EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

PURPOSE:

1. To estimate the Acid-Ionization Constant (Ka) for acetic acid by conductivity testing comparisons.
2. To become familiar with the pH meter and pH measurements.
3. To determine the Acid-Ionization Constant (Ka) for acetic acid by pH measurements.
4. To examine the effect of dilution on the degree of ionization of acetic acid.

PRINCIPLES:

All acids (strong and weak) react with water to produce hydrogen ions. This process is referred to as **Acid Ionization** or **Acid Dissociation**.

Since all strong acids ionize completely in solution (100%), for solutions of strong acids, the concentrations of ions strong acids produce in solution are determined by the stoichiometry of the reaction and the initial concentration of the acid.

	$HCl(aq)$	+ $H_2O(l)$	$\longrightarrow$ $H_3O^+(aq)$	+ $Cl^-(aq)$
Initial Concentration	0.1 M	excess	0	0
Final Concentration	0	excess	0.1 M	0.1 M

The situation is quite different for weak acids. Consider an aqueous solution of the weak acid, acetic acid ($HC_2H_3O_2$). When acetic acid is added to water, acetic acid ionizes only partially, and an equilibrium system is established, in which the molecules of acetic acid $HC_2H_3O_2$ are in equilibrium with the ions they produce by ionization (H^+ and $C_2H_3O_2^-$)

This equilibrium system, commonly referred to, as Acid Ionization Equilibrium, can be represented by the following equation:

	$HC_2H_3O_2(aq)$	+ $H_2O(l)$ $\rightleftarrows$	$H_3O^+(aq)$	+ $C_2H_3O_2^-(aq)$
Initial Concentration	0.1000 M	excess	0	0
Change	- 0.0013	excess	+ 0.0013	+ 0.0013
Equilibrium Concentration	0.0987	excess	0.0013 M	0.0013 M

As shown above, this equilibrium system strongly favors the reactants, as evidenced by the fact that in a solution of 0.1000 M of acetic acid, only 1.3 % of the acetic acid molecules change into ions.

$$\%\ \text{Ionization} = \frac{\text{Number of molecules ionized}}{\text{Total number of molecules originally present}} = \frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]} \times 100 = \frac{0.0013\ \text{M}}{0.1000\ \text{M}} \times 100$$

% Ionization = 1.3 %

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

At the level of arithmetic, this implies that out of 100 molecules of $HC_2H_3O_2$, only one molecule breaks apart to produce one hydronium ion (H_3O^+) and one acetate ion ($C_2H_3O_2^-$)
Since acidity is associated with the concentration of hydronium ions [H_3O^+] in the solution of an acid, the relative strength of various weak acids is judged by the concentration of the H_3O^+ they produce at a given concentration. The higher the extent of ionization of the weak acid, the more [H_3O^+] are produced and the relatively stronger the weak acid is.
The extent to which weak acids ionize differ greatly among weak acids. In order to objectively judge the relative strength of weak acids, a constant Ka, Acid-Ionization Constant has been introduced.
Ka is in fact the equilibrium constant for the ionization of a weak acid and it has a specific numerical value for every weak acid.
To find Ka, consider the Acid Ionization Equilibrium of a weak acid:

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

The corresponding equilibrium constant is:

$$Kc = \frac{[H_3O^+]\,[A^-]}{[HA]\,[H_2O]}$$

Assuming that this is a dilute solution and that the forward reaction occurs to only a small extent (as it is always the case for weak acids), the concentration of water can be considered constant.
Rearranging the equation, with the aim to include the concentration of H_2O in the equilibrium constant, Kc, yields:

$$Kc \times [H_2O] = \frac{[H_3O^+]\,[A^-]}{[HA]} = \textbf{Ka} = \textbf{Acid – Ionization Constant}.$$

This expression indicates that for weak acids that are relatively highly ionized – that is few molecules of HA in solution – the value of Ka will be relatively large. A relatively small value of Ka indicates that the weak acid is only slightly ionized.

Ka of a weak acid can be determined experimentally by one of two methods:

1. **<u>By conductivity testing comparison</u>**.
   This method is based on the premise that two solutions that exhibit the same electrical conductivity also have the same concentration of ions. An estimate of the number of ions produced by the ionization of a weak acid ($HC_2H_3O_2$) of known concentration is determined by comparing the conductivity of a solution of the weak acid of a known concentration with the conductivity of a solution of a strong acid (HCl) of a known concentration.
   Since the conductivity indicator used in this experiment provides only semi-quantitative data, this method often yields a very approximate estimate of the Acid-Ionization Constant, Ka.

EXPERIMENT 6

A STUDY OF AN ACID IONIZATION EQUILIBRIUM

Assume that the conductivity indicator used in this experiment indicates that a 0.10 M solution of a weak acid, HA, has the same conductivity as a 0.0040 M solution of the strong acid, HCl. This implies that that these two solutions contain the same concentration of ions.

The Acid – Ionization Constant, Ka, can be calculated as follows:

$$Ka = \frac{[H^+]\,[A^-]}{[HA]} = \frac{[H^+]\,[Cl^-]}{[HA]} = \frac{(0.0040)\,(0.0040)}{(0.10)} = 1.6 \times 10^{-4}$$

2. **<u>By pH measurements</u>**

   In this method, the pH of a solution of a weak acid of known concentration is determined experimentally with a pH meter. From the pH value, the equilibrium concentration of the $[H_3O^+]$ and of the other ion(s) can be calculated.

   Once the equilibrium concentrations of all species present in the equilibrium system are known, the Acid – Ionization Constant can be calculated.

PROCEDURE:

PART I:

ESTIMATING THE ACID IONIZATION CONSTANT OF A WEAK ACID (Ka) BY CONDUCTIVITY TESTING COMPARISONS

1. **<u>Preparing your glassware, plastic-ware and equipment</u>**

   You will need:
   - One clean 400 mL beaker/two students
   - One clean 100 mL graduated cylinder/two students
   - Four clean 250 mL flasks/two students
   - Four clean # 6 rubber stoppers/two students
   - Two clean Chemplates/two students

     The items listed above should be washed with tap water followed by three rinses with D.I water and drained off of excess D.I. water.
   - Disposable plastic droppers or a clean narrow diameter glass tube.
   - Two " Conductivity Tester Combos"/two students, composed of Power Converters connected to Conductivity Indicators.

2. **<u>Preparing your solutions</u>**

   Six solutions will be used in this experiment for conductivity testing.

Solution 1:	$HC_2H_3O_2(aq)$	(Acetic Acid)	0.60 M	- already available
Solution 2:	$HC_2H_3O_2(aq)$	(Acetic Acid)	0.060 M	- to be prepared
Solution 3:	$HC_2H_3O_2(aq)$	(Acetic Acid)	0.0060 M	- to be prepared
Solution 4:	HCl(aq)	(Hydrochloric Acid)	0.10 M	- already available
Solution 5:	HCl(aq)	(Hydrochloric Acid)	0.010 M	- to be prepared
Solution 6:	HCl(aq)	(Hydrochloric Acid)	0.0010 M	- to be prepared

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

- You will find available in the lab two stock solutions, which can be dispensed from either burets or dropper bottles:
  - Solution 1: $HC_2H_3O_2$(aq) (Acetic Acid) 0.60 M
  - Solution 4: HCl(aq) (Hydrochloric Acid) 0.10 M
- Label four 250 mL flasks as follows: 2, 3, 5 and 6

- Add the reagents to the labeled 250 mL flasks as indicated in the table below:
  - Measure and deliver the stock solutions (Solutions 1 & 4) by dispensing them from burets.
  - Measure and deliver the D.I. from a 100 mL graduated cylinder.

Solution	Solution 1 $HC_2H_3O_2$(aq 0.60 M from buret (mL)	Solution 4 HCl(aq) 0.10 M from buret (mL)	D.I.H_2O from graduated cylinder (mL)	Total Volume (mL)	Dilution
Solution 2 $HC_2H_3O_2$(aq) 0.060 M	10.00	----	90.0	100.0	ten fold
Solution 3 $HC_2H_3O_2$(aq) 0.0060 M	1.00	----	99.0	100.0	hundred fold
Solution 5 HCl(aq) 0.010 M	----	10.00	90.0	100.0	ten fold
Solution 6 HCl(aq) 0.0010 M	----	1.00	99.0	100.0	hundred fold

- Tightly seal the four 250 mL flasks with # 6 rubber stoppers.
- Mix the contents of the four 250 mL flasks by slowly inverting the flasks several times.
- Place a stirring magnet in each of the solutions. Depending on the number of stirring magnets and stirring plates available, you may have to proceed with one or two flasks at a time. Place the flask(s) on a stirring plate(s) so as to provide a gentle and uniform mixing of the solution in the flask(s). Stir each of the four solutions with the stirring magnet for about three minutes. Note that improper and/or insufficient mixing is the major sour of error in this experiment. Do not spill any of the contents of these flasks.
  If you do, discard the solution and make a new solution.

3. **Conductivity Testing**
   Check out **two** Conductivity Tester Combos and check the electrodes in the same manner it was done in the previous experiment (Experiment # 5: Acid & Base Strength).

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

4. <u>Testing the reliability of responses obtained from two different conductivity indicators</u>

A. In this experiment (Unlike in Experiment 5) there are **two factors** that determine the electrical conductivity of the solutions tested:

- the nature of the substance tested (strong electrolyte versus weak electrolyte) and
- the concentration of the solution tested (more concentrated versus less concentrated)

As such, in this experiment a good observer is able to detect more than the three standard responses of the conductivity indicator (+, +/- and -) previously detected.

This experiment will challenge your power of observation and your results and conclusions greatly depend on it.

A good and diligent observer will be able to distinguish between five possible responses from the Conductivity Indicator:

A.	Very Bright or Blinking Light indicates	Very Strong Conductivity	**(+/ +)**
B.	Bright Light indicates	Strong Conductivity	**(+)**
C.	Faint Light indicates	Weak Conductivity	**(+/-)**
D.	Very Faint Light indicates	Very Weak Conductivity	**(+/- -)**
E.	Barely Visible Faint Light indicates	Extremely Low Conductivity	**(+/- - -)**

B. Since you will be determining the matching brightness of light intensity of two different solutions, it is necessary to see the two responses of the conductivity indicators simultaneously.

The observation is only valid if the two conductivity indicators provide identical responses.

This will be checked by simultaneously observing the response of two different conductivity indicators immersed in identical solutions.

This step in the procedure is outlined below:

- Prepare two Chemplates and two Conductivity Combos.
- One team member works with one set and the other team member works with the other set.
- Plan ahead and record in which depressions you will place the six solutions to be tested.
- Fill the selected depressions completely with the six solutions to be tested.
  - ➢ Solution 1 & 4 can be dispensed directly from the dropper bottles provided.
  - ➢ Solution 2, 3, 5 & 6 (in 250 mL flasks) may be dispensed by using disposable plastic droppers (if available) or a hollow glass tube with a narrow diameter. If you use a glass tube, make sure you wash it with D.I. water and drain excess D.I water when switching from one solution to another.

C. Using the conductivity indicator, test the conductivity of identical solutions against each other in the same manner as it was done in Experiment # 5.

- Observe and record if you obtain the same response (**Match**) from the two conductivity indicators when testing two identical solutions simultaneously.
  If you do not obtain a match (**No Match**) consult your instructor.
- Record the response of the two conductivity indicators:

  (+/+) **(+)** **(+/-)** **(+/- -)** **(+/- - -)**

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

5. **<u>Testing different solutions of different concentrations for conductivity comparison by matching responses.</u>**
   Two different conductivity testers that have been verified to provide reliable responses will be used.
   - Record in your Lab Notebook the results of your Conductivity Testing.
     You may choose to use symbols:
     (+/+) **(+)** **(+/-)** **(+/- -)** **(+/- - -)**
     or letters:
     (A) (B) (C) (D) (E)
   - Check with your instructor if your observations and recorded data are correct.
   - If time allows it, do some or all of the calculations for this part.

6. **<u>Wrapping up PART I of the experiment</u>**
   a. **<u>The Solutions</u>**
      - **Keep** the solutions of **acetic acid, $HC_2H_3O_2$(aq), (Solutions 2 & 3)** since they will be used in Part II of the experiment.
      - **Discard** the solutions of **hydrochloric acid, HCl(aq) (Solutions 5 & 6)** in the appropriate waste container.
        THE SOLUTIONS SHOULD BE DISCARDED THROUGH A FUNNEL IN ORDER TO CATCH THE STIRRING MAGNETS!
        Wash the stirring magnet with plenty of water followed by a few rinses with D.I Water. **Return the stirring magnets to your laboratory instructor**!

   b. **<u>The Chemplates</u>:**
      - Should be washed well with tap water and rinsed with D.I. water.
      - Could be put away in your locker. They are no longer needed for this experiment.

   c. **<u>The Conductivity Tester Combos</u>**
      - Could be disconnected. They are no longer needed for this experiment.
      - Should be returned in the same condition they were, prior to being checked out.

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

PART II:
DETERMINING THE ACID IONIZATION CONSTANT OF A WEAK ACID (Ka) BY pH MEASUREMENTS

1. Preparing your glassware, plastic-ware and equipment

You will need:

- Three Shell Vials (available in your locker), labeled respectively:
  - Solution 1: $HC_2H_3O_2$(aq 0.60 M
  - Solution 2: $HC_2H_3O_2$(aq 0.060 M
  - Solution 3: $HC_2H_3O_2$(aq 0.0060 M
- Two plastic test tubes with caps (available in your locker), labeled respectively:
  - Buffer, pH = 7.01
  - Buffer, pH = 4.01

The items listed above should be washed with tap water followed by three rinses with D.I water and drained off of excess D.I. water.

2. Preparing your solutions for pH testing

- **The $HC_2H_3O_2$(aq) solutions**
  - **Solution 1**
    Dispense from the buret 10.00 mL of Solution 1: $HC_2H_3O_2$(aq 0.60 M
  - **Solution 2**
    Transfer about 10 mL of Solution 2 from the 250 mL flask into the shell vial, by matching the level of the solution transferred in the shell vial with the level of Solution 1.
  - **Solution 3**
    Transfer about 10 mL of Solution 3 from the 250 mL flask into the shell vial, by matching the level of the solution transferred in the shell vial with the level of Solution 1.
- **The Buffer solutions**
  - **Buffer, pH = 4.01**
    Measure out about 10 mL of buffer 4.01 solution into a graduated cylinder and transfer it into a labeled plastic test tube.
  - **Buffer, pH = 7.01**
    Transfer about 10 mL of buffer 7.01 solution into a second labeled plastic test tube, by matching the level of the buffer 7.01 in the second test tube with the level of the buffer 4.01 in the first labeled plastic test tube (no need to use the graduated cylinder again).
  - NOTE:
    Place each labeled plastic test tube containing the buffer solutions into the test tube- rack.
    This will:
    - Provide easy access to the pH meter during the calibration process, and
    - ensure that the pH meter will be in an upright position at all times.

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

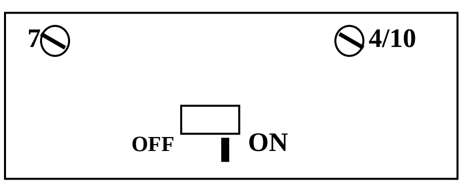
3. **<u>Preparing your equipment</u>**
   a. Check out a small screwdriver from your laboratory instructor.
   You will use the screwdriver to calibrate your pH-meter.
   b. Check out a pH-meter from your laboratory instructor.
   The pH meters are turned "OFF" and stored in a common plastic bucket containing pH = 7.01 buffer solution. **DO NOT TURN THE pH METER "ON" YET!**
   c. Place the pH meter in your labeled plastic test tube containing the pH = 7.01 buffer.

The pH-meter

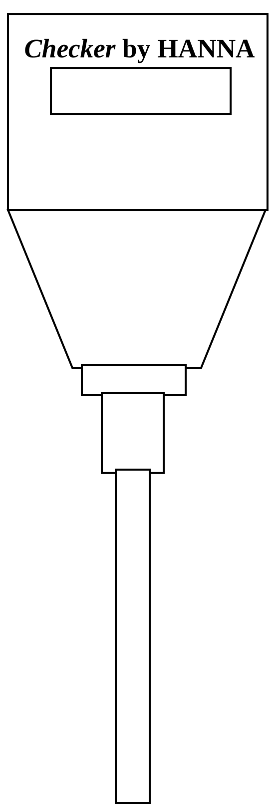
I. Description

A pH meter and its electrodes form a sensitive electrochemical device that will allow an accurate, reproducible, and reliable measurement of the pH of a solution. A pH meter is essentially a voltmeter that measures the voltage of an electric current flowing through a solution between two electrodes. There is a direct relationship between the voltage and the pH of a solution. As a result, the meter on the instrument is calibrated directly in pH units. You will be using a pH meter with a combination stick pH electrode equipped with a screw-type connector for calibration.

Top View



Front View



EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

I. <u>Operating instructions for the pH meter</u>

1. <u>General Guidelines</u>

- **When NOT Taking a pH measurement:**

  <u>DO</u>:
  - keep the pH meter turned "OFF"
  - store the pH meter immersed in the pH = 7.01 buffer solution.

  <u>DO NOT</u>:
  - store the pH meter immersed D.I. water
  - let the pH electrode dry out by lying it horizontally on your bench top.
  - contaminate the solutions when you transfer from one solution to another.

- **<u>When preparing to take a pH measurement</u>**

Photo by Andrew Huertas

➤ During the transfer:
- Rinse the electrode with a stream of deionized water.
- Remove the excess water from the electrode with tissue paper (Kimwipe) before you immerse the electrode in the next solution.
- Do not touch the electrode with your hands.
- Handle the electrode with care, since it is fragile.

2. <u>Calibrating the pH meter:</u>

(a) <u>First Calibration</u>

The pH meter must be **calibrated** with solutions whose pH values are known before you can measure with accuracy the unknown pH of a solution.
The solutions used for calibration are called **buffer solutions.**
Commonly three buffer solutions are used for calibration, with the following pH values: **pH = 4.01** **pH = 7.01** **pH = 10.01**
The first calibration is always done with pH = 7.01 buffer

Photo by Andrew Huertas

- Turn the pH meter "ON".
- Stir gently the electrode in the pH = 7.01 buffer solution and allow the pH reading to stabilize.
- Use a small screwdriver to adjust the pH = 7 trimmer until the display reads "**7.01**" (or as close as possible to this value).

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

(b) <u>Second Calibration</u>

The second calibration can be done with either pH = 4.01 buffer or pH = 10.01 buffer, depending on the character of the solution to be tested (acidic or basic)
To do the second calibration of the pH meter, you will need to use the buffer solution(s) whose pH value(s) is (are) in relatively close range with the assumed pH (acidic or basic) of your solution. In this experiment the pH-meter will be used to measure the pH of acidic solutions. As such, the second calibration of the pH meter will be performed with a pH = 4.01 buffer solution

- Remove the electrode from the pH= 7.01 buffer solution
- Rinse the electrode with a stream of deionized water over the sink or catch the water in a beaker.
- Remove the excess water from the electrode with tissue paper (Kimwipe) before you immerse the electrode in the pH = 4.01 buffer solution..
- Immerse the tip of the electrode (bottom 4 cm/1½) in a sample of pH = 4.01 buffer solution.
- Stir gently the electrode in the pH = 4.01 buffer solution and allow the pH reading to stabilize.
- Use a small screwdriver to adjust the pH 4/10 trimmer until the display reads "**4.01**" (or as close as possible to this value).
- ➢ Leave the pH meter in the pH = 4.01 buffer solution, until you are ready to take the pH measurements of your samples.
- ➢ Turn the pH meter "OFF"

3. <u>Taking the pH measurements</u>

- Remove the electrode from the pH= 4.01 buffer solution.
- Rinse the electrode with a stream of deionized water over the sink or catch the water in a beaker.
- Remove the excess water from the electrode with tissue paper (Kimwipe) before you immerse the electrode in the first solution to be tested.
- You will now be taking the pH measurements of the three $HC_2H_3O_2(aq)$, solutions placed in the three shell vials (Solution 1, 2 & 3):
- Follow the steps described below to measure the pH of the three solutions.
  <u>NOTE</u>:
  To avoid contamination through the electrodes, when transferring the pH meter from one solution to the other, it is best to test the solutions in order of increasing concentration, from the most dilute to the most concentrated.
  (First: Solution 3; Second: Solution 2; Third: Solution 1)

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM


Photo by Andrew Huertas

- Immerse the electrode (bottom 4 cm/1½) in the solution whose pH you are measuring.
- Turn the pH meter "ON".
- Stir gently the electrode in the solution you are testing and allow the pH reading to stabilize. You should be able to read the pH of the solution about 20 seconds after the electrode has been immersed. The reading should be steady and not suddenly changing.
- Record the pH reading in your Laboratory Notebook
- Turn the pH meter "OFF"

- ➢ Remove the electrode from the solution you have tested
- ➢ Rinse the electrode with a stream of deionized water and remove the excess water from the electrode with tissue paper (Kimwipe) before you immerse the electrode in the next solution to be tested.

4. **Storing the pH-meter after use.**
   - **Make sure the pH meter is turned "OFF"!**
   - Rinse the electrode with a stream of deionized water over the sink or catch the water in a beaker.
   - Remove the excess water from the electrode with tissue paper (Kimwipe)
   - **When finished, place the pH meter in the common plastic bucket containing the pH = 7.01 buffer solution.**

5. **Storing the two buffer solutions**
   - **DO NOT DISCARD THE BUFFER SOLUTIONS**!
   - The plastic test-tubes containing the two buffer solutions should be kept in your locker (capped) for the next experiment. The test-tubes come with a cap and they should be well capped to preserve the contents.

6. **Discarding the solutions**
   - Discard the solutions of acetic acid (Solutions 1, 2 & 3) in the appropriate waste container.
   - If a stirring magnet has been used**,** the solutions should be discarded through a funnel in order to catch the stirring magnet.
   - Wash the stirring magnet with plenty of water followed by a few rinses with D.I water.
   - Return the stirring magnets to your laboratory instructor, if so instructed.

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

PART III:
EFFECT OF DILUTION ON THE DEGREE OF IONIZATION OF A WEAK ACID

There is no experimental work involved in this part of the experiment.
If time allows, you may wish to work on this part of the experiment in class, taking advantage of the laboratory instructor's assistance. Alternatively, you may choose to complete the calculations and answer the questions at home.
The "PRINCIPLES" section of this experiment outlines the basic concepts behind this part of the experiment and the data tables given in the "Report Form" are almost self-explanatory.

Bibliography:
1. Nivaldo J. Tro "Chemistry: A Molecular Approach", Third Edition
2. William E, Bull, William T. Smith, Jr. and Jesse H. Wood "Laboratory Manual for College Chemistry", Sixth Edition

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

REPORT FORM

NAME: ____________________________ Date: _____________ Partner: ____________________________

PART I:

ESTIMATING THE ACID – IONIZATION CONSTANT OF A WEAK ACID (Ka) BY CONDUCTIVITY TESTING COMPARISONS.

1. **Testing the reliability of responses obtained from two different conductivity indicators.**
   Record your observations in the DATA TABLES # 1 & 2 below.
   One box in DATA TABLE # 2 is completed for you as an example.

DATA TABLE # 1

Conductivity Tester # 1 (columns); **Conductivity Tester # 2** (rows)

	Solution 4 HCl(aq) 0.10 M	Solution 5 HCl(aq) 0.010 M	Solution 6 HCl(aq) 0.0010 M
Solution 4 HCl(aq) 0.10 M	(+/+) MATCH (+/+)		
Solution 5 HCl(aq) 0.010 M			
Solution 6 HCl(aq) 0.0010 M			

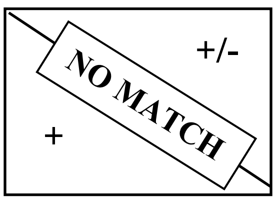
DATA TABLE # 2

Conductivity Tester # 1 (columns); **Conductivity Tester # 2** (rows)

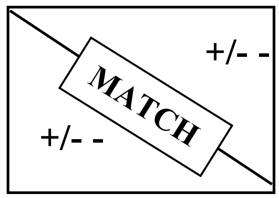
	Solution 1 $HC_2H_3O_2(aq)$ 0.60 M	Solution 2 $HC_2H_3O_2(aq)$ 0.060 M	Solution 3 $HC_2H_3O_2(aq)$ 0.0060 M
Solution 1 $HC_2H_3O_2(aq)$ 0.60 M			
Solution 2 $HC_2H_3O_2(aq)$ 0.060 M			
Solution 3 $HC_2H_3O_2(aq)$ 0.0060 M			

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

2. **Testing different solutions of different concentrations for conductivity comparison by matching responses.**
   Record your observations in the DATA TABLE # 3 below.
   Your observations can be entered in the appropriate boxes in one of the following manners:



OR



DATA TABLE # 3

Conductivity Tester # 2	**Conductivity Tester # 1**	Solution 1 $HC_2H_3O_2(aq)$ 0.60 M	Solution 2 $HC_2H_3O_2(aq)$ 0.060 M	Solution 3 $HC_2H_3O_2(aq)$ 0.0060 M
	Solution 4 HCl(aq) 0.10 M			
	Solution 5 HCl(aq) 0.010 M			
	Solution 6 HCl(aq) 0.0010 M			

3. **<u>Interpreting your observations</u>**:
   - What is the concentration of the strong acid (HCl) that exhibits the same electrical conductivity as one of the solutions of the weak acid ($HC_2H_3O_2$)?
     ____________ M HCl

   - If two solutions of two different acids and of two different concentrations exhibit the same electrical conductivity, what is the relationship between their **<u>concentrations of ions</u>** in their respective solutions?

     ____

     ____

Concentration of ions in _____________ M HCl(aq)	=	Concentration of ions in _____________ M $HC_2H_3O_2(aq)$

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

4. **<u>Calculations</u>**:
   Using the information above, construct and complete the Acid Ionization Equilibrium Table for the ionization of acetic acid ($HC_2H_3O_2$):

Acid Ionization Equilibrium Equation:			$\rightleftharpoons$		
Initial Conc.					
Change:					
Equil. Conc.					

1. Calculate the value of Ka for acetic acid ($HC_2H_3O_2$) using the equilibrium concentrations from the equilibrium table above.

2. The estimated value for the Acid ionization Constant, Ka, of acetic acid, $\mathbf{HC_2H_3O_2}$**,** as determined by the Conductivity Comparison Method is:

   Ka ($\mathbf{HC_2H_3O_2}$) estimated: **Ka =**

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

PART II:

DETERMINING THE ACID - IONIZATION CONSTANT OF A WEAK ACID (Ka) BY pH MEASUREMENTS.

1. Determine the equilibrium concentrations of all species in the acetic acid, $\mathbf{HC_2H_3O_2}$**,** solutions from pH measurements.

DATA TABLE 5

Solution	**Molarity of [$HC_2H_3O_2$]**	**Measured pH**	**[H^+](aq) or [H_3O^+](aq)**
1	**0.60 M**		
2	**0.060 M**		
3	**0.0060 M**		

Solution 1

Acid Ionization Equilibrium Equation: $HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$

Initial Concentrations		excess		
Change		excess		
Equilibrium Concentrations		excess		

Solution 2

Acid Ionization Equilibrium Equation: $HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$

Initial Concentrations		excess		
Change		excess		
Equilibrium Concentrations		excess		

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

Solution 3

Acid Ionization Equilibrium Equation:	$HC_2H_3O_2(aq)$	+ $H_2O(l)$	$\rightleftharpoons$	$H_3O^+(aq)$	+ $C_2H_3O_2^-(aq)$
Initial Concentrations		excess			
Change		excess			
Equilibrium Concentrations		excess			

2. Calculate Ka for acetic acid, $HC_2H_3O_2$, from the equilibrium concentrations of the species present In the Acid Ionization Equilibrium.

DATA TABLE 6

	$HC_2H_3O_2(aq)$ M	$H_3O^+(aq)$ M	$C_2H_3O_2^-(aq)$ M	Ka (refer to page 2)
Solution 1				
Solution 2				
Solution 3				
			Average Ka ⇨	

3. The calculated Average Value for the Acid ionization Constant, Ka, of acetic acid, **$HC_2H_3O_2$,** as determined by pH measurements is:

Ka ($HC_2H_3O_2$) calculated : | |

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

PART III:

1. **EFFECT OF DILUTION ON THE DEGREE OF IONIZATION OF A WEAK ACID [$HC_2H_3O_2(aq)$]**

   Calculate the percentage ionization of the solutions of the weak acid **($HC_2H_3O_2$)**, of decreasing concentrations, by using the concentration of [H^+] obtained from pH measurements.

Concentration of $HC_2H_3O_2(aq)$	pH (experimental)	$[H^+] = [C_2H_3O_2^-]$	% Ionization = $\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]} \times 100$
0.60 M			
0.060 M			
0.0060 M			

2. **CONCLUSIONS:**

(a) What is the relationship between the concentration of a solution of a weak acid [**$HC_2H_3O_2(aq)$**]and its Degree (%) Ionization?

The ____________________ the concentration, (the more____________________ the solution)
(lower/higher) (dilute/concentrated)

____________________ the Degree of Ionization.
(greater/smaller)

(b) Provide a rationale for your stated conclusion above (a), by completing the Equilibrium Table that illustrates the Acid Ionization Equilibrium of $HC_2H_3O_2(aq)$, being gradually diluted (water added to the solution).

Equation:	[ ]	+	[ ]	$\rightleftharpoons$	[ ]	+	[ ]	
Stress:....................................			[ ]					
Shift:..				[ ]				
New Equilibrium...... (Increased or Decreased)	[ ]	......	[ ]	....	[ ]	......	[ ]	

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM

(c) Provide a rationale for your stated conclusion above (a), by explaining it in terms of Le Chatelier's Principle.

____

____

____

____

____

3. EXTENDING THE CONCEPTS:

pH measurements have been taken for two different acidic solutions of the same concentration.
The pH of a 0.00001 M ($1x10^{-5}$ M) HCl(aq) solution had been determined to be about 5.0
The pH of a 0.00001 M ($1x10^{-5}$M) $HC_2H_3O_2$ (aq) solution had been determined to be 4.9

Question:

How do you account for the fact that the pH values of a solution of a strong acid and of a solution of a weak acid, of the same concentration, have almost the same pH values?

____

____

____

____

____

____

EXPERIMENT 6
A STUDY OF AN ACID IONIZATION EQUILIBRIUM