

# 1. Acids, Bases, and Buffered Systems

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## Objective

In this experiment, a study of the properties of acidic and basic substances will be made, using indicators and a pH meter to determine pH. A buffered system will be prepared by half-neutralization, and the properties of the buffered system will be compared to those of a nonbuffered medium.

## Introduction

One of the most important properties of aqueous solutions is the concentration of hydrogen (hydronium) ion present. The  $H^+$  [ $H_3O^+$ ,  $H_2O \cdot H^+$ ] ion greatly affects the solubility of many organic and inorganic substances; the nature of complex metal cations found in solution; and the rate of many chemical reactions; and may profoundly affect the reactions of biochemical species.

## The pH Scale

Typically the concentration of  $H^+$  ion in aqueous solutions may be small, requiring the use of scientific notation to describe the concentration numerically. For example, an acidic solution may have  $[H^+] = 2.3 \times 10^{-5} M$ , whereas a basic solution may have  $[H^+] = 4.1 \times 10^{-10} M$ . Because scientific notation may be difficult to deal with, especially when you make *comparisons* of numbers, a mathematical simplification, called pH, has been defined to describe the concentration of hydrogen ion in aqueous solutions:

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

By use of a logarithm, the power of ten of the scientific notation is converted to a "regular" number, and use of the minus sign in the definition of pH produces a positive value for the pH. For example, if  $[H^+] = 1.0 \times 10^{-4} M$ , then

$$\text{pH} = -\log[1.0 \times 10^{-4}] = -[-4.00] = 4.00$$

and similarly, if  $[H^+] = 5.0 \times 10^{-2} M$ , then

$$\text{pH} = -\log[5.0 \times 10^{-2}] = -[-1.30] = 1.30$$

Although basic solutions are usually considered to be solutions of *hydroxide* ion,  $OH^-$ , basic solutions also must contain a certain concentration of *hydrogen* ion,  $H^+$ , because of the equilibrium that exists in aqueous solutions due to the autoionization of water. The concentration of hydrogen ion in a basic solution can be determined by reference to the equilibrium constant for the autoionization of water,  $K_w$ :

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \quad (\text{at } 25^\circ\text{C})$$

For example, if a basic solution contains  $[\text{OH}^-]$  at the level of  $2.0 \times 10^{-3} M$ , then the solution also contains hydrogen ion at the level of

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

$$[\text{H}^+] = 1.0 \times 10^{-14}/[2.0 \times 10^{-3}] = 5 \times 10^{-12} M$$

In pure water, there must be *equal* concentrations of hydrogen ion and hydroxide ion (one of each ion is produced when a water molecule ionizes). So, in pure water

$$[\text{H}^+] = [\text{OH}^-] = (1.0 \times 10^{-14})^{1/2} = 1.0 \times 10^{-7} M$$

The pH of pure water is thus 7.00; this value serves as the *dividing line* for the pHs of aqueous solutions. A solution in which there is *more* hydrogen ion than hydroxide ion ( $[\text{H}^+] > 1.0 \times 10^{-7} M$ ) will be *acidic* and will have pH less than 7.00. Conversely, a solution with *more* hydroxide ion than hydrogen ion ( $[\text{OH}^-] > 1.0 \times 10^{-7} M$ ) will be *basic* and will have pH greater than 7.00.

If the pH of a solution is known (or has been measured experimentally), then the hydrogen ion concentration of the solution may be calculated

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

For example, a solution with  $\text{pH} = 4.20$  would have hydrogen ion concentration given by

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-4.20} = 6.3 \times 10^{-5} M$$

### ***Experimental Determination of pH***

The experimental determination of the pH of a solution is commonly performed by either of two methods. The first of these methods involves the use of chemical dyes called **indicators**. These substances are generally weak acids/bases and can exist in either of two colored forms, depending on whether or not the molecule is protonated or has been deprotonated. For example, let HIn represent the protonated form of the indicator. In aqueous solution, an equilibrium exists:



<i>first</i>	<i>second</i>
<i>color</i>	<i>color</i>

Depending on what other acidic/basic substances are present in the solution, the equilibrium of the indicator will shift, and one or the other colored form of the indicator will predominate and impart its color to the entire solution.

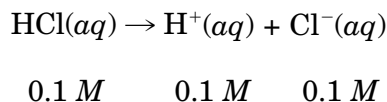
Indicators commonly change color over a relatively short pH range (about 2 pH units), and, when properly chosen, can be used to estimate the pH of a solution. Among the most common indicators are litmus (which is usually used in the form of test paper that has been impregnated with the indicator) and

phenolphthalein (which is most commonly used as a solution added to samples for titration analysis). Litmus changes color from red to blue as the pH of a solution increases from pH 6 to pH 8. Phenolphthalein goes from a colorless form to red as the pH of a solution increases from pH 8 to pH 10. A given indicator is useful for determining pHs only in the region in which it changes color. Indicators are available for measurement of pH in all the important ranges of acidity and basicity. By matching the color of a suitable indicator in a solution of known pH to the color of the indicator in an unknown solution, the pH of the unknown solution can be estimated.

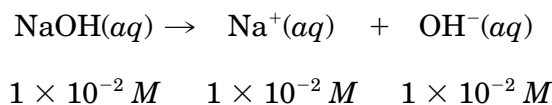
The second method for the determination of pH involves an instrument called a **pH meter**. The pH meter contains two electrodes, one of which is sensitive to the concentration of hydrogen ion. Typically the pH sensitive electrode is a *glass membrane* electrode, whereas the second electrode may consist of a silver/silver chloride system, or a calomel electrode. Frequently, a glass membrane pH electrode is combined in the same physical chamber as a silver/silver chloride electrode and is referred to as a “combination” pH electrode. In the solution being measured, the electrical potential between the two electrodes is a function of the hydrogen ion concentration of the solution. The pH meter has been designed so that the pH of the solution can be directly read from the scale provided on the face of the meter. It is essential always to calibrate a pH meter before use: a solution of known pH is measured with the electrodes, and the display of the meter is adjusted to read the correct pH for the known pH solution. A properly calibrated pH meter provides for much more precise determinations of pH than does the indicator method, and is ordinarily used when a very accurate determination of pH is needed.

### ***Strong and Weak Acids/Bases***

Some acids and bases undergo substantial ionization when dissolved in water and are called **strong acids** or **strong bases**. Strong acids and bases are completely ionized in dilute aqueous solutions. Other acids and bases, because they ionize *incompletely* in water (often to the extent of only a few percent), are called **weak acids** or **weak bases**. Hydrochloric acid (HCl) and sodium hydroxide (NaOH) are examples of a typical strong acid and a strong base, respectively. A 0.1 *M* solution of HCl contains hydrogen ion concentration at effectively 0.1 *M* concentration because of the complete ionization of HCl molecules.

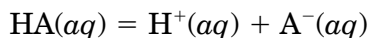


Similarly, a  $1 \times 10^{-2} \text{ M}$  solution of NaOH contains hydroxide ion concentration at the level of  $1 \times 10^{-2} \text{ M}$  because of the complete ionization of NaOH.



Acetic acid (HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> or CH<sub>3</sub>COOH) and ammonia (NH<sub>3</sub>) are typical examples of a weak acid and weak base, respectively. Weak acids and weak bases must be treated by the techniques of chemical equilibrium in determining the

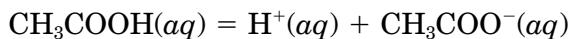
concentrations of hydrogen ion or hydroxide ion in their solutions. For example, for the general weak acid HA, the equilibrium reaction would be



and the equilibrium constant expression would be given by

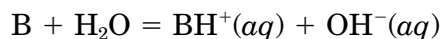
$$K_a = [\text{H}^+][\text{A}^-]/[\text{HA}]$$

For acetic acid, as an example, the ionization equilibrium reaction and  $K_a$  are represented by the following:



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

$K_a$  is a constant and is characteristic of the acid HA. For a given weak acid, the product of the concentrations in the expression will remain constant at equilibrium, regardless of the manner in which the solution of the acid was prepared. For a general weak base B, the equilibrium reaction would be



and the equilibrium constant expression would be given by

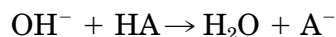
$$K_b = [\text{BH}^+][\text{OH}^-]/[\text{B}]$$

As an example, the ionization and  $K_b$  for the weak base ammonia are represented by



$$K_b = [\text{NH}_4^+][\text{OH}^-]/[\text{NH}_3]$$

The value of the equilibrium constant  $K_a$  for a weak acid (or  $K_b$  for a weak base) can be determined experimentally by several methods. One simple procedure involves little calculation, is accurate, and does not require knowledge of the actual concentration of the weak acid (or base) under study. For example, a sample of weak acid (HA) is dissolved in water and is then divided into two equal-volume portions. One portion is then titrated with a sodium hydroxide solution to a phenolphthalein endpoint, thereby converting all HA molecules present into  $\text{A}^-$  ions:



The number of moles of  $\text{A}^-$  ion produced by the titration is equal, of course, to the number of moles of HA in the original half-portion titrated, and is also equal to the number of moles of HA in the remaining unused portion of weak acid. The two portions are then mixed, and the pH of the combined solution is measured. Because  $[\text{HA}] = [\text{A}^-]$  in the combined solution, these terms (represented below by  $[n]$ ) cancel each other in the equilibrium constant expression for  $K_a$  for the acid

$$K_a = [\text{H}^+][\text{A}^-]/[\text{HA}] = [\text{H}^+][n]/[n]$$

and therefore the ionization equilibrium constant,  $K_a$ , for the weak acid will be given directly from a measurement of the hydrogen ion concentration

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

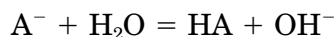
for the combined solution. In other words, the ionization equilibrium constant for a weak acid can be determined simply by measuring the pH of a half-neutralized sample of the acid.

### ***pH of Salt Solutions***

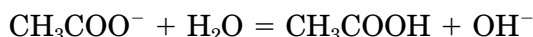
Salts that can be considered to have been formed from the complete neutralization of strong acids with strong bases—such as NaCl (which can be considered to have been produced by the neutralization of HCl with NaOH)—ionize completely in solution. Such strong acid/strong base salts do not react with water molecules to any appreciable degree (do not undergo hydrolysis) when they are dissolved in water. Solutions of such salts are neutral and have  $\text{pH} = 7$ . Other examples of such salts are KBr (from HBr and KOH) and  $\text{NaNO}_3$  (from  $\text{HNO}_3$  and NaOH).

However, when salts formed by the neutralization of a *weak* acid or base are dissolved in water, these salts furnish ions that tend to react to some extent with water, producing molecules of the weak acid or base and releasing some  $\text{H}^+$  or  $\text{OH}^-$  ion to the solution. Solutions of such salts will not be neutral, but rather will be acidic or basic themselves.

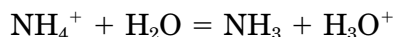
Consider the weak acid HA. If the sodium salt of this acid,  $\text{Na}^+\text{A}^-$  is dissolved in water, the  $\text{A}^-$  ions released to the solution will react with water molecules to some extent



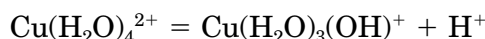
The solution of the salt will be *basic* because hydroxide ion has been released by the reaction of the  $\text{A}^-$  ions with water. For example, a solution of sodium acetate (a salt of the weak acid acetic acid) is basic because of reaction of the acetate ions with water molecules, releasing hydroxide ions:



Conversely, solutions of salts of weak bases (such as  $\text{NH}_4^+\text{Cl}^-$ , derived from the weak base  $\text{NH}_3$ ) will be *acidic*, because of reaction of the ions of the salt with water. For example,



Most salts of transition metal ions are acidic. A solution of  $\text{CuSO}_4$  or  $\text{FeCl}_3$  will typically be of pH 5 or lower. The salts are completely ionized in solution. The acidity comes from the fact that the metal cation is hydrated. For example,  $\text{Cu}(\text{H}_2\text{O})_4^{2+}$  better represents the state of a copper(II) ion in aqueous solution. The large positive charge on the metal cation attracts electrons from the O–H bonds in the water molecules, thereby weakening the bonds and producing some  $\text{H}^+$  ions in the solution



## ***Buffered Solutions***

Some solutions, called buffered solutions, are remarkably resistant to pH changes caused by the addition of an acid or base from an outside source. A **buffered solution** in general is a mixture of a weak acid and a weak base. The most common sort of buffered solution consists of a mixture of a conjugate acid/base pair. For example, a mixture of the salt of a weak acid or base, and the weak acid or base itself would constitute a buffered solution. The half-neutralized solution of a weak acid described earlier as a means of determining  $K_a$  would therefore represent a buffered system. That solution contained equal amounts of weak acid HA and of the anion  $A^-$  present in its salt. If a small amount of strong acid were added to the buffered system from an external source, the  $H^+$  ion introduced by the strong acid would tend to react with the  $A^-$  ion of the salt, thereby preventing a change in pH. Similarly, if a strong base were added from an external source to the HA/ $A^-$  system, the  $OH^-$  introduced by the strong base would tend to be neutralized by the HA present in the buffered system. If similar small amounts of strong acid or strong base were added to a nonbuffered system (such as plain distilled water), the pH would be changed drastically by the addition.

### **SAFETY PRECAUTIONS**

- **Wear safety glasses at all times while in the laboratory.**
- **The acids, bases, and salts to be used are in fairly dilute solution, but may be irritating to the skin. Wash if they are spilled and inform the instructor.**
- **The pH meter uses electricity. Do not spill water or a solution near the meter; electrical shock may result.**

### **Apparatus/Reagents Required**

pH meter and electrodes, buret and clamp, unknown sample for indicator determination, known pH 0–6 samples, indicators, unknown acid sample (for half-neutralization and buffered solution study), 0.2 M NaOH solution, 0.1 M HCl, 0.1 M NaOH, salt solutions for pH determination [NaCl, KBr,  $NaNO_3$ ,  $K_2SO_4$ ,  $NaC_2H_3O_2$ ,  $NH_4Cl$ ,  $K_2CO_3$ ,  $(NH_4)_2SO_4$ ,  $CuSO_4$ , and  $FeCl_3$ ], Universal indicator solution

### **Procedure**

Record all data and observations directly in your notebook in ink.

#### ***A. pH Using Indicators***

Obtain a solution of unknown pH for indicator determination from your instructor. Place about half an inch of this solution in a small test tube, and add 1–2 drops of one of the indicators in the table of indicators that follows. Record the color of the solution after addition of the indicator.

<b>Indicator</b>	<b>Color Change</b>	<b>pH of Color Change</b>
methyl violet	yellow to violet	-1 to 1.7
malachite green	yellow to green	0.2 to 1.8
cresol red	red to yellow	1 to 2
thymol blue	red to yellow	1.2 to 2.8
bromphenol blue	yellow to blue	3 to 4.7
methyl orange	red to yellow	3.2 to 4.4
bromcresol green	yellow to blue	3.8 to 5.4
methyl red	red to yellow	4.4 to 6.0
methyl purple	purple to green	4.5 to 5.1
bromthymol blue	yellow to blue	6.0 to 7.6
litmus	red to blue	4.7 to 6.8

Repeat the test with each of the indicators listed, using a small amount of the unknown solution and 1–2 drops of indicator.

From the information given in the table, estimate the pH of your unknown solution. Note that the color of an indicator is most indicative of the pH in the region in which the indicator changes color.

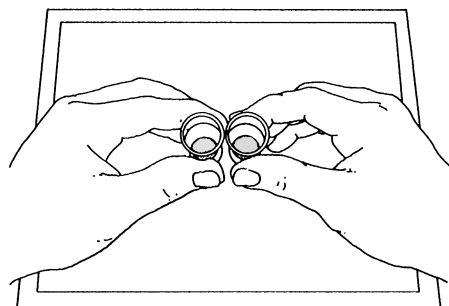
Having established the pH of your unknown to within approximately one pH unit, obtain a known solution with a pH about equal to that of your unknown. Stock solutions of standard reference buffered solutions with integral pH values from 0 to 6 have been prepared for you.

Test 1-mL samples of the known pH sample with the indicators most useful in determining the approximate pH of the unknown sample. Compare the colors of the known and unknown solutions. On the basis of the comparisons, decide whether the pH of the unknown is slightly higher or lower than the pH of the known sample.

Select a second known solution, with the aim of bracketing the pH of your unknown solution between two known solutions. Treat the second known solution with the useful indicator(s) and compare the color(s) with the unknown sample(s). Continue testing known pH solutions until you find two that differ by one pH unit, one of which has a higher pH than the unknown, and one of which has a lower pH than the unknown.

For a final estimation of the pH of your unknown, use a 5-mL sample of the unknown and of the two known solutions that bracket its pH. Add 2 drops of the indicator to each solution. View the solutions against a well-lighted sheet of white paper, and estimate the pH of your unknown (as compared to the two known solutions) to  $\pm 0.3$  pH units. (See Figure 1-1.)

FIGURE 1-1  
Comparison  
of known and  
unknown  
solution  
colors. Make  
certain the  
background is  
well-lighted.

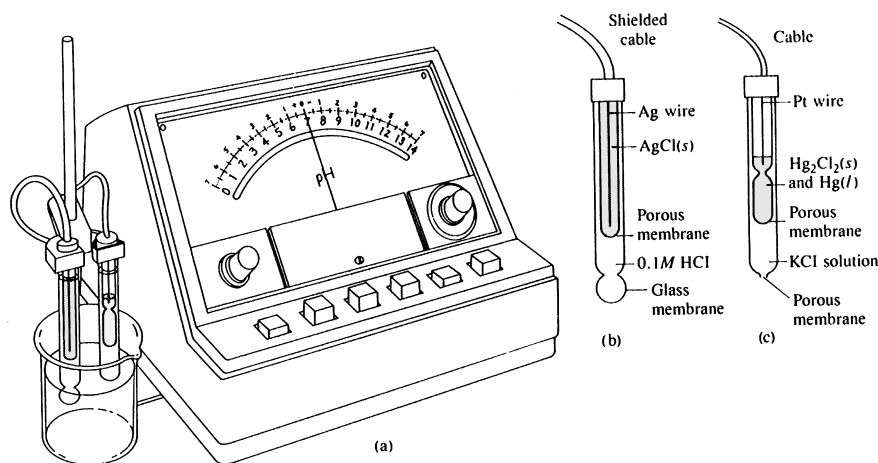


## B. Calibration of the pH Meter

Typically, several models of pH meters are available in the laboratory (see Figure 1-2). Your instructor will provide specific instructions for the operation of the meter you will use.

FIGURE 1-2

(a) A typical pH meter/ electrode setup. The two electrodes may be combined in one container to form a combination pH-sensing electrode. (b) pH sensing electrode. The glass membrane measures the difference in  $[H^+]$  between the interior of the electrode and the solution being tested. (c) Calomel reference electrode. This completes the electrical circuit in the solution and is very stable and reproducible.



In all cases, the meters must be *calibrated* before being used for pH measurements. Samples of pH standard reference buffered solutions are available for calibrating the meters. Generally, the combination electrode is dipped into one of the reference buffered solutions and allowed to stand for several minutes; this permits the electrode to come to equilibrium with the buffered solution. The set or calibrate knob on the face of the meter is then adjusted until the meter display reads the correct pH for the reference buffered solution. For precise work, a two-point calibration, in which the meter is checked in two different reference buffers, may be necessary. Your instructor will explain the two-point calibration method if you will be using it.

Remember that the pH sensing electrode is made of *glass* and is therefore very *fragile*. (Combination pH electrodes are very expensive, and you may be asked to pay for the electrode if you break it.) Handle the electrode gently, do not stir solutions with the electrode, and keep the electrode in a beaker of distilled water when not in use to keep it from drying out. Rinse the electrode with distilled water when transferring it from one solution to another.

## C. pH of Salt Solutions

Note: If the number of pH meters in the laboratory is limited, your instructor may ask you to perform this qualitative portion of the experiment using *Universal indicator*; rather than the pH meter. Universal indicator is a dye that exhibits different colors over the entire pH range. Universal indicator does not permit as sensitive a measurement of pH as would be possible with a pH meter or with the method using several indicators as discussed in Part A of this experiment. For a quick “ballpark” estimate of pH, however, Universal indicator can be very useful.

In this case, obtain 10 drops of each salt solution, and test for pH with one drop of Universal indicator. Use the color chart provided with the indicator to estimate the pH of each salt solution. Write an equation accounting for the pH observed.

If sufficient pH meters are available in the lab, obtain 20–25 mL (in a small beaker) of one of the 0.1 *M* salt solutions listed in the table that follows. Rinse the pH electrode with distilled water and immerse the electrode in the salt solution.

<b>Salt</b>	<b>Formula</b>
sodium chloride	NaCl
potassium bromide	KBr
sodium nitrate	NaNO <sub>3</sub>
potassium sulfate	K <sub>2</sub> SO <sub>4</sub>
sodium acetate	NaC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>
ammonium chloride	NH <sub>4</sub> Cl
potassium carbonate	K <sub>2</sub> CO <sub>3</sub>
ammonium sulfate	(NH <sub>4</sub> ) <sub>2</sub> SO <sub>4</sub>
copper(II) sulfate	CuSO <sub>4</sub>
iron(III) chloride	FeCl <sub>3</sub>

Allow the electrode to stand in the solution for 2–3 minutes so that it may come to equilibrium with the solution; then record the pH of the salt solution.

Rinse the pH electrode with distilled water, and determine the pH of each of the other salt solutions listed in the table.

Based on the pH measured for each salt solution, write a chemical equation that will explain the observed pH.

#### ***D. Determination of $K_a$ for a Weak Acid***

Obtain a sample of an unknown solid acid for  $K_a$  determination.

If the acid is provided as a solution, use your graduated cylinder to measure out two 50-mL portions into 250-mL Erlenmeyer flasks. If the acid is provided as a solid, place the solid acid sample in 100 mL of distilled water, stir thoroughly to dissolve the acid, and use your graduated cylinder to divide the acid solution into two equal volume portions and place into two 250-mL Erlenmeyer flasks.

Clean, rinse, and fill a buret with 0.2 *M* NaOH solution.

Take one portion of your unknown acid solution, add 3–4 drops of phenolphthalein indicator, and titrate the acid solution until a faint pink color appears and persists. At this point, you have neutralized half the original acid provided and have a solution that contains the salt of the weak acid. It is not necessary in this experiment to determine the exact volume necessary for the titration.

Mix the neutralized solution from the titration with the remaining portion of weak acid solution. Stir well.

Check the calibration of the pH meter with the standard reference buffers provided to make sure that the meter has not drifted in its readings.

Rinse off the electrode of the pH meter with distilled water, and transfer the electrode to the mixture of unknown weak acid and salt. Allow the electrode to stand for 2–3 minutes, and record the pH of the solution.

From the observed pH, calculate  $[H^+]$  in the solution and  $K_a$  for the weak acid unknown. Save the solution for use later.

### ***E. Properties of a Buffered Solution***

The half-neutralized solution prepared in Part D is a buffered system since it contains approximately equal amounts of a free weak acid (HA) and of a salt of the weak acid ( $Na^+A^-$ ) from the titration. This buffered system should be very resistant to changes in pH, compared to a nonbuffered system, when small quantities of strong acid/strong base are added from an external source.

Check the calibration of the pH meter with the standard reference buffers provided to make sure that the meter has not drifted in its readings.

Place 25 mL of the unknown acid buffered mixture in a small beaker and measure its pH with the meter. Add 5 drops of 0.1 M strong acid HCl to the buffered mixture and stir. Allow the electrode to stand for 2–3 minutes and record the pH. The pH should only drop by a small fraction of a pH unit (if at all).

Place another 25-mL sample of the buffered mixture in a small beaker. Add 5 drops of 0.1 M strong base NaOH to the buffered mixture and stir. Record the pH after allowing the electrode to stand for 2–3 minutes in the solution. The pH should only have risen by a small fraction of a pH unit (if at all).

For comparison, repeat the addition of small amounts of 0.1 M HCl and 0.1 M NaOH to 25-mL portions of one of the standard reference buffered solutions available. Determine the pH of the buffered solution itself, as well as the pHs after the addition of strong acid or strong base.

To contrast a buffered system with an unbuffered solution, place 25 mL of distilled water in a small beaker, immerse the electrode, and record the pH. (Do not be surprised if the pH is not exactly 7.0—only pure water that is not in contact with air or glass containers gives pH exactly 7.00.) Add 5 drops of 0.1 M HCl to the distilled water, and record the pH. It should decrease by several pH units.

Place another 25 mL of distilled water in a beaker. Add 5 drops of 0.1 M NaOH solution and record the pH. It should increase by several pH units.

Write net ionic equations demonstrating why an  $HA/A^-$  buffered system resists changes in its pH when strong acid ( $H^+$ ) or strong base ( $OH^-$ ) is added.

# *Acids, Bases, and Buffered Systems*

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Date: ..... Student name: .....  
Course: ..... Team members: .....  
Section: .....  
Instructor: .....

## **Prelaboratory Questions**

1. Using your textbook or a handbook of chemical data, list five strong acids and five strong bases. Write an equation for each acid and each base showing its ionization in water.

2. Using a handbook of chemical data, find the ionization equilibrium constants ( $K_a$  or  $K_b$ ) for the following weak acids and weak bases. Write an equation for each acid or base showing its ionization in water, and write the expression in terms of chemical symbols for the ionization equilibrium constant for the weak acid or base.

Hydrocyanic acid, HCN

Hydrofluoric acid, HF

Formic acid, HCOOH

Methylamine, CH<sub>3</sub>NH<sub>2</sub>

Diethylamine, (CH<sub>3</sub>CH<sub>2</sub>)<sub>2</sub>NH

3. Using a handbook or chemical encyclopedia, find five acid/base indicators other than the indicators listed in the *Procedure* for this experiment, and give their names and pH color change range.

# Acids, Bases, and Buffered Systems

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Date: ..... Student name: .....  
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Instructor: .....

## Results/Observations

### A. pH Using Indicators

Identification number of unknown sample .....

Approximate pH of unknown .....

Which indicators bracketed the pH of your unknown?

Lower pH ..... Higher pH .....

Best estimate of unknown's pH .....

### B. Calibration of pH meter

Describe how you calibrated the pH meter for subsequent determinations of pH. Which standard reference buffers did you use?

### C. pH of Salt Solutions

Salt	pH	Equation
NaCl	.....	.....
KBr	.....	.....
NaNO <sub>3</sub>	.....	.....
K <sub>2</sub> SO <sub>4</sub>	.....	.....
NaC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	.....	.....
NH <sub>4</sub> Cl	.....	.....
K <sub>2</sub> CO <sub>3</sub>	.....	.....

$(\text{NH}_4)_2\text{SO}_4$  .....

$\text{CuSO}_4$  .....

$\text{FeCl}_3$  .....

#### D. Determination of $K_a$ for a Weak Acid

Identification number of unknown acid .....

pH of half-neutralized solution of unknown .....

$[\text{H}^+] =$  .....  $K_a =$  .....

#### E. Properties of a Buffered Solution

pH of buffer (half-neutralized acid unknown) before addition of other reagents .....

pH of buffer with 5 drops HCl .....

pH of buffer with 5 drops NaOH .....

pH of buffer (standard reference) before addition of other reagents .....

pH of buffer with 5 drops HCl .....

pH of buffer with 5 drops NaOH .....

pH of distilled water .....

pH of water with 5 drops HCl .....

pH of water with 5 drops NaOH .....

Balanced equations demonstrating buffering action:

#### Questions

1. It was mentioned that you should not expect the pH of distilled water to be 7.00 as ordinarily measured in the laboratory. What component(s) of the atmosphere might cause distilled water in the laboratory to have  $\text{pH} < 7.0$ ? Write an equation to show this.

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2. The human body contains several systems that show buffering capabilities. Why would buffering action be vital in these systems? Use your textbook or a chemical encyclopedia to find some of the buffer systems found in the body. Write equations showing how the components of these buffers maintain pH in the body.
  
  
  
  
  
  
  
  
  
  
3. Calculate the pH of a buffer solution made by adding 10.0 g of anhydrous sodium acetate ( $\text{NaC}_2\text{H}_3\text{O}_2$ ) to 100 mL of 0.100 M acetic acid. Assume there is no change in volume on adding the salt to the acid.  $\text{p}K_a$  for acetic acid is 4.74.
  
  
  
  
  
  
  
  
  
  
4. In addition to pH, another important aspect of a buffer system is called the *buffer capacity*. Use your textbook or a chemical dictionary to define buffer capacity.

