

Solution Conductivities and Titrations

Minneapolis Community and Technical College

v.9.17

Objective: Introduce you to aqueous solutions, ion concentrations, conductivity measurements and acid-base titrations and techniques.

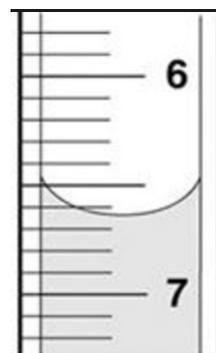
Prelab Questions: Read through this lab handout and answer the following questions before coming to lab. There will be a quiz at the beginning of lab over this handout and its contents.

1. List the ions formed when each of the following is dissolved in water: CuCl_2 , Na_2SO_4 , $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, KBr , $\text{C}_6\text{H}_6\text{O}_2$,
2. Consider a 2.30 Molar solution of FeCl_3 . What is the concentration of each ion? What is the TIC for this solution?
3. A student measures a solution's conductivity and observes the following fluctuating readings on the computer screen:

12345 μS , 12490 μS , 12302 μS , 12289 μS , 12388 μS , 12201 μS , 12475 μS .

What should be reported as the solution's conductivity?

4. How is the solution stirred when performing the titration experiment?
5. Which of the following solutions should have the greatest conductivity?
1.00 M KCl , 0.75 M Na_2SO_4 , 0.80 M CaCl_2
6. What is the correct reading for the burette measurement at right?
7. Why does aqueous acetic acid conduct electricity better than pure acetic acid liquid?



Introduction:

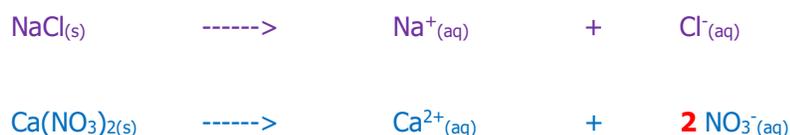
Ionic solids don't conduct electricity until mixed with water to form an aqueous solution. Once dissolved, the ions are free to move and conduct electricity. If a molecular solid like sugar is dissolved in water, no additional ions are produced and the solution conducts electricity no better than the pure water did.

We use the following three terms to describe how well liquids and solutions conduct electricity:

- **Strong electrolyte**
 - Conducts electricity well
 - Solution contains a lot of cations and anions
 - Example: Table salt (NaCl) dissolved in water
- **Weak electrolyte**
 - Conducts electricity moderately
 - Solution contains a modest number of ions
 - Example: Acetic acid (CH_3COOH) dissolved in water
- **Non-electrolyte**
 - Conducts very little electricity
 - Solution contains few or no ions.
 - Example: Sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) dissolved in water

A solution conducts electricity in direct proportion to the number of ions present. For example, consider these two solutions: **1.5 M Ca(NO₃)₂** and **1.5 M NaCl**. We will determine which of the two solutions contains more ions and is therefore the stronger electrolyte.

Being ionic substances (metal/non-metal combinations) each salt dissociates in water as follows. Note that NO₃⁻ is a polyatomic anion that doesn't break apart.



Using the provided 1.5 M concentrations, we determine the concentrations of individual ions and the **Total Ion Concentration (TIC)**. Consider the 1.5 M NaCl. Because 1 sodium ion is produced when the salt dissociates, we determine that the Na⁺ concentration is 1.5 M. The same reasoning applies to the chloride ion and its concentration is also 1.5 M.

However, the conductivity of the solution depends on the contributions of both the sodium and chloride ions combined. Therefore, we calculate the TIC by adding the individual ion concentrations together:

$$\begin{array}{l} \text{TIC}_{\text{NaCl}} \quad = \quad \text{Concentration}_{\text{Na}^+} \quad + \quad \text{Concentration}_{\text{Cl}^-} \\ \text{TIC}_{\text{NaCl}} \quad = \quad 1.5\text{M}_{\text{Na}^+} \quad + \quad 1.5\text{M}_{\text{Cl}^-} \\ \\ \mathbf{\text{TIC}_{\text{NaCl}} \quad = \quad 3.0 \text{ M}} \end{array}$$

Having finished with NaCl, we now turn to the 1.5 M Ca(NO₃)₂ solution. Like NaCl above, only one Ca²⁺ ion is produced when the solid salt dissociates. Therefore, the concentration of the Ca²⁺ ions is 1.5 M. However, **two** NO₃⁻ ions are produced making their concentration **2 X (1.5M) = 3.0 M**. Calculating the total ion concentration we have:

$$\begin{array}{l} \text{TIC}_{\text{Ca(NO}_3)_2} \quad = \quad \text{Concentration}_{\text{Ca}^{2+}} \quad + \quad \text{Concentration}_{\text{NO}_3^-} \\ \text{TIC}_{\text{Ca(NO}_3)_2} \quad = \quad 1.5\text{M}_{\text{Ca}^{2+}} \quad + \quad \mathbf{2 \text{ X } (1.5\text{M})}_{\text{NO}_3^-} \\ \\ \mathbf{\text{TIC}_{\text{Ca(NO}_3)_2} \quad = \quad 4.5\text{M}} \end{array}$$

Because the Ca(NO₃)₂ solution has the larger TIC value, it contains more ions and will be more conductive. In fact, we can use TIC values to predict relative conductivities mathematically. Let's say that the NaCl solution is measured and found to have conductivity of 15000 μS (micro Seimans is a unit of conductivity). The Ca(NO₃)₂ conductivity can be predicted using ratios:

$$\frac{\text{TIC}_{\text{Ca(NO}_3)_2}}{\text{TIC}_{\text{NaCl}}} = \frac{\text{Conductivity}_{\text{Ca(NO}_3)_2}}{\text{Conductivity}_{\text{NaCl}}}$$

Using the TIC values derived above for each solution we have:

$$\frac{4.5 \text{ M}}{3.0 \text{ M}} = \frac{\text{Conductivity}_{\text{Ca(NO}_3)_2}}{15000 \text{ } \mu\text{S}}$$

Solving for the Ca(NO₃)₂ conductivity we get **22,500 μS**. This is a 50% increase in conductivity in comparison to the 1.5 M NaCl solution even though both solutions were labeled as 1.5 M.

Molecular liquids, like glacial * acetic acid (CH₃COOH_(l)) don't conduct electricity well and are considered non-electrolytes. This is because the covalent bonds that hold the molecules together are strong and resist breaking. Few, if any ions are produced and the liquid doesn't conduct electricity.

* Glacial acetic acid is a pure substance and gets its name from the fact that when frozen, it forms long, beautiful crystals that resemble glacial ice.

However, when mixed with water, the acetic acid *solution* is more conductive than either the pure water or glacial acetic acid alone. This fact is due to water's attraction to the acetic acid molecule via a force known as hydrogen bonding. Several water molecules, working together pull and break the bond connecting the **hydrogen atom** to the rest of the acetic acid molecule. This process is illustrated by the following chemical equation:



Note that there are **ionic products** and these are responsible for the solution's moderate conductivity.

We refer to the "left to right" conversion of reactants to products in a reaction like this as the "forward reaction." However, once the forward reaction has created enough ions, it becomes likely that they will bump into one another and convert back into non-conducting reactants; something known as the "reverse reaction."

Both the forward and reverse reactions are indicated in the chemical equation above by the **double arrow**.

At some point, the levels of non-conductive reactants and conductive products stabilize reaching a state known as dynamic equilibrium. Conductivity measurements reflect the equilibrium levels of product ions in the solution.

Experiment: Calibration of the Conductivity Probe

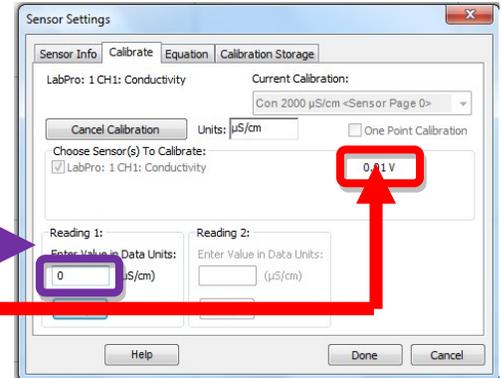
1. Log into a laboratory computer and connect it to the LabPro device.
2. Position the conductivity probe sensitivity switch on the 20,000 setting. This setting will allow the greatest range of conductivity measurements.
3. Plug the conductivity probe into Channel#1
4. Open the LoggerPro application:
The display gives solution conductivity in micro-Siemens (μS).
5. Calibrate your conductivity probe using two solutions of known conductivity
Liquid 1: Distilled water (Conductivity = 0) and NaCl (Conductivity = $1000\mu\text{S}$)



Calibration Procedure

The screen display is locked during the calibration process.

1. Click Experiment → Calibrate → Conductivity Probe
2. Click "Calibrate Now" in the sensor settings window
3. Immerse the conductivity probe in the distilled water sample
4. Stir using an up and down motion with the conductivity probe
5. Enter "0" in the text box
6. Click the "Keep" button when the VOLTAGE reading stabilizes
7. Rise off the probe with distilled water and blot dry with a tissue
8. Immerse the probe in the $1000\mu\text{S}$ calibration solution.
9. Stir using an up and down motion with the conductivity probe (don't splash)
10. Click "Keep" when the voltage reading stabilizes
11. Click "Done"
12. Re-measure the conductivities of the two standard solutions to be sure the conductivity probe reads correctly.



Experiment: Solution Conductivities

1. Inspect the solution tray provided. You should have two water samples, three ionic solutions and four molecular samples.
2. Remove the cap from one vial and insert the conductivity probe directly into the solution. Be sure the conductivity probe's lower electrode hole is totally immersed.
3. Gently agitate the solution with an up/down motion of the conductivity probe. Be careful not to splash liquids out of vials.
4. Record the conductivity on your data sheet while stirring (up/down). Not all digits are significant and it is your responsibility to decide what digits should be recorded.
5. Replace cap on vial when finished.
6. Rinse the conductivity probe over the waste beaker provided with each tray using a wash bottle and *small amounts* of distilled water. One beaker is provided for the water/salt solution samples and another beaker for the molecular samples.
7. Gently shake excess water from the probe and dry it by gently blotting the outside of the electrode assembly with a tissue. It isn't necessary to dry the interior (hole) in the electrode assembly.
8. Use the above procedure to test all remaining solutions. (Be sure to collect rinse solutions correctly)

Experiment: Burette Operation

The burette is a device that is used to dispense a liquid, the **titrant**, into a flask or beaker containing a different solution, the **titer**. Before use, the burette must be rinsed with the titrant solution to remove contamination that may be the result of prior use and/or dishwasher residue.

Note also the position of the burette's valve handle. When perpendicular to the body of the burette, the valve is off. The burets in the figure at right are turned off.

Rinse the burette by *first closing the valve* and then adding approximately 2-4 mL of titrant solution (record the bottle number). Now tilt the burette (almost horizontal) and allow the titrant solution to run in the direction of the open end. Just as the solution approaches the open end, *spin the burette between your fingers* but don't let the solution run out of the burette's open end onto your fingers!! Discard the rinse solution as instructed. Now rinse the burette a second time.

Fill the burette using a small plastic funnel (valve off). Run out enough liquid to remove any bubbles that may be trapped in the tip. Record the initial reading to 2 decimal digits (see figure at right).

Note that burets are read from the top down UNLIKE graduated cylinders that read from the bottom up. Don't make the mistake of reading the first buret (figure at above) as 1.73 mL. *The correct reading should be 0.27 +/- 0.01 mL.*

Slowly add titrant solution until you near the endpoint, signaled by a faint color change (pink for phenolphthalein) that lasts for progressively longer times.

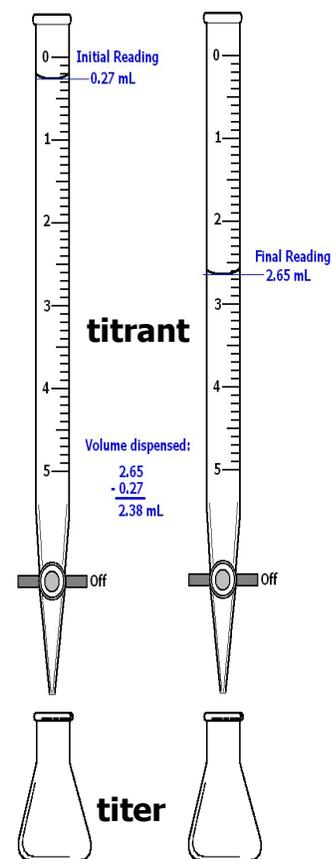
Watch for bubbles that may be hiding in the valve assembly. If you see one, it must NOT be allowed to flow out the burette's tip. If the bubble comes out, the titration will be in error and must be performed again. You may be able to continue by *slowing* down the titration.

Look for a color change early on as the titrant solution first encounters the solution in the beaker or flask below. Failure to see a color change means you forgot to add the indicator!

Note: During a titration, the meniscus must always fall on the burette's graduated scale! Don't let the meniscus drop below the lowest mark or you will have to repeat the trial!

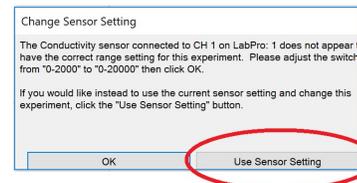
Two methods may be used to add the last, small amounts of titrant near the endpoint.

1. Quickly turn the valve 180 degrees from one off position to the next. When performed quickly enough only a small amount will be added to the beaker. PRACTICE THIS TECHNIQUE PRIOR TO DETERMINING YOUR ENDOINT.
2. Slowly open the valve and allow a small partial droplet to form at the end of the burette. Rinse this droplet from the tip of the burette using a small amount of water into the beaker below. The small amount of water should have no effect on your result.



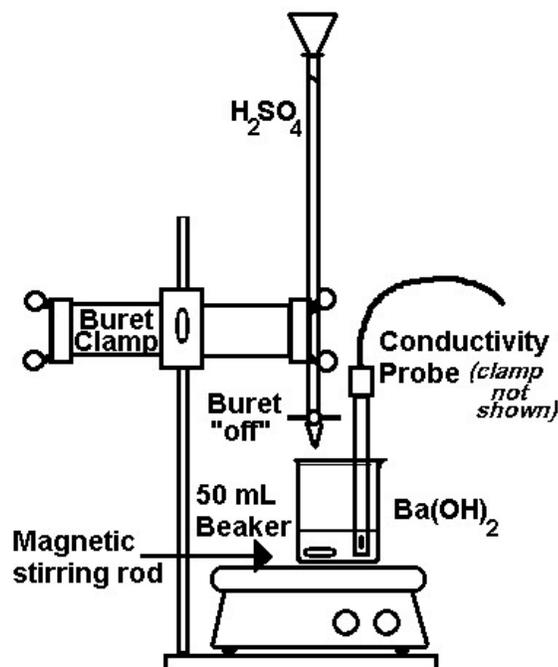
Experiment: Conductivity and an acid/base titration

1. Obtain a $\text{Ba}(\text{OH})_2$ solution from the instructor and record its number on your data sheet.
2. On the conductivity control box, move the toggle switch to the 0-2000 μS position. We will not be recalibrating the conductivity probe for this experiment.
3. When asked, click on the "Use Sensor Setting" button that appears on screen (Figure at right)
4. Obtain approximately 35 mL of H_2SO_4 and write the concentration of the acid on your data sheet.
5. Rinse the burette twice with 1-2 mL portions of the sulfuric acid. Remove the funnel before starting your titration.



6. Use the small funnel to fill the burette with H_2SO_4 and eliminate any bubbles from the tip of the burette.
7. Adjust the liquid level in the burette to read 0.00 mL.
8. Use a graduated cylinder to measure out 20.00 mL of the $\text{Ba}(\text{OH})_2$ solution and pour it into a clean 50 mL beaker.
9. Place a small magnetic stirring rod in the beaker and position it on the stir plate.
10. Adjust the speed control on the stirrer to mix well but without splashing.
11. Use the articulated probe holder (not shown in figure) to support the conductivity probe and position it in the 50 mL beaker.

Be sure the electrode hole at the bottom of the probe is completely immersed in solution.



12. Add 4 drops of thymol blue indicator to the beaker
13. Position the buret containing the H_2SO_4 above the 50 mL beaker.
14. Record the initial burette reading (0.00 mL) and conductivity on your datasheet.
15. Add a small amount (~ 0.5 mL or less) of sulfuric acid to the beaker and wait several seconds for the reading to stabilize. Record the buret reading and solution's conductivity on your data sheet.
16. Continue adding small amounts of sulfuric acid. With each addition, record the buret reading and solution conductivity.
17. As you get closer to the equivalence point (minimum in conductivity), add progressively smaller amounts of sulfuric acid. The quality of your titration will be graded according to the number of data points you obtain in the near zero conductivity region of the graph.
18. After reaching the equivalence, go beyond it in 1 mL increments for a total of 4 mL. **THIS STEP IS IMPORTANT!!!** Be sure you have at least 4 more data points beyond the conductivity minimum.
19. Dispose of your waste solutions as instructed (not down the drain).

Answers to pre-lab questions:

1. CuCl_2 : Cu^{2+} Cl^- Cl^- Na_2SO_4 : Na^+ Na^+ SO_4^{2-} $\text{C}_{12}\text{H}_{22}\text{O}_{11}$: molecular ... no ions
KBr : K^+ Br^- $\text{C}_6\text{H}_6\text{O}_2$: molecular ... no ions

2. Fe^{3+} concentration = 2.30 M Cl^- concentration = 6.90 M TIC = 9.20 M

3. **"1" and "2"** don't fluctuate and are "certain digits." The third digit fluctuates but averages a **"3"**.
Reported value **12300** μS where **zeros** are used as placeholders and not significant.
Represents a range of values from 12200 ... 12300 ... 12400 μS

4. A small magnetic stir bar is placed in the 50 mL beaker. The hotplate/stir-plate spins the stir-bar and stirs the solution. The solution is stirred briskly.

5. Which of the following solutions should have the greatest conductivity?

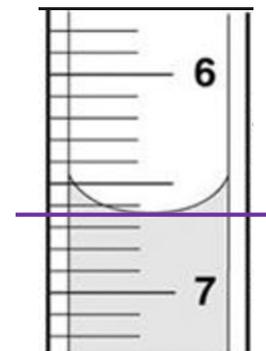
1.0 M KCl TIC = 2.00 M

0.75 M Na_2SO_4 TIC = 2.25 M

0.80 M CaCl_2 TIC = 2.40 M (Most concentrated, most conductive)

6. 6.64 +/- 0.01 mL

7. Water molecules interact with the acetic acid molecules pulling them apart to form ions.



CHEM1151 Data Sheet

Solution Conductivity

Instructor Initials_____

Name_____

Date of Exp. _____

Lab Section _____

*Your individual experimental report will be due at the beginning of class next week.**It should include 1. Signed data page 2. Answers to questions 3. Conductivity vs. mL titration graph**Data Tables: (All entries must be in written in ink before you leave the lab).***Solution Conductivities**

Description	Liquid	Conductivity (μS)
Liquid water	H ₂ O (Tap)	
Liquid water	H ₂ O (distilled)	
Aqueous Salt Solution	0.05 M NaCl (aq)	
Aqueous Salt Solution	0.05 M CaCl ₂ (aq)	
Aqueous Salt Solution	0.05 M AlCl ₃ (aq)	
Liquid Methanol	CH ₃ OH(aq)	
Aqueous Methanol Solution	0.05 M CH ₃ OH(aq)	
Liquid Glacial Acetic Acid liquid	CH ₃ COOH (l)	
Aqueous Acetic Acid Solution	0.05 M CH ₃ COOH (aq)	

Titration of an acid and base**Unknown # _____ H₂SO₄ concentration (bottle) _____ M**

	Buret Reading (mL)	Conductivity (μS)		Buret Reading (mL)	Conductivity (μS)		Buret Reading (mL)	Conductivity (μS)
1	0.00 mL		11			21		
2			12			22		
3			13			23		
4			14			24		
5			15			25		
6			16			26		
7			17			27		
8			18			28		
9			19			29		
10			20			30		

Questions: Your answers in your own words for credit.

1. Which water sample has the greatest ion concentration? Where does the more conductive sample get its ions?

2. All three salt solutions have the same 0.05 M concentration. What is the TIC for each solution?

_____ AlCl_3

_____ CaCl_2

_____ NaCl

3. Use your conductivity measurements to calculate how many times more conductive the AlCl_3 solution is compared to the NaCl solution (Show work below).

How does your value compare to what you would expect from the TIC values?

4. Compare the four acetic acid and methanol conductivity measurements. What do they suggest about the bond that attaches the hydrogen atom to the methanol molecule compared to acetic acid?

5. Write the balanced acid/base chemical reaction that occurs between barium hydroxide and sulfuric acid.

6. Referring to #5, how do the conductivities of the reactant species compare to the conductivities of the product species?

7. For the first half of the titration, conductivity and ion concentration decrease as H_2SO_4 is added. Why?

8. For the second half of the titration, conductivity and ion concentration increase as H_2SO_4 is added. Why?

Graph: Use your titration data and Excel to construct a conductivity vs. Volume_{H₂SO₄} graph.

Your graph should include the following:

1. Graph title (including your name and the experimental date)
2. Axis titles and units
3. Text box labels properly positioned for:
 - a. "Increasing Ion Concentration"
 - b. "Decreasing Ion Concentration"
4. A vertical line centered accurately at the conductivity's lowest point. Label the line "Equivalence Point" and include the H₂SO₄ volume from the horizontal axis with one decimal digit.
5. The following information in a text box:
 - a. H₂SO₄ Volume at equivalence point (1 decimal digit).
 - b. Moles of H₂SO₄ at equivalence point.
 - c. Moles of Ba(OH)₂ at equivalence point.
 - d. Molar concentration of Ba(OH)₂

***Note that the graph (right) is incomplete. ***
***Your graph should include all listed information ***

