at a particular intensity. The best way to observe the light intensity after immersing the electrodes into a conducting test solution is to view the LED from the top, not the side.

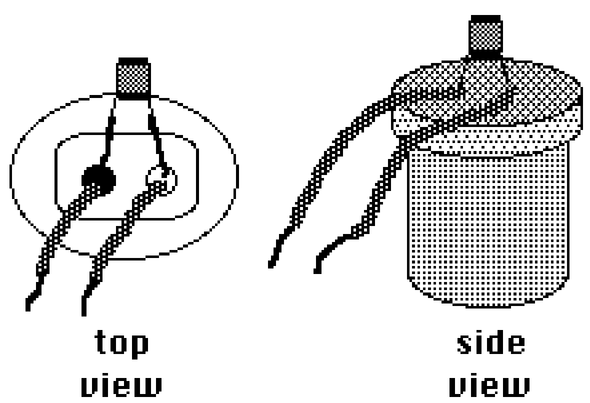


Figure II

EXPERIMENT 6: CONDUCTIVITY

EQUIPMENT:

beaker, 50 mL.........................1	petri dish.............................1
conductivity apparatus...............1	plate, 96-well.......................1
dropper.................................1	spatula or microspatula 2
paper towels..........................several	watch glass..........................2

PART I: Introduction to Conductivity

Check the conductivity apparatus using a test solution. Place 4 drops of the test solution in H12 and insert the electrodes of the conductivity apparatus into the test solution. Check that the LED (light emitting diode) glows brightly. When viewing the LED, look down from above the LED rather than from the side. It may also help to slightly darken the room. If the conductivity apparatus works, remove the electrodes from the solution and wash the electrodes with deionized water and dry with a paper towel. *It is important to wash the electrodes with deionized water following each measurement.* Try not to immerse the electrodes so deep that solution leaks underneath the plastic sheath covering the wires. If this happens clean the electrodes carefully to prevent erroneous observations.

Using the 96–well plate and a dropper, fill A1 with tap water and introduce the electrodes to a depth of about 5 mm. Note whether the LED glows brilliantly, faintly, or not at all.

Obs. #1
The LED glows faintly.

Try immersing the electrodes more and more deeply into the tap water. Record your results.

Obs. #2
The LED glows glows somewhat more brightly the deeper the electrodes are immersed into the solution.

NOTE: For each of the tests that follow, you should immerse the electrodes to approximately the same depth.

Dry the electrodes. In A3, test deionized water using the conductivity apparatus. Does the LED glow brilliantly, moderately, faintly, or not at all?

Obs. #3
The LED does not glow at all.

Test dry sucrose ($C_{12}H_{22}O_{11}$) with the conductivity apparatus. Use a microspatula to half-fill A7. Be careful not to spill sucrose in the wells surrounding A7 and be sure the electrodes are clean and dry before testing the sample.

Obs. #4
The LED does not glow.

With the electrodes still in contact with the solid sucrose in A7, add a few drops of deionized water and observe LED. Record your observations.

Obs. #5
The LED does not glow.

In A9, test dry sodium chloride (NaCl) with the conductivity apparatus. Use a dry microspatula to place the NaCl into the well. Be sure the electrodes are dry before testing the sample.

Obs. #6
The LED does not glow.

With the electrodes still in contact with the dry sodium chloride in A9, add a few drops of deionized water. Record your observations.

Obs. #7
The LED glows brilliantly.

Clean and dry the conductivity apparatus electrodes. Place 4 drops of KI solution in A11 and then test the conductivity of KI. Record your observations.

Obs. #8
The LED glows brilliantly.

Clean and dry the conductivity apparatus electrodes. Test a sample of $Pb(NO_3)_2$ solution in C1. Record your observations.

Obs. #9
The LED glows brilliantly.

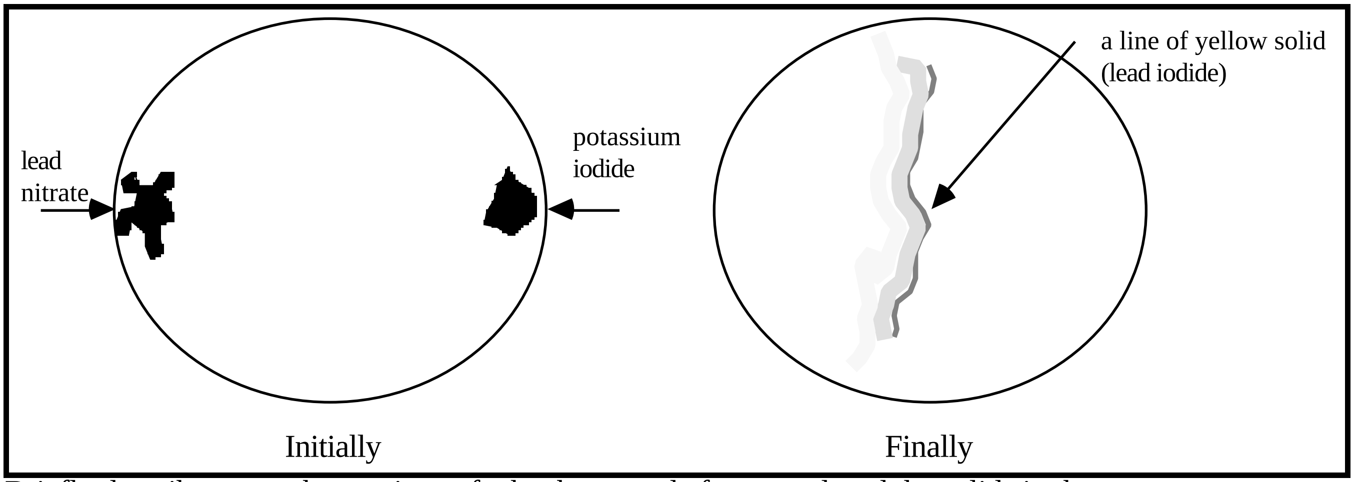
Place a small sample of solid lead nitrate on a watch glass. On a second watch glass, place a sample of solid potassium iodide.
Briefly describe the initial appearance of the dry potassium iodide and lead nitrate solids.

Obs. #10
Both compounds are white crystalline solids.

Clean a Petri dish. Cover the bottom of the dish with a thin layer of deionized water. Use a spatula to carefully place a few crystals of lead nitrate into the water close to one side of the Petri dish. Try not to agitate the water when adding the solid. Use a clean spatula to carefully add a few crystals of potassium iodide to the opposite side of the Petri dish. Do not bump the Petri dish. It is important that the water not be agitated during the experiment. Watch what happens.

Draw the arrangement of the Petri dish and the samples of lead nitrate and potassium iodide in the space below. Draw a second picture showing what happened as time passed.

Picture #1



Briefly describe your observations of what happened after you placed the solids in the dish.

Obs. #11

When the solids were added to the water, they both sank to the bottom of the petri dish. After a moment, they dissolved. For approximately 90 seconds it appeared that nothing was happening. Then a fine yellow line of crystals began to develop at the center of the petri dish. It continued to spread and widen towards the sides of the petri dish.

Do not continue experimenting until your instructor has completed the group discussion to clarify the experiment just performed. Following the discussion explain your observations. In your explanation include the use of the following terms: anion, cation, electrolyte, precipitate, soluble, hydration, and electrical conductivity.

Expl. #1

Discuss the students observations as suggested in the article by Vos and Verdonk from the *Journal of Chemical Education 62(8), 649*. Some suggestions from the article for the discussion are included below. Frequent student responses have been included in parenthesis.
How did the yellow line form? ("molecules" of lead nitrate and potassium iodide move through the water and collide in the center of the dish to form the product.) Why do the particles move through the water? How do they move? (1. When the solid dissolves, ions become hydrated and migrate through the water. The solution is an electrolyte and will conduct electricity. When the lead ions encounter iodide ions, the reaction occurs. 2. The particles are attracted to one another and want to react. They are willing to move long distances to do so.) Should answer number 2. be given, be prepared to dispute. Suggest students try an experiment in which the potassium iodide is introduced to the dish three or four minutes before the lead nitrate is added. The yellow precipitate will form immediately. Students should then abandon the attraction argument in favor of the migration (or diffusion) argument, recognizing that the movement of one species is independent of the presence of the other.

Explain your observation of the conductivities of deionized water and tap water.

Expl. #2

Tap water conducts electricity because of the presence of ions, such as magnesium, calcium, chloride and nitrate. Deionized water does not conduct electricity because these ions are absent or at very low concentrations.

Explain your observations of the conductivities of solid sucrose and of sucrose solution.

Expl. #3

Neither the solid nor the aqueous solution conducts electricity because of the lack of ions in both the solid and in aqueous solution. Sucrose is a covalent compound which does not ionize after dissolving in water.

Explain your observations of the conductivities of solid sodium chloride and of sodium chloride solution.

Expl. #4

Solid NaCl does not conduct electricity even though it does contain Na^+ and Cl^- ionsbecause the ions in the crystal lattice are not mobile. The ions in a solid are held rigidly in place by electrostatic attractions. When water is added to the solid, the crystal lattice is broken down as the ions are hydrated by water molecules. In solution, the ions are mobile. As ions move toward the electrodes, the LED glows.

Explain the difference in the conductivities of the sodium chloride solution and sucrose solution.

Expl. #5

The sodium chloride solution conducts electricity because of the presence of ions (Na^+ and Cl^-) in the solution. Sucrose does not conduct because there are no ions in solution.

EXPERIMENT 6: CONDUCTIVITY

EQUIPMENT:
beaker, 50 mL.........................1
conductivity apparatus...............1
dropper................................1
paper towels..........................several
plate, 96-well.........................1

PART II: Strong and Weak Electrolytes

Place four drops of 0.1 M hydrochloric acid (HCl) in C3, four drops of 0.1 M acetic acid ($HC_2H_3O_2$) in C5, and four drops of 0.1 M sulfuric acid (H_2SO_4) in C7. Compare the conductivities of the three solutions by observing the intensity of the LED, for example, no glow (non conductor), faint (poor conductor) or brilliant (good conductor). Be sure to clean and dry the electrodes after each use. Classify each acid as either a nonelectrolyte, weak electrolyte or strong electrolyte. Identify the ions present in each solution which account for the conductivity.

Obs. #12

Solution Tested	LED Intensity	Electrolyte Strength	Ions Present
HCl	**bright**	**strong**	**H^+ and Cl^-**
faint	**weak**	**H^+ and $C_2H_3O_2^-$**	$HC_2H_3O_2$
H_2SO_4	**bright**	**strong**	**H^+ and HSO_4^-**

Place four drops of 0.1 M sodium hydroxide (NaOH) in C9 and four drops of 0.1 M ammonia (NH_3) in C11. Compare the conductivities of the two solutions by observing the intensity of the LED. Be sure to clean and dry the electrodes after each use. Classify each base solution as either a nonelectrolyte, weak electrolyte or strong electrolyte. Identify the ions that are present in each solution which would account for the conductivity.

Obs. #13

Solution Tested	LED Intensity	Electrolyte Strength	Ions Present
NaOH	**bright**	**strong**	**Na^+ and OH^-**
faint	**weak**	**NH_4^+ and OH^-**	NH_3

Place four drops of 0.1 M sodium acetate ($NaC_2H_3O_2$) in E1, four drops of 0.1 M sodium chloride (NaCl) in E3, four drops of 0.1 M ammonium acetate ($NH_4C_2H_3O_2$) in E5, and four drops of 0.1 M ammonium chloride (NH_4Cl) in E7. Use the conductivity apparatus to check each solution. Be sure to clean and dry the electrodes after each test. Record your results. Classify each solution as either a nonelectrolyte, weak electrolyte or strong electrolyte. Identify the ions that are present in each solution which would account for the conductivity.

Obs. #14

Solution Tested	LED Intensity	Electrolyte Strength	Ions Present
$NaC_2H_3O_2$	**bright**	**strong**	**Na^+ and $C_2H_3O_2^-$**
NaCl	**bright**	**strong**	**Na^+ and Cl^-**
$NH_4C_2H_3O_2$	**bright**	**strong**	**NH_4^+ and $C_2H_3O_2^-$**
NH_4Cl	**bright**	**strong**	**NH_4^+ and Cl^-**

Place four drops of methanol (CH_3OH) in E9 and four drops of ethanol (C_2H_5OH) in E11. Compare the conductivities of the two solutions by observing the intensities of the LED, for example, no glow (non conductor), faint (poor conductor) or brilliant (good conductors). Be sure to clean and dry the electrodes after each use. Classify each compound as either a nonelectrolyte, weak electrolyte or strong electrolyte. Identify the ions that are present in each solution which would account for the conductivity.

Obs. #15

Solution Tested	LED Intensity	Electrolyte Strength	Ions Present
CH_3OH	**none**	**none**	**no ions**
none	**none**	**no ions**	C_2H_5OH

Consider the following sample set of observations and the equation that results. Jennifer Redleg, an aspiring chemistry student, tested the conductivity of a 0.1 M nitric acid solution (HNO_3) and found that the LED glowed brightly. Jennifer concluded that the HNO_3 is a strong electrolyte. To demonstrate her knowledge of the ions formed in the solution Jennifer wrote the following equation:

$$HNO_{3}(aq) \rightarrow H^{+}(aq) + NO_{3}^{-}(aq) .$$

The equation can be interpreted the following way: An aqueous solution of HNO_3 contains the hydrated ions H^+ and NO_3^-. Jennifer realized that she could write HNO_{3}*(aq)* or H^+*(aq)* and NO_3^-*(aq)* to indicate an aqueous nitric acid solution.

Write a similar chemical equation which indicates the species that are in solution for each of the compounds whose conductivities were measured.

Reactions

$C_{12}H_{22}O_{11}(aq) \rightarrow$ ***DNI (does not ionize)***

$NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$

$KI(aq) \rightarrow K^+(aq) + I^-(aq)$

$Pb(NO_3)_2(aq) \rightarrow Pb^{2+}(aq) + 2NO_3^-(aq)$

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$

$HC_2H_3O_2(aq) \rightarrow H^+(aq) + C_2H_3O_2^-(aq)$

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

$NH_3(aq) \rightarrow NH_4^+(aq) + OH^-(aq)$

$NaC_2H_3O_2(aq) \rightarrow Na^+(aq) + C_2H_3O_2^-(aq)$

$NH_4C_2H_3O_2(aq) \rightarrow NH_4^+(aq) + C_2H_3O_2^-(aq)$

$NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

$CH_3OH(l) \rightarrow DNI$

$CH_3CH_2OH(l) \rightarrow DNI$

You have now collected data on a variety of substances. You have classified each solution as a nonelectrolyte, weak electrolyte or strong electrolyte. The terms strong electrolyte, weak electrolyte, or nonelectrolyte are used to summarize the experimental observations and refer to the ability of the compound to conduct electricity. In the case of the strong and weak electrolytes, you identified the ions in solution that were responsible for the observed ability to conduct electricity.

List the solutions that are strong electrolytes:

NaCl, KI, $Pb(NO_3)_2$, $NaC_2H_3O_2$, $NH_4C_2H_3O_2$, NH_4Cl
HCl, H_2SO_4
NaOH

List the solutions that are weak electrolytes:

$HC_2H_3O_2$
NH_3

List the solutions that are nonelectrolytes:

$C_{12}H_{22}O_{11}$, CH_3OH, C_2H_5OH

Based on your observations of the above 14 solutions, what classes of compounds are strong electrolytes when dissolved in water? Weak electrolytes? Nonelectrolytes? (Note: See the preparatory questions you answered at the beginning of the experiment .)

Expl. #6

Solutions containing covalent compounds

1) Organic compounds ($C_{12}H_{22}O_{11}$, CH_3OH, C_2H_5OH)-nonelectrolytes
2) Basic compounds (NH_3)-weak electrolytes
3) Acidic compounds ($HC_2H_3O_2$, HCl and H_2SO_4)-weak and strong electrolytes

Solutions containing ionic Compounds

1) Salts ($NaC_2H_3O_2$, NaCl, $NH_4C_2H_3O_2$, NH_4Cl)-strong electrolytes
2) Bases (NaOH)-strong electrolytes

Covalent bonded compounds can be nonelectrolytes, weak electrolytes or strong electrolytes. The covalent acids can be strong or weak electrolytes. The organic compounds are all nonelectrolytes. It may be necessary to accumulate more data to establish a better theory for the covalently bonded compounds.

All of the ionic compounds are strong electrolytes in solution.

EXPERIMENT 6: CONDUCTIVITY

EQUIPMENT:

laboratory balance	1	paper towels	3
beaker, 50 mL	2	plate, 96-well	1
capillary tube	1	ring stand	2
conductivity apparatus	1	spatula	1
dropper	1	stopper, 1-hole #6	1
flask, Erlenmeyer, 250 mL	1	trough or pan	1
glass bend	1	tubing,1' rubber	1
glass tubing	1	utility clamps	2
graduated cylinder, 50 mL	1	weighing paper	3

PART III: Electrolyte Strength and Reaction Rate

You will assemble an apparatus such as that shown in Figure 1. Start by filling the trough and the graduated cylinder with deionized water. Invert the graduated cylinder into the trough. After you have inverted the graduated cylinder be sure that it remains full of water. Fit the Erlenmeyer flask with a 1-holed rubber stopper. **Your instructor will demonstrate how to insert glass tubing into a rubber stopper. Follow these instructions carefully. Please ask if you have questions. The procedure is also demonstrated on the pre-lab video tape for Experiment #1.** Set up your apparatus as shown in Figure I. Complete your setup and be sure that all parts fit snugly so that no gas can escape. Have your instructor check your apparatus before you begin the experiment.

Apparatus Checked _________________

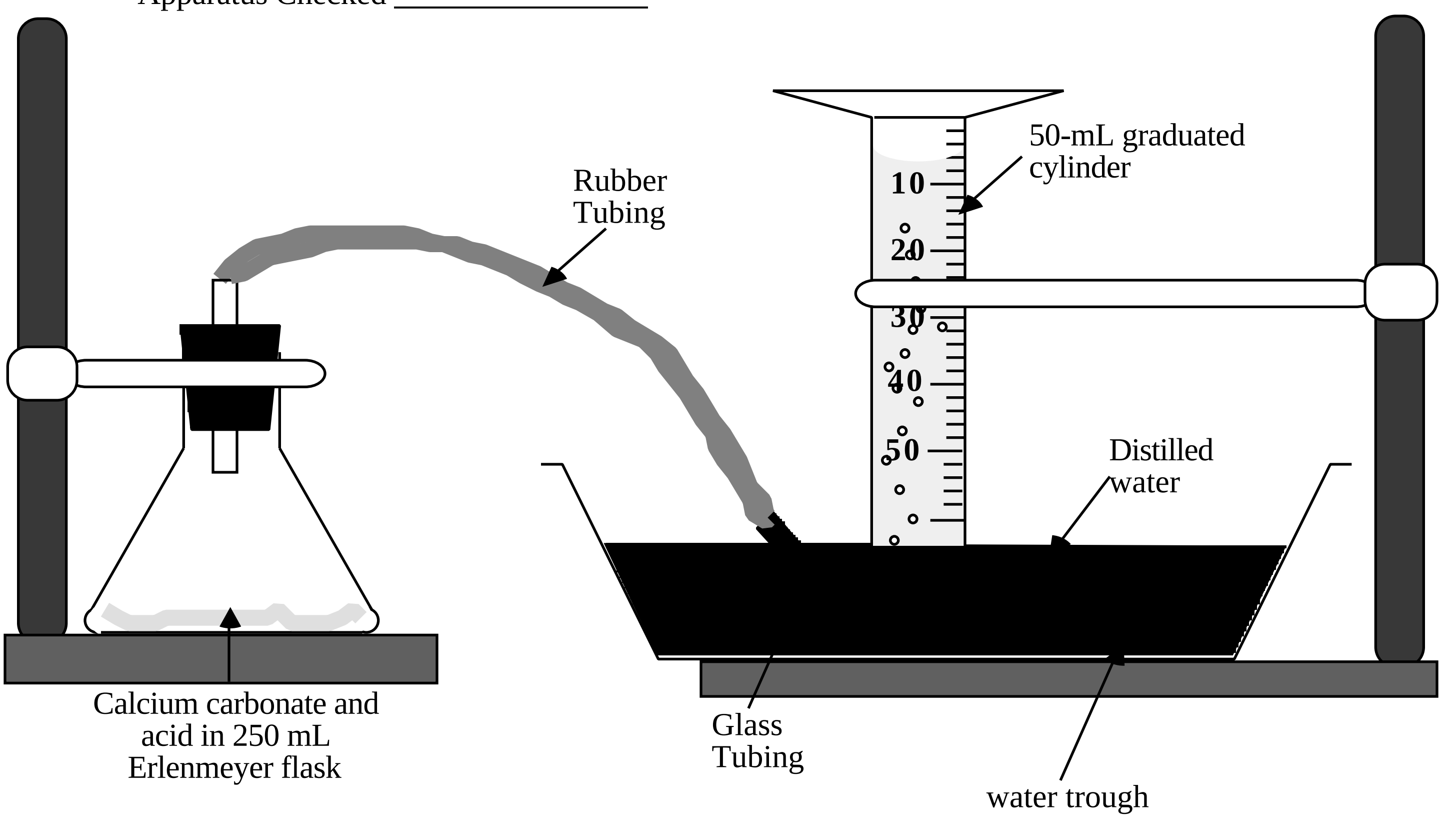


Figure I.

Measure approximately 2 grams of powerded calcium carbonate ($CaCO_3$) onto a small piece of paper. Be sure the sample texture and particle size is uniform. Obtain 30 mL of the 1 M HCl in a small beaker. After the instructor has checked your apparatus, and you are ready, add the acid to the Erlenmeyer flask. Then add the calcium carbonate to the acid in the Erlenmeyer flask, quickly stopper the flask and collect the escaping gas in the graduated cylinder. Note the time required to collect 20 mL of gas. The acid may react with $CaCO_3$ very rapidly, generating the 20 mL of gas quickly. You may wish to indicate the time as less than a second. Clean and rinse the flask and repeat the experiment for the other two acids (be sure to use 1 M $HC_2H_3O_2$ and 0.5 M H_2SO_4). It is not necessary to try to time these reactions. Simply compare and note the relative rate of evolution of gas. (Hint: Two of these should be about the same and the third noticeably different.)

	1 M HCl	1 M $HC_2H_3O_2$	0.5 M H_2SO_4
Rate of evolution of 20 mL of gas	**very fast**	**slow**	**very fast**

Record your observations for each acid.

Obs. #16

As soon as the solid calcium carbonate is added to the acid solutions a chemical reaction occurs producing a colorless gas. The speed of the reaction is very fast with HCl and H_2SO_4 and much slower with $HC_2H_3O_2$.

On the basis of your experimental data arrange the acids in the order of decreasing strength.

$$HCl = H_2SO_4 > HC_2H_3O_2$$

Write the reaction that occurs in the Erlenmeyer flask for each acid.

Reactions

(Students may need help with these equations.)

$2HCl(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + CO_2(g) + H_2O(l)$

$H_2SO_4(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + SO_4^{2-}(aq) + CO_2(g) + H_2O(l)$

$2HC_2H_3O_2(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + 2C_2H_3O_2^-(aq) + CO_2(g) + H_2O(l)$

How do you explain the difference in the time required to generate 20 mL of gas?

Expl. #7

The relative acid strength is directly related to formation of ions in solution. The students observed that both HCl and H_2SO_4 are strong electrolytes and $HC_2H_3O_2$ is a weak electrolyte. (Obs. #12). They have identified the ions formed in the solution. They should be able to conclude the reactivity is related to the amount of H^+ in solution. Both HCl and H_2SO_4 have high concentrations of H^+ (they are strong electrolytes) and the reaction rapidly produces CO_2. On the other hand $HC_2H_3O_2$ is a weak electrolyte and has a low concentration of H^+ ion and CO_2 is evolved slowly. Conclusion: When an acid is added to $CaCO_3$ the rate of evolution of CO_2 depends on the concentration of H^+.

Why was it important to have the three samples of $CaCO_3$ match in texture and particle size?

Expl. #8

If the sample texture and particle size are not similar, it would introduce another variable that would affect the rate of reaction of the acids with the $CaCO_3$. Large clumps of $CaCO_3$ have smaller surface area, and the reaction with the acids would be slower. The rate of CO_2 evolution can be more directly correlated with the strength of the acid if the three samples have similar textures and particle sizes.

Why was 0.5 M H_2SO_4 used instead of 1 M H_2SO_4?

Expl. #9

0.5 M H_2SO_4 and 1.0 M HCl have very nearly equal amounts (moles) of H^+ ion in equal volumes of solution. If concentrations in these ratios were not used there would not be equal concentrations of H^+ ion in the solutions and this would change the amount of CO_2 liberated.

The concentration of hydrogen ion, H^+ in 1.0 M $HC_2H_3O_2$ is 0.0042 M. How would the rate or reaction between 1.0 M $HC_2H_3O_2$ and $CaCO_3$ compare with the rate of reaction between 0.0042 M HCl and $CaCO_3$? Explain.

Expl. #10

The reaction rates should be equal, because the concentration of hydrogen ion in the two solutions is equal.

Place four drops of saturated (0.1 M) calcium hydroxide ($Ca(OH)_2$) in G7 and test its conductivity. Make sure the well is dry and free of contamination. Note: Be sure the sample of saturated calcium hydroxide you use is clear and colorless. If the sample is cloudy, check with the instructor.

Obs. #17
The solution is a good conductor.

Remove the electrodes of the conductivity apparatus from the solution of calcium hydroxide. Blow through a capillary tube into the solution and then measure its conductivity again. (Note: Be sure you are wearing your goggles). Keep repeating the procedure. Clean and dry the electrodes between measurements. Measure the conductivity of the solution and record your observations. Be sure to note any changes in the appearance of the solution.

Obs. #18

The solution becomes more and more cloudy when blowing through the capillary tube. The conductivity of the solution drops as more exhaled air, containing CO_2, is blown through the solution. When the well plate is allowed to stand undisturbed, a white precipitate falls to the bottom of the well.

Write a chemical reaction which describes what is occurring when you exhale into a saturated solution of calcium hydroxide.

Reaction

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

$Ca(OH)_2(aq) + H_2CO_3(aq) \rightarrow CaCO_3(s) + 2H_2O(l)$

$Ca^{2+}(aq) + 2OH^{-}(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

Explain your observations of the conductivity of the solution.

Expl. #11

An aqueous solution of $Ca(OH)_2$ contains Ca^{2+} and OH^{-} ions and is a good conductor. However, when blowing into the solution, CO_2 from exhaled gas precipitates $CaCO_3$ from the solution. The concentration of ions drops and the conductivity of the solution also drops.

EXPERIMENT 6: CONDUCTIVITY

EQUIPMENT:

beaker, 50 mL.........................1
conductivity apparatus................1
dropper................................1
paper towels......................several
plate, 96-well 2
spatula 1

PART IV: Chemical Reactions

Observe the color of 0.1 M solutions of the salts listed below and record the formula and color for both the cation and the anion. (If a solution is colorless, the ions it contains must also be colorless. The color of an ion is independent of the color of any other ions in the solution.)

Obs. #19

Salt Solution	Color of Solution	Cation and Color	Anion and Color
KCl	**colorless**	**K^+, colorless**	**Cl^-, colorless**
$NiCl_2$	**green**	**Ni^{2+}, green**	**Cl^-, colorless**
$CoCl_2$	**pink**	**Co^{2+}, pink**	**Cl^-, colorless**
$AgNO_3$	**colorless**	**Ag^+, colorless**	**NO_3^-, colorless**
$Ni(NO_3)_2$	**green**	**Ni^{2+}, green**	**NO_3^-, colorless**
KNO_3	**colorless**	**K^+, colorless**	**NO_3^-, colorless**

Place two drops of 0.1 M $AgNO_3$ with four drops of 0.1 M $NiCl_2$ in G9. Observe what happens, then test the conductivity of the solution.

Obs. #20

A white precipitate forms. The solution is green. The LED glows brightly indicating the presence of a strong electrolyte.

Identify the precipitate (if any) and the ions present in solution. Explain how you arrived at your conclusions.

Expl. #12

The solution mixture is light green therefore the solution contains Ni^{2+} ions. The precipitate must contain Ag^+, as it is colorless in solution. The anion precipitating with silver has to be Cl^-, not NO_3^- because we know that $AgNO_3$ is soluble.

Write an ionic equation for the reaction.

Reaction

$\mathbf{2Ag^{+}}$*(aq)* **+** $\mathbf{2NO_3^{-}}$*(aq)* **+** $\mathbf{Ni^{2+}}$*(aq)* **+** $\mathbf{2Cl^{-}}$*(aq)* → $\mathbf{2AgCl}$*(s)* **+** $\mathbf{Ni^{2+}}$*(aq)* **+** $\mathbf{2NO_3^{-}}$*(aq)*

Cite two methods you could use to test for ions in solution.

Expl. #13

1) **Test the conductivity of the solution. If the LED glows, ions must be present in the solution.**
2) **Observe the color of the solution. In this case the green color indicates** $\mathbf{Ni^{2+}}$ **ion is in the solution.**

Ions that remain unreacted in a solution are called spectator ions. A net ionic equation can be obtained by algebraically cancelling all the spectator ions. Write the net ionic equation for the above reaction.

Reaction

$\mathbf{Ag^{+}}$*(aq)* **+** $\mathbf{Cl^{-}}$*(aq)* → $\mathbf{AgCl}$*(s)*

Place two drops of 0.1 M $AgNO_3$ and four drops of 0.1 M $CoCl_2$ in G11. Observe what happens. Test the conductivity of the solution. (Be sure to clean the electrodes of the conductivity device.)

Obs. #21

A white precipitate forms. The solution is pink.

The LED glows brightly.

Write the ionic and net ionic equations for the reaction.

Reaction

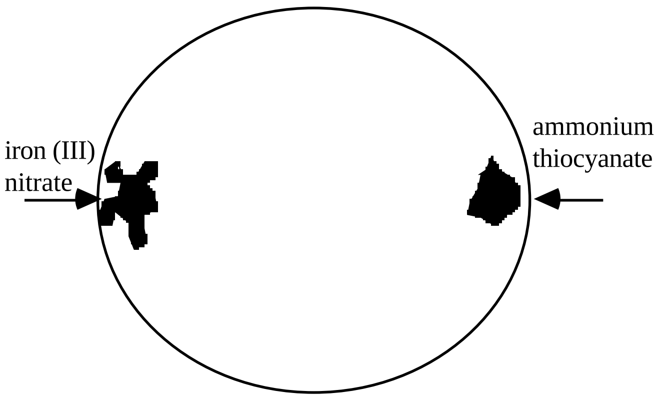
$\mathbf{2Ag^{+}}$*(aq)* **+** $\mathbf{2NO_3^{-}}$*(aq)* **+** $\mathbf{Co^{2+}}$*(aq)* **+** $\mathbf{2Cl^{-}}$*(aq)* → $\mathbf{2AgCl}$*(s)* **+** $\mathbf{Co^{2+}}$*(aq)* **+** $\mathbf{2NO_3^{-}}$*(aq)*

$\mathbf{Ag^{+}}$*(aq)* **+** $\mathbf{Cl^{-}}$*(aq)* → $\mathbf{AgCl}$*(s)*

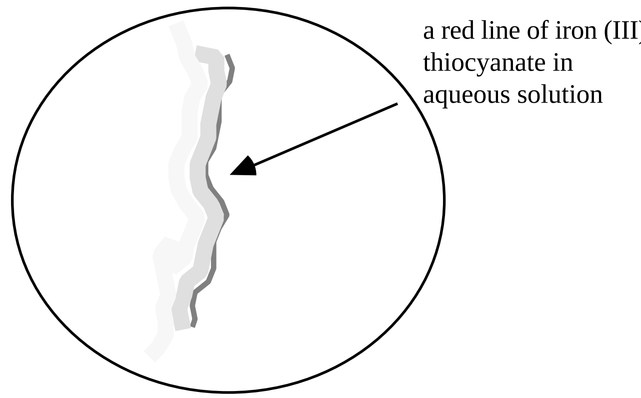
Add a crystal of iron(III) nitrate, $Fe(NO_3)_3$, to one side of a Petri dish containing deionized water and then add a few crystals of ammonium thiocyanate, NH_4SCN, to the other side. Describe what happens. Include a picture.

Obs. #22

The reaction which is occurring in the solution is commonly represented as, $Fe^{3+} + SCN^- \rightarrow FeSCN^{2+}$.



Initially



Finally

Explain what must be happening in the solution to account for your observations.

Expl. #14

Both iron(III) nitrate and ammonium thiocyanate are strong electrolytes. The appearance of the red line in solution indicates that there has been a chemical reaction. Since the solids dissolve and then ionize the ions in the solution mix and the reaction is probably between one of the cations and one of the anions. It is unlikely that the reaction is as a result of the combination of NH_4^+ and the NO_3^-, so it must be a reaction between Fe^{3+} and SCN^-.

Write the ionic and net ionic equations for the reaction.

Reaction

Note: Students will need help writing the net ionic equation.

$Fe(NO_3)_3(aq) + NH_4SCN(aq) \rightarrow FeSCN(NO_3)_2(aq) + NH_4NO_3(aq)$

$Fe^{3+}(aq) + 3NO_3^-(aq) + NH_4^+(aq) + SCN^-(aq) \rightarrow FeSCN^{2+}(aq) + 3NO_3^-(aq) + NH_4^+(aq)$

$Fe^{3+}(aq) + SCN^-(aq) \rightarrow FeSCN^{2+}(aq)$

Post-lab Questions:

The answers to the following problems must accompany your laboratory report.

1. Write the ionic and net ionic equations for each of the following combinations. Identify the color of any precipitate and the supernatant solution in each case.

   (a) HNO_3*(aq)* and NaOH*(aq)*

$H^+(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + H_2O(l)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

colorless

   (b) KNO_3*(aq)* and $NiCl_2$*(aq)*

$K^+(aq) + NO_3^-(aq) + Ni^{2+}(aq) + 2Cl^-(aq) \rightarrow K^+(aq) + NO_3^-(aq) + Ni^{2+}(aq) + 2Cl^-(aq)$

no reaction **green solution**

   (c) $AgNO_3$*(aq)* and KCl*(aq)*

$Ag^+(aq) + NO_3^-(aq) + K^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + NO_3^-(aq) + K^+(aq)$

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

white solid, colorless solution

   (d) $Ni(NO_3)_2$*(aq)* and $AgNO_3$*(aq)*

$Ni^{2+}(aq) + 2NO_3^-(aq) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow Ni^{2+}(aq) + 2Ag^+(aq) + 4NO_3^-(aq)$

no reaction **green solution**

   (e) H_2SO_4*(aq)* and $CaCO_3$*(aq)*

$2H^+(aq) + 2HSO_4^-(aq) + Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow Ca^{2+}(aq) + 2HSO_4^-(aq) + H_2O(l) + CO_2(g)$

$2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

colorless solution, colorless gas

   (f) HCl*(aq)* and Na_2CO_3*(s)*

$2H^+(aq) + 2Cl^-(aq) + Na_2CO_3(s) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + H_2O(l) + CO_2(g)$

$2H^+(aq) + Na_2CO_3(s) \rightarrow 2Na^+(aq) + H_2O(l) + CO_2(g)$

2. Why is it necessary to use deionized water when testing the conductivity of aqueous solutions?

   Any measured conductivity should be due to the solute particles and not the solvent particles.

3. Aqueous ammonia, NH_3*(aq)*, and acetic acid, $HC_2H_3O_2$*(aq)*, solutions of equal concentrations, conduct electric current equally well. Explain why the addition of one solution to the other results in a substantial increase in electrical conductivity.

   $\mathbf{NH_3}(aq) + \mathbf{HC_2H_3O_2}(aq) \rightarrow \mathbf{C_2H_3O_2^-}(aq) + \mathbf{NH_4^+}(aq)$

   The reaction is a neutralization reaction between a weak base and a weak acid. The product is a salt ($NH_4C_2H_3O_2$) that is a strong electrolyte, therefore an increase in the conductivity of the solution is observed.

4. Ammonium sulfate and barium hydroxide solutions are each very good conductors. However, when equal volumes of solutions of equal concentrations are mixed, a dramatic decrease in conductivity is observed. Explain.

   $\mathbf{NH_4SO_4}(aq) + \mathbf{Ba(OH)_2}(aq) \rightarrow \mathbf{BaSO_4}(s) + \mathbf{H_2O}(l) + \mathbf{NH_3}(g)$

   In the reaction above there is a neutralization as well as a precipitation reaction. Both of the reactants are strong electrolytes, but the products formed are a solid, a liquid and a weak electrolyte (NH_3). So as the reaction proceeds one observes the conductivity decreasing. At the equivalence point the conductivity of the solution is low because the only ions present are NH_4^+ and OH^-.

5. Making predictions based on previous experimental evidence is an important goal for a chemist. Use your classification system, (**Obs. #6**), to predict the conductivity of each of the following solutions. Predict whether the substance is a strong electrolyte, a weak electrolyte, or a nonelectrolyte.

   (a) $HClO_4$ **strong electrolyte**

   (b) $Ca(NO_3)_2$ **strong electrolyte**

   (c) NH_2CONH_2 (urea) **nonelectrolyte**

   (d) HBr **strong electrolyte**

   (e) H_3PO_4 **weak electrolyte**

   (f) $(NH_4)_2CO_3$ **strong electrolyte**

   (g) $PbCl_2$ **strong electrolyte**

   (h) KOH **strong electrolyte**

   (i) $C_3H_5(OH)_3$ (glycerol) **nonelectrolyte**

   (j) PbI_2 **strong electrolyte**

   (k) $CH_3CH_2CH_2OH$ **nonelectrolyte**

   Note: Even though both PbI_2 and $PbCl_2$ are insoluble in water, both salts are sparingly soluble. The amount of both salts that <u>does</u> dissolve, completely dissociates.

INSTRUCTOR EVALUATION NAME: ____________________
EXPERIMENT 6: CONDUCTIVITY SCHOOL: ____________________

Please complete the form as soon as possible after your students have completed the laboratory. Include any comments you have on each section of the experiment. If the answer to any question is "no" please note the specific problems or difficulties encountered. Attach extra sheets if necessary. At the end of the semester, return all forms to **Dr. John Gelder, Department of Chemistry, Oklahoma State University, Stillwater, OK 74078**. Your comments and suggestions are very important in helping to correct errors and improve the overall quality of this manual.

1. How much time was required to complete the experiment? _______ hours
   Briefly describe those sections of the experiment which were completed during each laboratory period. (Note: You may include Part numbers or page numers for simplicity.)

	NO	YES
2. Was the pre-lab exercise ...		
A. ...completed by the students?	_____	_____
B. ... adequate introduction to the ideas introduced in the experiment?	_____	_____

Comments:

	NO	YES
3. Were the laboratory instructions ...		
A. ...understood by the students with little or no assistance from you?	_____	_____
B. ...leading to the collection of necessary data ?	_____	_____
C. ...resulting in data with acceptable experimental error?	_____	_____

Comments:

	NO	YES
4. Were the questions and calculations included in the experiment ...		
A. ...completed by most students?	_____	_____
B. ...relevant to the experiment?	_____	_____

Comments:

	NO	YES
5. Were the post-lab problems ...		
A. ...completed by most students?	_____	_____
B. ...relevant to the experiment?	_____	_____
C. ...sufficient to illustrate the overall goals of the experiment?	_____	_____

Comments:

	NO	YES
6. Was the experiment as a whole ...		
A. ...interesting to the students?	_____	_____
B. ...relevant to the course work?	_____	_____
C. ...written at an appropriate level of difficulty?	_____	_____

Comments: