Experiment 12: Buffer Solution

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2026.02.24

I. Objective

The objective of this experiment is to understand buffer solutions and explore the factors that influence their buffer capacity.

II. Technique

In this experiment, you will learn experimental skills such as using a pH meter, volumetric pipette, heating stirrer, and preparing chemical solutions.

III. Introduction

(i) Buffer solution

A buffer solution is a solution whose pH does not change significantly when small amounts of acid or base are added, or when it is diluted with water. Buffer solutions can be relatively concentrated strong acids or strong bases, or mixtures of a weak acid (base) and its conjugate base (acid). For example, 0.2 M hydrochloric acid (HCl) is a good buffer solution because the solution already contains a high concentration of hydrogen ions (H^+); therefore, adding a small amount of acid or base has little effect on its hydrogen-ion concentration or pH value.

A commonly used buffer solution consists of a weak acid (HA) and its conjugate base (A^-), or a weak base (B) and its conjugate acid (BH^+), forming a common-ion system in solution. Because the conjugate acid–base pair maintains a dynamic equilibrium in solution, as shown in Equations 12-1 and 12-2, where K_a represents the acid dissociation constant:

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq) \qquad (12\text{-}1)$$

$$K_a = \frac{[H^+][A^-]}{[HA]} \qquad (12\text{-}2)$$

When a small amount of acid is added to the solution, the basic species (A^-) reacts with the added acid:

$$A^-(aq) + H^+(aq) \rightarrow HA(aq) \qquad (12\text{-}3)$$

When a base is added, the acidic species (HA) in the solution reacts with the added base:

$$\mathrm{HA(aq)} + \mathrm{OH^-(aq)} \rightarrow \mathrm{H_2O(l)} + \mathrm{A^-(aq)} \quad (12\text{-}4)$$

According to the Henderson equation (Equation 12-7):

$$[\mathrm{H^+}] = \frac{K_a[\mathrm{HA}]}{[\mathrm{A^-}]} \quad (12\text{-}5)$$

$$\mathrm{pH} = -\log[\mathrm{H^+}] = -\log K_a - \log\left(\frac{[\mathrm{HA}]}{[\mathrm{A^-}]}\right) \quad (12\text{-}6)$$

$$\mathrm{pH} = \mathrm{p}K_a + \log\left(\frac{[\mathrm{A^-}]}{[\mathrm{HA}]}\right) \quad (12\text{-}7)$$

The pH of a buffer solution is determined by K_a and the concentration ratio of the conjugate acid–base pair ($[A^-]/[HA]$). Therefore, when small amounts of acid or base are added, the change in $\log([A^-]/[HA])$ is minimal, allowing the solution to maintain a nearly constant pH.

Buffer capacity refers to the ability of a buffer solution to resist changes in pH. In this experiment, it is defined as the number of equivalents of strong acid or strong base required to change the pH of one liter of buffer solution by one unit.

The higher the concentration of the buffer solution, the greater its buffer capacity. In addition, the closer the concentration ratio of the conjugate acid–base pair ($[A^-]/[HA]$) is to 1, the greater the buffer capacity; this ratio should not exceed 10. Therefore, an effective buffer solution has a pH within the range of $pK_a \pm 1$ of the weak acid.

In this experiment, a buffer solution is prepared using acetic acid (CH_3COOH, abbreviated as HOAc) and sodium acetate (CH_3COONa, abbreviated as NaOAc). A measured amount of strong acid or strong base is then added to the solution, and the resulting change in pH is determined using a pH meter to compare how concentration affects buffer capacity.

(ii) pH meter

A pH meter is an instrument used to measure the pH of a solution. It consists of three main components. The first is a thermometer probe, which measures the temperature of the solution. The second is a combination pH electrode, which contains a reference electrode (usually a silver/silver chloride electrode with a constant potential) and an indicator electrode (typically a glass electrode whose potential depends on the hydrogen ion concentration of the solution). The third component is a voltmeter that measures the potential difference between the two electrodes.

When the thermometer probe and the pH electrode are immersed in the test solution, the pH meter performs automatic temperature compensation based on the measured temperature and directly converts the measured potential into a pH value. The relationship between the measured potential and the pH of the solution is shown in Equation 12-8:

$$E_m = K - \frac{2.3RT\text{pH}}{nF} \qquad (12\text{-}8)$$

Where:

E_m: measured potential — K: constant (its value depends on the electrode)

R: gas constant — T: absolute temperature of the solution

pH: pH of the solution — n: number of electrons transferred in the electrode reaction

F: Faraday constant

Since R, n, and F are constants during measurement, Equation 12-8 can be simplified as:

$$E_m = mT(\text{pH}) + K \qquad (12\text{-}9)$$

When a set of measurements is carried out at a constant temperature, Equation 12-9 shows that E_m and pH have a linear relationship, with a slope of mT.

Therefore, a pH meter is typically calibrated at a fixed temperature using two standard buffer solutions. The first standard solution (usually pH 7.00) is used to set the absolute reading of the instrument, while the second standard solution (usually pH 4.00) is used to determine the slope of the linear relationship, as illustrated in Figure 12-1.

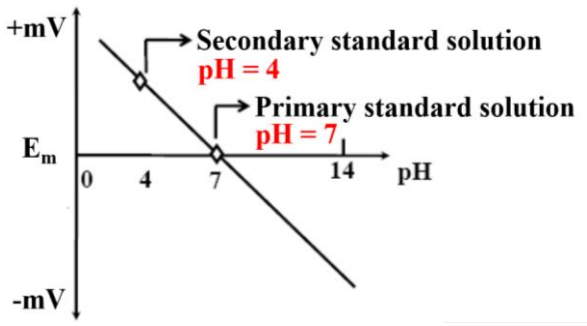


Figure 12-1. Relationship between the measured potential and the pH of the solution.

IV. Equipment

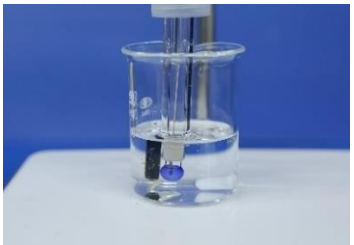
Group Equipment/Materials (in Cabinet)		Provided by Teaching Assistant
500 mL Beaker (1)	Pipet filler (1)	Dropper
250 mL Beaker (2)	Graduated pipet 10 mL (1)	pH meter (1)
100 mL Beaker (2)	Graduated cylinder 10 mL (1)	Cleaning wipes
Glass rod (1)	Wash bottle (1)	(one box shared by two groups)

V. Chemicals

0.05 M CH_3COOH	0.05 M CH_3COONa
1.00 M HCl	1.00 M NaOH
Standard buffer solutions (Primary standard buffer solution, pH = 7.00) (Secondary standard buffer solution, pH = 4.00)	

VI. Procedure

	Procedure	Figure
1.	**Calibration of the pH Meter:** (1) Connect the main power supply and press the "POWER" button to turn on the instrument. Allow it to warm up for 10 minutes. (2) Press the "MODE" button until "°C" appears at the lower right corner of the screen. Confirm and record the solution temperature. Then press the "MODE" button again until "pH" appears at the upper right corner to begin calibration. (3) Press the "HOLD" button to lock the displayed value. Rinse the electrode and thermometer probe with deionized water from a wash bottle, gently blot them dry with tissue paper, then immerse them in the pH 7.00 buffer solution. Press the "HOLD" button again to release the lock.	 

1.	(4) Adjust the “CALIB.” knob until the display reads “7.00.” (5) Rinse the electrode, then place it in the pH 4.00 buffer solution. Adjust the “SLOPE” knob until the display reads “4.00” to complete the calibration. (6) After rinsing the electrode and thermometer probe, immerse them in clean deionized water for storage. **Note:** The pH electrode is fragile and expensive; handle it with care.	 
2.	Prepare 250 mL each of 0.050 M HOAc and 0.050 M NaOAc solutions using 17 M concentrated acetic acid and sodium acetate trihydrate, respectively. Pour the solutions into 250 mL beakers for later use. **Note:** Write the preparation procedure in the pre-lab report.	
3.	Place about 5 mL each of 1.0 M NaOH and 1.0 M HCl solutions into clean, oven-dried test tubes for later use. Use a 10 mL graduated cylinder to measure the volume of 30 drops from the dropper solution in order to calculate the average volume of one drop of NaOH and HCl solution. **Note:** The droppers used for adding NaOH and HCl must be intact and undamaged to avoid large variations in drop volume. Alternatively, use designated plastic droppers.	
4.	According to Table 12-1, use a 10 mL graduated pipette to measure 15 mL each of 0.050 M HOAc and 0.050 M NaOAc into a 50 mL beaker to prepare the test solution (a).	
5.	Set up the apparatus according to Figure 12-2 and measure the initial pH of the test solution. **Note:** Ensure that the salt bridge at the tip of the electrode is immersed in the solution. Adjust the position of the magnetic stir bar to prevent it from hitting the glass electrode, and stir the solution at low speed.	

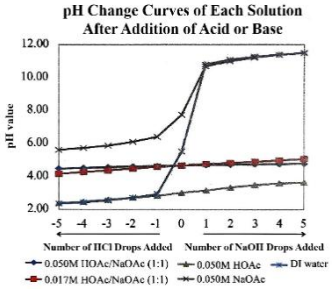
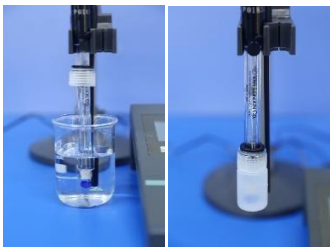
6.	Use a dropper to add one drop of 1.0 M HCl to solution (a) from Step 4. After thorough mixing, read and record the pH value. Repeat the addition for a total of five drops, recording the pH after each addition. **Note:** Keep the dropper in a vertical position while adding the solution to ensure consistent drop volume.	
7.	Measure a fresh portion of solution (a) into a 50 mL beaker. Repeat Steps 5 and 6, but add 1.0 M NaOH instead, recording the pH after each addition of base.	
8.	Repeat Steps 4–7 using the four different formulations (b), (c), (d), and (e) listed in Table 12-1. **Note 1:** Use a 50 mL graduated cylinder to measure the deionized water. **Note 2:** The pH of deionized water is easily affected by trace amounts of acid or base; it is recommended to measure it last.	
9.	(1) The waste solution is an aqueous salt solution; after neutralization, it may be disposed of down the sink. (2) After the experiment, rinse the electrode and immerse it in clean deionized water, or place it in an electrode sleeve containing 3 M KCl. (3) Turn off the pH meter. (4) Return the magnetic stir bar.	

Table 12-1. Volume measurements of test solutions.

Solution	0.050 M HOAc（mL）	0.050 M NaOAc（mL）	DI water（mL）	Total volume（mL）
(a)	15	15	0	30
(b)	5	5	20	
(c)	30	0	0	
(d)	0	30	0	
(e)	0	0	30	

References

1. King, E. J. *Qualitative Analysis and Electrolytic Solutions*; Harcourt, Brace: New York, **1976**.

2. *Advanced General Chemistry C/F 1507 Laboratory Manual*; Columbia Univ., **1991**.

3. 北京大學化學系普通化學教研室，*普通化學實驗*；修訂本，藝軒圖書出版社；台北，**1995**。

4. Russel, A. A.; Siebert, E. D. *Experiments for General Chemistry*; 5th ed., Burgess International Group, Inc.: Minnesota, **1989**; p 18.

5. Skoog, D. A.; Leary, J. J. *Principles of Instrumental Analysis*; 4th ed., Saunders College Publishing: Chicago, **1988**; pp 477, 479, 495, 496.

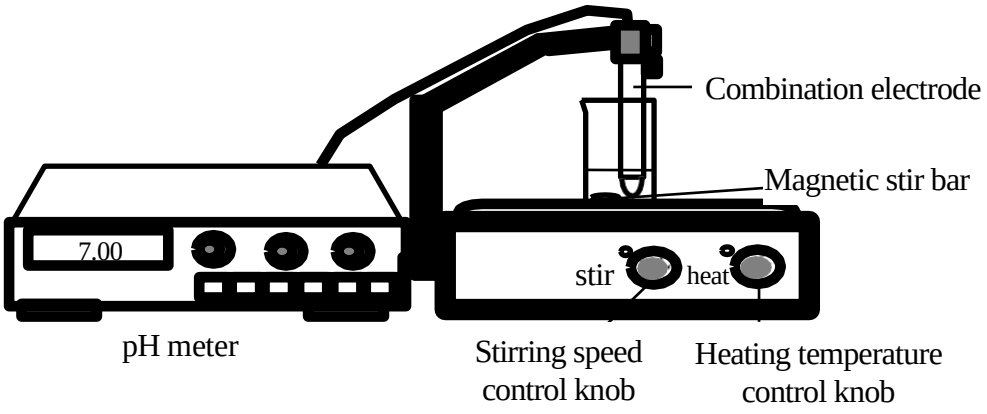


Figure 12-2. Schematic diagram of the pH meter setup for measuring the pH of a solution.

Experiment 12: Buffer Solution

Name: _____ **Student ID:** _____

Department: _____ **Group:** _____ **Date:** _____

I. Experimental Data and Results

(A) Calculation of Acid/Base Drop Volume

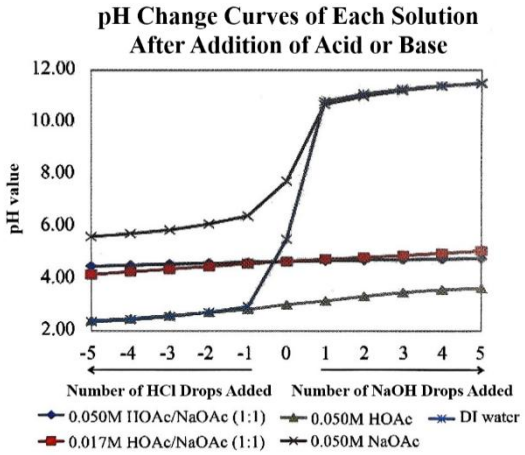
Acid/Base Solution	Number of Drops	Volume (mL)	Average Volume (mL/drop)
1.0 M HCl			
1.0 M NaOH			

(B) Effect of Concentration on Buffer Capacity

Acid/Base Solution		(a) 0.05 M HOAc / NaOAc	(b) 0.017M HOAc / NaOAc	(c) 0.05 M HOAc	(d) 0.05 M NaOAc	(e) DI water
1.0 M HCl	5					
	4					
	3					
	2					
	1					
	0					
Average (0)						
1.0 M NaOH	0					
	1					
	2					
	3					
	4					
	5					

(C) Data Analysis and Plotting

1. Plot the pH value (y-axis) against the number of acid/base drops added (x-axis) for solutions (a)–(e), as shown in the figure below.



2. Select the middle five data points from solutions (a) and (b), then create an XY scatter plot and determine the least-squares linear fit.

   Note: The plot can be made using MS Excel. Choose a scatter plot that displays only the data points, then add a linear trendline along with the trendline equation and the R^2 value. Print the data table and the graph and include them in the lab report.

3. Using the slope of the linear trendline (i.e., the least-squares line, ΔpH/drop), the volume of one drop of solution (V_{drop}), the concentration of HCl or NaOH (1.0 M), and the buffer solution volume (30 mL), calculate the buffer capacity of solutions (a) and (b).

$$\text{Buffer capacity}\left(\frac{\text{meq}}{\text{L}\cdot\text{pH}}\right) = \frac{1}{\text{slope}}\left(\frac{\text{drop}}{\text{pH}}\right) \times V_{\text{drop}}\left(\frac{\text{mL}}{\text{drop}}\right) \times C_{\text{M}}\left(\frac{\text{eq}}{\text{L}}\right) \times \frac{1000\ (\text{mL/L})}{30\ \text{mL}}$$

4. Compare the buffer capacities of solutions (a) and (b). Which one is the better buffer solution?

II. Conclusion

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III. Question and Discussion

1. For the prepared solutions (a) 0.050 M HOAc / NaOAc (1:1) and (b) 0.017 M HOAc / NaOAc (1:1) in this experiment, what are their theoretical pH values? Show the equations used to determine them.

2. If a buffer solution is prepared using NaH_2PO_4 and Na_2HPO_4, what is its optimal buffering pH range? Write the reaction equations showing how this buffer neutralizes added H^+ and OH^-.