

Experiment 8

pH, Acid-Base Indicators, and K_a Equilibrium

OBJECTIVE

The colours of several acid-base indicator substances in water solution will be observed over the pH range 2 to 12. A wide-range pH paper and a pH meter will be used to measure the pH of several common substances. A series of acetic acid solutions of different concentrations will be prepared. The pH of each solution will be measured; the K_a and percent ionization of the acetic acid in each solution will be calculated.

- The colour of several acid-base indicator substances will be observed at different pH values.
- Wide-range pH paper and a pH meter will be used to measure pH of various common substances.
- Volumetric glassware will be used to carry out a serial and multiple dilution of an acetic acid solution of known molarity.
- The pH meter will be used to measure the pH of the initial solution and all of the diluted solutions of the acetic acid.
- Calculations will be performed to determine the concentration, the acid ionization constant, and the percent ionization of all of the acetic acid solutions.

REFERENCES

Timberlake	3 rd Edition	4 th Edition
➤ Molarity and Dilution	Chapter 10 8.5 Pages 316-320	Chapter 10 8.4, 8.5 Pages 315-321
➤ Strengths of Acids and Bases ➤ Ionization of Water ➤ The pH Scale ➤ Reactions of Acids and Bases ➤ Buffers	Chapter 10 10.2 Pages 377-382 10.3 Pages 383-386 10.4 Pages 386-392 10.5 Pages 393-398 10.7 Pages 401-404	Chapter 10 10.2 Pages 375-381 10.3 Pages 381-385 10.4 Pages 384-391 10.5 Pages 391-396 10.6 Pages 397-401
➤ Figure 10.4	Chapter 10 Page 387	Chapter 10 Page 385

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Wikipedia	URL
➤ Acids	http://en.wikipedia.org/wiki/Acid
➤ pH	http://en.wikipedia.org/wiki/PH
➤ pH Indicator	http://en.wikipedia.org/wiki/PH_indicator

INTRODUCTION

The pH Scale

The concentration of the hydrated hydrogen ion, H^+ (aq), also written and referred to as the hydronium ion, H_3O^+ (aq), in aqueous solutions varies from a value above **10 M** in concentrated acid solutions to less than **1.0×10^{-14} M** in concentrated alkali solutions.

To simplify thinking about this huge range of concentration values, a negative logarithmic scale called the pH scale has been adopted:

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+ (\text{aq})] = -\log_{10} [\text{H}^+] \quad [\text{H}^+] = 10^{-\text{pH}}$$

The now universally used measure of the hydrogen ion concentration was introduced in a publication of the Danish biochemist Sørensen in 1909, relating to work carried out at the Carlsberg Breweries. The hydrogen ion concentration of aqueous solutions is a major factor in the chemistry of all living organisms and biochemical substances. It is very important in many areas of industrial chemistry, in food chemistry, in biotechnology, and in the chemistry of the environment. Measurement and control of pH is big business.

Pure water dissociates very slightly, producing equal concentrations of hydronium and hydroxide ions.



The equilibrium constant expression for this reaction is:

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14}$$

At **25 °C**, the molar concentration of hydronium ion, $[\text{H}_3\text{O}^+]$, equals the molar concentration of the hydroxide ion, $[\text{OH}^-]$ and both equal **1.0×10^{-7} M** resulting in a neutral pH of 7 for pure water.

The value of pH and pOH can also be determined using the formula:

$$\text{pH} + \text{pOH} = 14 \quad (\text{Where } \text{pOH} = -\log [\text{OH}^-])$$

Calculating pH and pOH of Aqueous Solutions

Addition of an acid to a solution increases the $[\text{H}_3\text{O}^+]$ of a solution, causing a pH of less than 7. This also decreases the value of $[\text{OH}^-]$.

Adding a base increases the $[\text{OH}^-]$ in the solution, causing a pH greater than 7. This also decreases the value of $[\text{H}_3\text{O}^+]$.

For strong acids such as **HCl**, **HNO₃**, and **H₂SO₄** the hydronium ion concentration is assumed to be equal to the stoichiometric concentration calculated from the molarity of the acid.

For strong alkalis such as **NaOH**, **KOH**, and **Ca(OH)₂** the hydroxide ion concentration is assumed to be equal to the stoichiometric concentration calculated from the molarity of the alkali.

Acid-Base Indicators

Acid-base indicator substances are usually organic chemical substances with large complex molecules. They are dyestuffs which are also weak acids (or weak bases). In water solution, they ionize and / or change their structure by reacting with water, or with acid or alkali. An acid-base indicator substance must have at least two forms of different visible colour to be useful. Some indicators have more than two forms (see **Box** below).

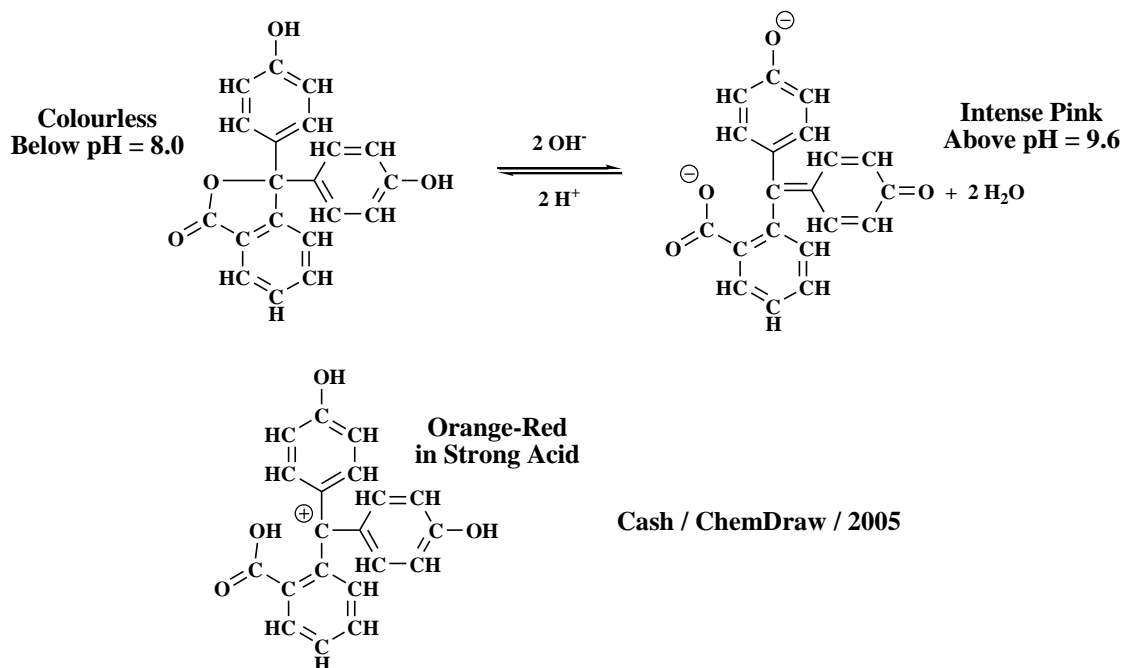
Of the two useful forms of the indicator substance at least one must absorb visible light intensely. An indicator changes colour at the hydrogen ion concentration that matches its acid ionization constant value (when $\text{pH} = \text{pK}_{\text{In}}$). Below this pH, the colour of a conjugate acid form is seen, and above this pH, the colour of a conjugate base form is seen. In some cases, an intermediate or mixed colour is seen at the pH of the change.

Box: The Three Forms and Colours of the Acid – Base Indicator Phenolphthalein¹

For more than 100 years, phenolphthalein was used as the active medicinal ingredient in many over-the-counter laxative products. This ended in 1997 when the U.S. Food and Drug Administration (FDA) classed it as a possible cancer causing chemical.

For chemists, phenolphthalein is a very useful acid-base indicator substance.

The most useful transition for phenolphthalein is between pH 8.0 and 9.6. The colour change from colourless below pH 8.0 to intensely pink above pH 9.6 makes this indicator very useful for weak acid – strong base titrations. Phenolphthalein has another form in very strong acid that is orange-red in colour.



¹ G. Wittke, J. Chem. Ed., **60**, 239, 1983; Daniel C. Harris, Quantitative Chemical Analysis, 3rd Edition, Freeman, 1991, pages 243-244.

Universal Indicator

A single indicator can only be used to determine that a solution is below or above a particular value of pH. By selecting a set of four or five indicators, each one of which changes colour at a different pH, it is possible to make an indicator mixture that displays numerous colours. A well chosen mixture will show a different and distinct colour at every pH value. A mixed indicator can be narrow-range, over 2 or 3 pH units, or wide-range (universal), over the entire pH range. Mixed indicator, whether purchased as a solid or as a pre-dissolved solution, is more expensive than a single indicator. Cost of universal indicator solution (0 – 14 pH) is about \$ 11 per 25 mL (2005).

pH Paper

A pH paper is produced by soaking a porous paper, similar to filter paper, in a solution containing a single or mixed indicator, and then drying the paper. Litmus paper contains only the single indicator **litmus**, while narrow-range and wide-range papers contain indicator mixtures. The mixed indicator papers are relatively expensive.

Cost of red / blue litmus indicator paper is about \$ 2.40 per 100 strips (2005).

Cost of universal indicator paper (0 – 14 pH) is about \$ 20.25 per 100 strips (2005).

Buffers

A buffer solution is a solution containing a substance or a mixture of substances capable of opposing a change in pH due to the addition of either an acidic substance or a basic substance to the solution. Up to its capacity limit, a buffer solution will maintain a stable value of pH. All biological fluids are buffered, since even small changes in pH are very dangerous for living organisms.

The pH Meter

The pH of an aqueous solution unknown can be estimated approximately using acid-base indicators or pH paper, but a more accurate and precise measurement requires an instrument called a pH meter. A pH meter is a voltmeter connected to two electrodes that are dipped into the solution unknown. One of the electrodes, called the reference electrode, is not affected by hydrogen ions. The other electrode, called the pH glass electrode, is affected by the concentration of the hydrogen ions present. The voltage between the electrodes varies in a way that is well understood and can be related to the hydrogen ion concentration or pH of the unknown solution. The pH value of the solution tested is usually shown on a dial or a digital display. The present generation of pH meters are solid state, microprocessor controlled electronic instruments.

The pH Glass Electrode

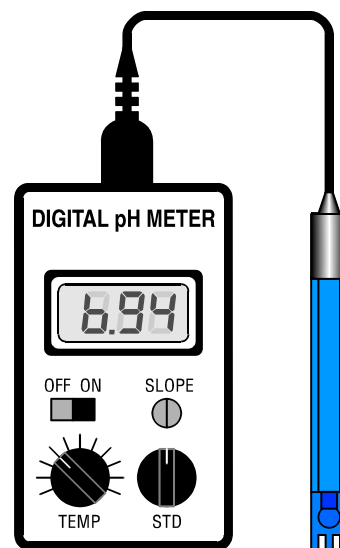
The pH glass electrode is an example of an ion-selective electrode, an electrode that is affected by the variation of the concentration of a specific substance in a solution. The tip of the electrode is made of a very thin layer of a special glass. This glass tip is very delicate and also it must not be allowed to dry out. If it dries out, it may require several days of soaking in water before it is usable.

The Combination Electrode

The reference electrode and the pH glass electrode and their wiring are both built into a single probe called a combination pH electrode. It may seem as if there is only one electrode, but there really must be two present.

Combination electrodes are relatively robust, but the fragile glass of the tip is protected by a plastic guard. When stored away, a water soaked cap is used to keep the tip wet.

A new generation of electrodes is now available with solid state electrodes, using an electric field effect. They are very robust, they need not be kept wet, but they are expensive.



The pH Meter and Combination Electrode

The Acid Ionization Constant of Acetic Acid

Vinegar is a dilute solution of the weak acid acetic acid in water. The equation of the reaction of acetic acid with water is given below. Also given is the equation of the K_a expression, referred to as the acid ionization constant (sometimes called the acid dissociation constant):



$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

By convention (tradition), the hydrated hydrogen (hydronium) ion is abbreviated as H^+ in equations for simplicity.

- In solutions of weak acids, it is assumed as an approximation that all of the hydrogen ion present has been produced by the reaction of the acid with water.
- It is also assumed that the concentration of the conjugate base of the acid (acetate ion in this case) is present at the same concentration as hydrogen ion since they are formed in a 1 mol to 1 mol reaction process.
- Both of these assumptions fail if the acid is extremely weak or if the solution is extremely dilute.

Calculation of the Value of the Acid Ionization Constant of Acetic Acid

If the nominal or original or **label concentration** of an acetic acid solution is known, then the value of K_a for the acid can be determined by a pH measurement and calculations.

Use a simplified equation, and assume that the H^+ ion in the solution is entirely due to dissociation of the acid.



$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

Substance	CH_3COOH	H^+	CH_3COO^-
Initial Molarity	Label Value	0	0
Molarity at Equilibrium	Label Value - x	+ x	+ x

The value of “x” is determined from the pH of the solution: $x = [H^+] = 10^{-pH}$

The Percent Ionization of Acetic Acid

The percent ionization (sometimes referred to as the percent dissociation) of an acetic acid solution is calculated from the expression:

$$\text{Percent Ionization (\% Ion.)} = \frac{[CH_3COO^-]_{\text{actual}}}{[CH_3COOH]_{\text{label}}} \times 100\%$$

Concentration at Equilibrium
 ↙
 ↘
 Label Concentration
 (Not Concentration at Equilibrium)

Sample Calculations

Example 1

Calculate the $[H^+]$ and the **pH** of a 0.076 M HNO_3 (**Nitric Acid**) solution. State the pH to 2 places after the decimal point.

Answer

Nitric acid is a strong acid which dissociates completely in water:



$$[H^+] = [HNO_3] = \underline{0.076 \text{ M}}$$

$$pH = -\log [H^+] = -\log_{10} (0.076) = \underline{1.12}$$

The Sample Calculations Section Continues on the Next Page →

Sample Calculations (Cont.)**Example 2**

Calculate the $[\text{OH}^-]$, the **pOH**, and the **pH** of a $\underline{1.5} \times 10^{-3} \text{ M NaOH}$ (Sodium Hydroxide) solution. State the pOH and the pH to 2 places after the decimal point.

Answer

Sodium hydroxide is a strong alkali which dissociates completely in water.



$$[\text{OH}^-] = [\text{NaOH}] = \underline{1.5} \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log [1.5 \times 10^{-3}] = \underline{2.82}$$

$$\text{pH} = 14 - \text{pOH} = 14.00 - 2.82 = \underline{11.18}$$

Example 3

The pH of an acetic acid solution with a label concentration of $\underline{0.750} \text{ M}$ was found to be $\underline{2.45}$. Calculate the value of the hydrogen ion concentration in the solution. State the answer to 3 significant digits.

Answer

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+ \text{ (aq)}] = -\log_{10} [\text{H}^+]$$

$$[\text{H}^+] = 10^{-\text{pH}}$$

$$[\text{H}^+] = 10^{-2.45} = \underline{0.00355} \text{ M}$$

The Sample Calculations Section Continues on the Next Page →

Sample Calculations (Cont.)**Example 4**

Calculate the experimental value of the actual equilibrium concentration of acetate ion and acetic acid in the solution of **Example 3**. State the answers to 3 significant digits.

Answer

Substance	CH ₃ COOH	H ⁺	CH ₃ COO ⁻
Initial Molarity	Label Value	0	0
Molarity at Equilibrium	Label Value - x	+ x	+ x

Assume that $[\text{CH}_3\text{COO}^-]_{\text{actual}} = [\text{H}^+] = 10^{-2.45} = \underline{0.00355 \text{ M}}$

Then $[\text{CH}_3\text{COOH}]_{\text{actual}} = [\text{CH}_3\text{COOH}]_{\text{label}} - [\text{CH}_3\text{COO}^-]_{\text{actual}}$

$$= (0.750 \text{ M} - 0.00355 \text{ M}) = 0.74645 \text{ M} = \underline{0.746 \text{ M}}$$

Example 5

Calculate the experimental value of K_a for the acetic acid solution of **Example 3**. State the answer in scientific notation to 3 significant figures.

Answer

Use the equilibrium constant expression and the values calculated in **Examples 3 and 4**:

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{(0.00355)^2}{0.746} = \underline{1.68} \times 10^{-5}$$

The Sample Calculations Section Continues on the Next Page →

Sample Calculations (Cont.)**Example 6**

Calculate the experimental **Percent Ionization (% Ion.)** of the acetic acid solution of **Example 3**.

State the answer to 2 significant figures.

Answer

$$\text{Percent Ionization (\% Ion.)} = \frac{[\text{CH}_3\text{COO}^-]_{\text{actual}}}{[\text{CH}_3\text{COOH}]_{\text{label}}} \times 100 \%$$

$$\text{Percent Ionization (\% Ion.)} = \frac{0.00355 \text{ M}}{0.750 \text{ M}} \times 100 \%$$

$$\% \text{ Ion.} = \underline{0.47} \%$$

Example 7

Calculate the **label molarity** of an acetic acid solution prepared by the precise dilution of **20.00 mL** of the acetic acid solution of **Example 3 (0.750 M Acetic Acid)** to a total volume of **200.0 mL**.

State the answer to 3 significant figures.

Answer

Review **Experiment 7, Dilution** on page Error! Bookmark not defined. and **Example 4** on page Error! Bookmark not defined..

Chose any convenient form of the dilution equation: $C_{\text{conc}} V_{\text{conc}} = C_{\text{dil}} V_{\text{dil}}$

$$C_{\text{conc}} = \underline{0.750 \text{ M}} \quad V_{\text{conc}} = \underline{20.00 \text{ mL}} \quad C_{\text{dil}} = \underline{\text{Unknown}} \quad V_{\text{dil}} = \underline{200.0 \text{ mL}}$$

Rearrange the equation, substitute and solve:

$$\begin{aligned} C_{\text{dilute}} &= C_{\text{concentrated}} \times \frac{V_{\text{concentrated}}}{V_{\text{dilute}}} = 0.750 \text{ M} \times \frac{20.00 \text{ mL}}{200.0 \text{ mL}} \\ &= \underline{0.0750 \text{ M Acetic Acid}} \end{aligned}$$

The Sample Calculations Section Continues on the Next Page →

Sample Calculations (Cont.)**Example 8**

Calculate the **label molarity** of an acetic acid solution prepared by the precise dilution of **10.00 mL** of the acetic acid solution of **Example 7 (0.0750 M Acetic Acid)** to a total volume of **200.0 mL**.

State the answer to 3 significant figures.

Answer

Chose any convenient form of the dilution equation: $C_{\text{conc}} V_{\text{conc}} = C_{\text{dil}} V_{\text{dil}}$

$$C_{\text{conc}} = \underline{0.0750 \text{ M}} \quad V_{\text{conc}} = \underline{10.00 \text{ mL}} \quad C_{\text{dil}} = \underline{\text{Unknown}} \quad V_{\text{dil}} = \underline{200.0 \text{ mL}}$$

Rearrange the equation, substitute and solve:

$$\begin{aligned} C_{\text{dilute}} &= C_{\text{concentrated}} \times \frac{V_{\text{concentrated}}}{V_{\text{dilute}}} = 0.0750 \text{ M} \times \frac{10.00 \text{ mL}}{200.0 \text{ mL}} \\ &= \underline{0.00375 \text{ M Acetic Acid}} \end{aligned}$$

Name		Day		Start Time	
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PROCEDURE

Parts A, B and C will be done in bench groups as a rotation.

Parts D, E, and F should be done with a partner.

One or more pH meters will be set up for use during the experiment. The use of the pH meter is described in the General Information section (The pH Meter).

- The instructor will demonstrate the use and care of the pH meter and combination electrode.
- Do not touch the tip of the glass electrode and do not allow it to dry out.
- Always leave the electrode immersed in distilled water when not in use.

A. Acid-Base Indicator Colours (Bench Groups)

- A-1. Clean six of the white glazed-porcelain spot plates containing 3×4 rows of small hemispherical depressions. Place the plates so that there are six horizontal rows of twelve columns of depressions in front of you.

	pH of Buffer: 2 3 4 5 6 7 8 9 10 11 12											
Methyl Orange →	○	○	○	○	○	○	○	○	○	○	○	○
Bromophenol Blue →	○	○	○	○	○	○	○	○	○	○	○	○
Litmus →	○	○	○	○	○	○	○	○	○	○	○	○
Phenolphthalein →	○	○	○	○	○	○	○	○	○	○	○	○
Thymol Blue →	○	○	○	○	○	○	○	○	○	○	○	○
Universal Indicator →	○	○	○	○	○	○	○	○	○	○	○	○

- A-2. Small squeeze bottles containing **buffer solutions** of pH from 2 up to 12 will be supplied. Half-fill the six depressions in the first **vertical column** with pH 2 buffer. Continue, half-filling each successive vertical column with pH 3, 4, etc. up to pH 12.
- A-3. Small squeeze bottles containing **acid-base indicator solutions** will be supplied. Place one drop of **Methyl Orange** indicator solution into each depression of one **horizontal row** (there will be one drop at each pH from 2 to 12). Add more indicator if the colour is faint.

Stir if necessary to mix well, but do not contaminate other depressions.
Rinse the stirring rod with distilled water after each use.

- A-4. Repeat, placing all of the other indicator substance solutions listed in the illustration above, each in a separate row. Observe the colour at each pH for each substance. Describe or name each of the colours seen as a function of pH in the **Indicator Substances Table** in the **REPORT PAGES** Section on page 209.

This Section of the PROCEDURE Continues on the Next Page →

A. Acid-Base Indicator Colours (Bench Groups) (Cont.)

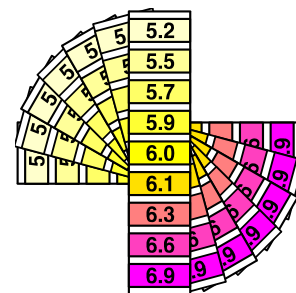
- A-5. Compare the colours observed for the **universal indicator** substance to the colour chart provided for that indicator. Note that this chart applies **only** to this particular universal indicator substance from this particular manufacturer.
- A-6. **Instructor Demonstration In the Fumehood: Verify Fan Operation.**
Into one **clean, dry** depression of a spot plate place **3 drops** of **concentrated** sulphuric acid. Add **3 drops** of phenolphthalein indicator solution. Observe and record the colour seen in the **Indicator Substances Table** on page 209.
- A-7. Have the instructor initial your table when it is completed.

B. pH Determination of Buffer Solutions (Bench Groups)

- B-1. You will require a number of clean **small** beakers and stirring rods.
- B-2. A **small** strip of **wide-range** pH paper and a colour chart will be supplied. This paper is to be used in **Parts B** and **C**. This paper is expensive, but very little is needed.

Cut the paper into small squares or strips with a pair of scissors, about 2 or 3 mm on a side.

- B-3. Place these **small** pieces of paper on a white background, well separated for use. The depressions of a spot plate or a clean round of filter paper can be used as a white background.



**Narrow-Range pH Paper
Colour Chart**

The pH paper works only when wetted with the solution to be tested.

Tests are made by bringing the solution to be tested to the paper, in the form of a small drop on a stirring rod.

- B-4. Two buffer solutions are provided for testing, one of **pH** about **4** (acidic) and the other of **pH** about **9** (alkaline). Test each solution in turn. Take about **10 mL** of each buffer solution into a clean **small** beaker. Rinse the beaker with the solution, and discard that portion.
- B-5. Take a fresh **25 mL** portion of buffer into its beaker. Test the pH of each buffer solution using the wide range pH paper. Record the colour observed and your pH conclusion for the buffer in the **Buffer Solutions Table** in the **REPORT PAGES** Section on page 20.

Use the colour comparison chart supplied with the paper to estimate the approximate pH of each buffer. Note that this chart applies only to this particular paper from this manufacturer.

This Section of the PROCEDURE Continues on the Next Page →

B. pH Determination of Buffer Solutions (Bench Groups) (Cont.)

- B-6. Next, use the pH meter to measure the pH of each of the two buffer solutions. In each case, rinse the electrodes with the portion of the test solution in the beaker and discard that portion.

Use a second, fresh portion of the buffer solution for the pH measurement. Swirl the solution gently around the electrodes to be sure there are no air bubbles on the surface.

Turn the meter to **On or pH**. Read and record the value observed within 15 seconds in the **Buffer Solutions Table**.



Photo: Emily Girard 3CH 2005

Record the value of the pH to 2 places after the decimal point.

- B-7. Return the meter to **Standby** and rinse the electrodes, leaving them in distilled water unless you have further measurements to be made. Have the instructor initial your table.

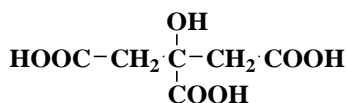
C. pH Determination of Common Substances (Bench Groups)

- C-1. See **Steps B-1 to B-3** for instruction in the use of pH paper.
- C-2. Use the wide-range pH paper and the colour chart to determine the approximate pH of each of the available common substances listed in the **Common Substances Table** in the **REPORT PAGES** Section on page 211 (see also the **Box** on page 204).

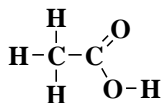
If there are other available substances to test, list them in the **Common Substances Table**. Record your observations of colour and your approximate pH assignments in the **Table**.

- C-3. Use the pH meter to measure the pH of each solution listed in the **Common Substances Table**. First, rinse the electrodes with a portion of the test solution, and discard that portion.
- C-4. Use a second, fresh portion for the pH measurement. Swirl the solution gently around the electrodes to be sure there are no air bubbles on the surface. Turn the meter to **On or pH**. Read and record the value observed within 15 seconds in the **Common Substances Table**. Record the value of the pH to 2 places after the decimal point.
- C-5. Return the meter to **Standby** and rinse the electrodes, leaving them in distilled water unless you have further measurements to be made. Have the instructor initial your table.

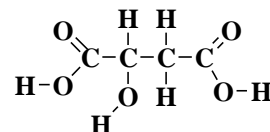
Box: Substances in Common Fluids²



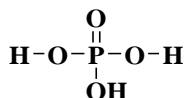
Citric Acid $K_1 = 8.4 \times 10^{-4}$
(Lemon Juice and 7-UP / Sprite)



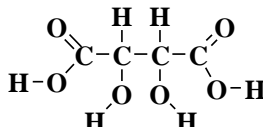
Acetic Acid $K_a = 1.8 \times 10^{-5}$
(Vinegar)



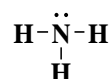
Malic Acid $K_1 = 3.9 \times 10^{-4}$
(Apple Juice, Wine)



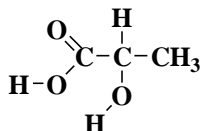
Phosphoric Acid $K_1 = 7.5 \times 10^{-3}$
(Cola Drinks)



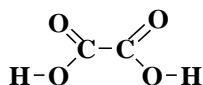
Tartaric Acid $K_1 = 1.0 \times 10^{-3}$
(Grape Juice, Wine)



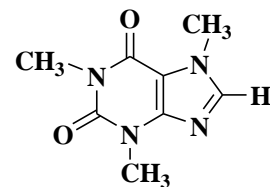
Ammonia $K_b = 1.8 \times 10^{-5}$
(Window Cleaner)



Lactic Acid $K_a = 8.4 \times 10^{-4}$
(Sour Milk, Wine)



Oxalic Acid $K_1 = 6.5 \times 10^{-2}$
(Plant Leaves)



Caffeine $K_b = 2.4 \times 10^{-11}$
(Coffee, Tea, Cola)

² Organic Structures and Acid Ionization Constants from: Chemistry in The Marketplace – A Consumer Guide, 4th Edition, Ben Selinger, HBJ, 1989, Table 13.7, page 551.

Work with a partner assigned by the instructor on Parts D, E, and F.

Report the name of your partner on page 213.

Four (4) acetic acid solutions of known concentration will be prepared by precise dilution of a supplied acetic acid solution.

- The five (5) solutions will form a series in which each successive solution will be more dilute (less concentrated) than the previous solution.
- The pH of each of the five acetic acid solutions will be measured and recorded.

D. pH Determination of an Acetic Acid Solution

D-1. Clean and rinse a **small** beaker.

D-2. A solution of acetic acid of known molarity (about 0.8 to 1.0 M) will be supplied. This solution will be referred to as **Solution 1**. Record the given molarity of **Solution 1** as M_1 in the **Acetic Acid Label Molarities Table** in the **REPORT PAGES** Section on page 212.

D-3. If possible, immediately measure the pH of **Solution 1** using the pH meter. Use a **10 mL** sample of **Solution 1** to rinse the small beaker. Discard the rinse portion of the solution.

D-4. Use a second, fresh **25 mL** sample of **Solution 1** to rinse the pH electrodes. Discard the second rinse sample.

D-5. Take a third, fresh **25 mL** sample of **Solution 1** for the pH measurement. Gently swirl the solution around the electrodes to remove any air bubbles.

D-6. Turn the meter to **On or pH**. Read and record the value observed within 15 seconds in the **Acetic Acid pH Data Table** in the **REPORT PAGES** Section on page 213.

Record the value of the pH to 2 places after the decimal point.

Return the meter to **Standby** and rinse the electrodes, leaving them in distilled water unless you have further measurements to be made.

D-7. Discard the portion of the solution that was tested **but keep the beaker for use in Part E without cleaning it.**

If time allows, make all pH measurements twice with repeat portions of the solutions, and use mean values in the calculations.

E. Dilution of an Acetic Acid Solution and pH Determination

- E-1. Clean and rinse a **10.00 mL** transfer pipet and a **200.0 mL** volumetric flask and stopper. Dry the outside of the pipet and the flask.
- E-2. Take about **40 mL** of fresh, uncontaminated **Solution 1** into the small beaker already used for **Part D**. Use the beaker volume markings.

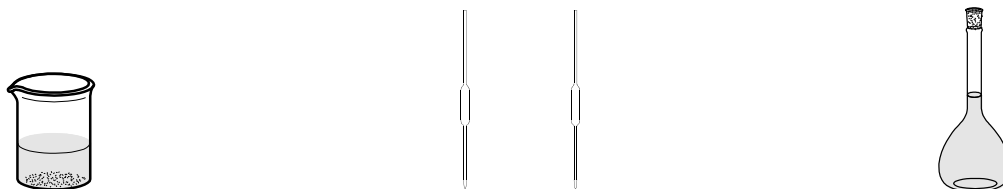
Be careful not to contaminate or dilute this solution.

From Solution 1 four (4) diluted acetic acid solutions will be prepared by serial and multiple dilution.

- The instructions assume that a 10.00 mL pipet and a 200.0 mL volumetric flask will be used.
- The volumes used in the calculations must match the glassware.

- E-3. Make a **1 to 10** dilution of **Solution 1** with distilled water to prepare a new solution in the **200.0 mL** volumetric flask. This solution will be called **Solution 2**.

Mix **Solution 2** completely and uniformly. Invert slowly how many times?



Undiluted and Uncontaminated
Sample of Solution 1

Two 10.00 mL Portions Added by Transfer
Pipet to a Clean 200.0 mL Volumetric Flask³

Solution 2 Uniformly Mixed in a
200.0 mL Volumetric Flask

- E-4. Measure the pH value of **Solution 2** immediately and record the pH values in the **Acetic Acid pH Data Table**. Otherwise label and set aside some of **Solution 2** and measure the pH later. The volumes used in this dilution have been entered into the **Acetic Acid Label Molarities Table** in the **REPORT PAGES** section on page 22 (**Solution 2** row).

Do not waste any of **Solution 2**; this is needed to prepare **Solutions 3, 4** and **5**.

³ **Note:** The preferred method for this dilution and those following would be to use a 20 mL and a 30 mL transfer pipet where appropriate instead of making multiple deliveries with a 10 mL pipet. However, for practice in using a pipet and to save the cost of the additional pipets, you are to use only a 10 mL pipet.

F. Multiple Dilutions of an Acetic Acid Solution and pH Determinations

F-1. **Be careful to conserve Solution 2.** Prepare a **small** beaker and a **10.00 mL** transfer pipet by rinsing several times with **small volumes** of **Solution 2**.

F-2. From **Solution 2** prepare three other solutions by multiple dilution, as follows:

Solution 3 (30.00 mL to 200.0 mL);


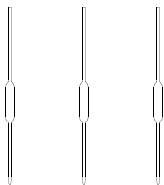

Solution 4 (20.00 mL to 200.0 mL);

Solution 5 (10.00 mL to 200.0 mL).


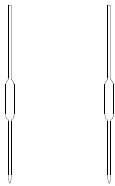

The volumes used in the preparations of these solutions have been entered into the **Acetic Acid Label Molarities Table**.

Mix each solution **thoroughly** and **completely** after preparing it.




Preparation of Solution 3

		
Undiluted and Uncontaminated Sample of Solution 2	3×10.00 mL Portions Added by Transfer Pipet to a Clean 200.0 mL Volumetric Flask	Solution 3 Uniformly Mixed in a 200.0 mL Volumetric Flask

Preparation of Solution 4

		
Undiluted and Uncontaminated Sample of Solution 2	2×10.00 mL Portions Added by Transfer Pipet to a Clean 200.0 mL Volumetric Flask	Solution 4 Uniformly Mixed in a 200.0 mL Volumetric Flask

Preparation of Solution 5

		
Undiluted and Uncontaminated Sample of Solution 2	1×10.00 mL Portion Added by Transfer Pipet to a Clean 200.0 mL Volumetric Flask	Solution 5 Uniformly Mixed in a 200.0 mL Volumetric Flask

F-3. Follow the technique of **Steps D-3 to D-6** to read and record the pH of each of **Solutions 3, 4, and 5**. Measure the pH of each solution as you make it, or after storage (labeled).

F-4. Record the pH values to 2 places after the decimal point in all cases in the **Acetic Acid pH Data Table** in the **REPORT PAGES** Section on page 213. Have the instructor initial your table.

G. Acetic Acid Solution Calculations

These calculated values are required for your report. Show a sample set of calculations for **Solution 3** in the spaces allowed in the **REPORT PAGES** Section.

- G-1. Calculate the **label concentration** of the acetic acid in each diluted solution using the dilution equation. Enter the calculated concentration values in the **Acetic Acid Label Molarities Table** in the **REPORT PAGES** Section on page 212 as M_{dil} in each case. State each concentration value to 3 significant figures.
- The value of M_{conc} for the preparation of **Solution 2** is the given concentration of **Solution 1** ($= M_1$).
 - The value of M_{conc} for the preparation of **Solutions 3, 4 and 5** is the calculated concentration of **Solution 2** ($= M_2$).
- G-2. Transcribe (copy) the five molarity values of the five solutions into the **Acetic Acid Calculated Values Table** in the **REPORT PAGES** Section on page 213.
- G-3. If you have measured the pH of each solution twice, calculate a mean pH value for each solution. Enter the values in the **Acetic Acid pH Data Table** in the **REPORT PAGES** Section on page 23. State each pH value to 2 places after the decimal point.
- G-4. Transcribe (copy) the five mean pH values of the five solutions into the **Acetic Acid Calculated Values Table** in the **REPORT PAGES** Section on page 213.
- G-5. For each of the five solutions, calculate the equilibrium value of the hydrogen ion concentration ($[H^+]$) corresponding to the observed mean pH value. Enter the five calculated values in the **Table**. State each concentration value to 3 significant figures.
- G-6. For each of the five solutions, calculate the equilibrium value of the acetate ion concentration ($[CH_3COO^-]$). Enter the five calculated values in the **Table**. State each concentration value to 3 significant figures.
- G-7. For each of the five solutions, calculate the equilibrium value of the acetic acid concentration ($[CH_3COOH]$). Enter the five calculated values in the **Table**. State each concentration value to 3 significant figures.
- G-8. For each of the five solutions, calculate the value of the equilibrium constant K_a . Enter all of the values calculated into the **Table**. State each value to 3 significant figures.
- G-9. For each of the five solutions, calculate the value of the **% Ionization**. Enter all of the values calculated into the **Table**. State each value to 2 significant figures.
- G-10. Calculate the mean value of the experimental equilibrium constant K_a to 3 significant figures and the range of the K_a values. Record these values below the **Table**.

Name		Day		Start Time	
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REPORT PAGES for Experiment 8

This section will be handed in as your report for this experiment.

Indicator Substances **Instructor's Initials:** _____

Buffer pH Indicator Solution	2	3	4	5	6	7	8	9	10	11	12
Methyl Orange											
Bromophenol Blue											
Litmus											
Phenolphthalein											
Thymol Blue											
Universal Indicator											
Phenolphthalein In Conc. H ₂ SO ₄	(Step A-6. See Box on Page 3 for an Explanation.)										

(marked under data recording rubric)

REPORT PAGES (Cont.)

Question 1: Suppose that you made a mixed indicator solution using **Methyl Orange**, **Litmus** and **Phenolphthalein** indicators. Fill in the Table below. What colour would this mixed indicator display: (i) at pH 3? (ii) at pH 6? (iii) at pH 9? (4 marks)

Indicator	Colour at pH 3	Colour at pH 6	Colour at pH 9
Methyl Orange			
Litmus			
Phenolphthalein			
Predicted Colour of a Mixed Indicator			

Explain your thinking process. (1 mark)

Buffer Solutions (pH Meter Readings in Ink to 2 Places After the Decimal Point)

Sample	pH Paper (Colour Observation)	pH Paper (Approx. pH Conclusion)	pH Meter Reading (In Ink)
Buffer Solution pH 4			
Buffer Solution pH 9			

(part of data recording rubric)

Instructor's Initials: _____

Question 2: What is the purpose of rinsing the electrodes of the pH meter with a portion of the test solution before making the reading with a second portion? (2 marks)

The REPORT PAGES Section Continues on the Next Page →

Name		Day		Start Time	
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REPORT PAGES (Cont.)

Question 3: If the pH paper and the pH meter are not in agreement, which would you consider more correct? Explain your answer. (3 marks)

Common Substances (pH Meter Readings in Ink to 2 Places After the Decimal Point)

Substance	pH Paper Colour	pH Conclusion	pH Meter Reading	pH Active Ingredient
Lemon Juice				Citric Acid
Apple Juice				Malic Acid
Vinegar				Acetic Acid
Purple Grape Juice				Tartaric Acid
Cola				Phosphoric Acid
7-UP™ / Sprite™				Citric Acid
Milk of Magnesia				Magnesium Hydroxide
Tomato Juice				Unknown
Windex				Ammonia
Spic and Span / Mr Clean				Sodium Hydroxide or other Alkaline Ingredients
Laboratory "SOAP" Solution				Alkaline Ingredients

(part of data recording rubric)

Instructor's Initials: _____

The REPORT PAGES Section Continues on the Next Page →

REPORT PAGES (Cont.)

Question 4: “Using universal pH paper to test the pH of purple grape juice is unlikely to be successful.” Explain this statement. (3 marks)

Acetic Acid Dilutions

Molarity of Acetic Acid Solution 1 = M_1 = _____ mol / L

Acetic Acid Label Molarities (State Concentration Values to 3 Significant Figures)

Values Preparation of:	Label Molarity M_{conc} (mol / L)	V_{conc} (L)	V_{dil} (L)	Label Molarity M_{dil} (mol / L)
Solution 2	$M_1 =$ _____	0.0200 L	0.200 L	$M_2 =$ _____
Solution 3	$M_2 =$ _____	0.0300 L	0.200 L	$M_3 =$ _____
Solution 4	$M_2 =$ _____	0.0200 L	0.200 L	$M_4 =$ _____
Solution 5	$M_2 =$ _____	0.0100 L	0.200 L	$M_5 =$ _____

(4 marks)

Calculate the Label Molarity (mol / L) of **Solution 3 (M_3)**.

State the answer to 3 significant figures. Show your work. (2 marks)

Label Molarity of Solution 3 (M_3): $[\text{CH}_3\text{COOH}]$ (mol / L) = _____ mol / L

The REPORT PAGES Section Continues on the Next Page →

Name		Day		Start Time	
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REPORT PAGES (Cont.)

Partner's Name: _____

Acetic Acid pH Data Table (pH Meter Readings in Ink to 2 Places After the Decimal Point)

	Solution 1	Solution 2	Solution 3	Solution 4	Solution 5
Part	D	E	F	F	F
Trial 1					
Trial 2					
Mean					

(2 marks)

Instructor's Initials: _____

Acetic Acid Calculated Values (State Concentration Values to 3 Significant Figures)

	Solution 1	Solution 2	Solution 3	Solution 4	Solution 5
Label Molarity (mol / L)	$M_1 =$ _____	$M_2 =$ _____	$M_3 =$ _____	$M_4 =$ _____	$M_5 =$ _____
Mean pH Value					
$[H^+]$ (at equil.)					
$[CH_3COO^-]$ (at equil.)					
$[CH_3COOH]$ (at equil.)					
K_a					
% Ion.					

(10 marks)

Mean $K_a =$ _____ Range = _____

(2 marks)

The REPORT PAGES Section Continues on the Next Page →

REPORT PAGES (Cont.)**Sample Calculations for Solution 3**

Calculate the Hydrogen ion Concentration ($[\text{H}^+]$) at Equilibrium (mol / L).

State the answer to 3 significant figures. Show your work. (1 mark)

Mean value of pH = _____

Equilibrium Value of $[\text{H}^+]$ (mol / L) = _____ mol / L

Calculate the Acetate Ion Concentration ($[\text{CH}_3\text{COO}^-]$) at Equilibrium (mol / L).

State the answer to 3 significant figures. Show your work. (1 mark)

Equilibrium Value of $[\text{CH}_3\text{COO}^-]$ (mol / L) = _____ mol / L

Calculate the Acetic Acid Concentration ($[\text{CH}_3\text{COOH}]$) at Equilibrium of (mol / L).

State the answer to 3 significant figures. Show your work. (1 mark)

Equilibrium Value of $[\text{CH}_3\text{COOH}]$ (mol / L) = _____ mol/L

State the K_a Expression and Calculate the Value of K_a .

State the answer to 3 significant figures. Show your work. (2 marks)

Value of K_a = _____

The REPORT PAGES Section Continues on the Next Page →

Name		Day		Start Time	
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REPORT PAGES (Cont.)

State the Percent Ionization Expression and Calculate % Ion. Value.

State the answer to 2 significant figures. Show your work. (3 marks)

% Ion. = _____ %

Question 5: Do you predict that the **pH** values will go down (more acidic) or up (less acidic) from **Solution 1** to **Solution 5** as the acetic acid solutions become more dilute? **Explain the reasoning behind your answer.** Is this what you observed in your measured values? (3 marks)

Question 6: What happened to the values of your calculated **equilibrium constants (K_a)** from **Solution 1** to **Solution 5** as the acetic acid solutions become more dilute? K_a equilibrium constants should be constant, give a reason why they may not have been. (3 marks)

Question 7: What happened to the values of your calculated **percent ionization (% Ion.)** from **Solution 1** to **Solution 5** as the acetic acid solutions become more dilute? **Explain the reasoning behind your answer.** (3 marks)