A pH Study of Acids and Salt Solutions

Cautions

Acetic acid and hydrochloric acid are toxic, irritants and corrosive. Ionic salts may be irritants and toxic.

Avoid skin contact from all chemicals and rinse thoroughly with running water if contact occurs.

Purpose

To compare calculated and measured pH values of a series of hydrochloric acid solutions, acetic acid solutions, and various salt solutions. Based on the measured pH values, the acid ionization constant (K_a) of acetic acid will be calculated.

Introduction

I. Acid-Base Properties of Water

Many types of compounds will dissolve in water to form aqueous solutions. A factor that contributes to the power of water as a solvent is the fact that it undergoes a self-ionization (often referred to as "autoionization") process

$$2\,H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + OH^-_{(aq)} \qquad (1)$$

Due to this autoionization process, small amounts of hydronium ion, H_3O^+, and hydroxide ion, OH^- are present. The equilibrium constant expression for this reaction is known as K_w, where:

$$K_w = [H_3O^+]\,[OH^-] \qquad (2)$$

For pure water at 25 °C the, the value of K_w is $1.0\text{x}10^{-14}$; based on this expression, in pure water at 25 °C, the molar concentrations of $[H_3O^+]$ and $[OH^-]$ are both $1.0 \times 10^{-7} M$. Different aqueous solutions may have widely different $[H_3O^+]$ and $[OH^-]$ values but their product will always equal 1.0×10^{-14} at 25 °C.

Sorensen, a Danish chemist, devised the exponential system for expressing molar concentration of H_3O^+ called pH, which provides information about the acidic nature of a solution in whole numbers. The lowercase p indicates "take the negative log of", the H refers to $[H^+]$ or $[H_3O^+]$. Thus, the expression for calculating the pH is:

$$pH = -\log [H_3O^+] \quad or \quad [H_3O^+] = 10^{-pH} \qquad (3)$$

A similar expression relates the basic nature of a solution to $[OH^-]$:

$$pOH = -\log [OH^-] \quad or \quad [OH^-] = 10^{-pOH} \qquad (4)$$

Using the K_w expression and rules of logs, the relationship between pH and pOH is:

$$pH + pOH = 14.00 \qquad (5)$$

II. Strong Acids and Bases in water

When Arrhenius or Bronsted-Lowry acids are dissolved in water, $[H_3O^+]$ increases. This change in concentration is due to the *dissociation* of at least one H^+ ion from a single acid molecule. The strength of an acid is determined by the extent to which it dissociates in water. Strong acids are assumed to dissociate completely, and $[H_3O^+]$ is determined by:

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$$[H_3O^+] = n[H_nA] \qquad (6)$$

where H_nA is a strong acid having *n* ionizable protons. Hydrochloric acid (HCl) and nitric acid (HNO_3) are examples of **strong acids**.

Similar statements regarding complete dissociation can be made for bases. Strong bases are compounds that completely dissociate in water to form one or more cations and hydroxide ions (OH^-). Sodium hydroxide (NaOH) and potassium hydroxide (KOH) are examples of strong bases.

III. Weak Acids

Many acids only dissociate to a small extent when dissolved in water; these substances are considered ***weak acids***. A weak acid solution will have $[H_3O^+]$ much less than the concentration of acid added to the solution ($[H_3O^+] << [HA]$). The dissociation of a weak acid in water is an equilibrium process:

$$HA_{(aq)} + H_2O_{(l)} \leftrightarrows H_3O^+_{(aq)} + A^-_{(aq)} \qquad (7)$$

and a corresponding acid dissociation equilibrium expression (***K_a***) can be written:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \qquad (8)$$

where $[A^-]$ is the concentration of conjugate base and [HA] is the concentration of conjugate acid. For a given weak acid solution, if one knows the initial acid concentration and K_a, a theoretical pH can be determined. If instead, the pH and initial concentration of acid are known, a value of K_a can be experimentally determined.

In some cases, the mathematical problem can be simplified due to insignificant dissociation of the weak acid. When the ***K_a*** of the weak acid is quite small compared to the bulk concentration (typically $>10^3$ difference), the concentration of acid at equilibrium can be assumed to be equal to the initial concentration of weak acid before dissociation.

IV. Weak Bases

Weak bases are compounds that produce small amounts of the hydroxide ion in solution when dissolved in water. Like weak acids, only a small percentage of the molecules of a weak base dissociate in water. When ammonia gas is bubbled into water, the process:

$$NH_{3\,(g)} + H_2O_{(l)} \leftrightarrows NH_4^+{}_{(aq)} + OH^-_{(aq)} \qquad (9)$$

occurs. A base ionization constant expression, ***K_b***, for the process is written:

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5} \qquad (8)$$

Notice that the ***K_b*** value is significantly less than 1, consistent with a weak base. This expression is analogous to the ***K_a*** for a weak acid. Based on the ***K_b*** expression the concentrations $[NH_4^+] = [OH^-]$ will be equal *at equilibrium* for this reaction.

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V. Aqueous Salt Solutions

Salts are ionic compounds that form a cation and an anion when dissolved in water. Once dissociated from each other, the ions of the salt may react with water in a process called ***hydrolysis***, which may influence the pH of the solution. If no hydrolysis or very little occurs with either ion (e.g. KBr) the pH of the resulting solution will be close to 7.0 or neutral. The anions of strong acids, such as NO_3^-, SO_4^{2-}, and Br^-, and the cations of strong ionic bases, such as Na^+, K^+, and Ca^{2+}, do not hydrolyze. If hydrolysis occurs, then the pH of the solution may be acidic or basic depending upon the ions present in the original salt. For example, when sodium carbonate is dissolved in water, it breaks into ions:

$$Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-} \qquad (10)$$

The sodium cation (Na^+) will not react significantly with OH^- in solution (NaOH is a strong base); however, the carbonate anion may react with H_2O, producing bicarbonate and hydroxide:

$$CO_3^{2-} + H_2O \leftrightarrows HCO_3^- + OH^- \qquad (11)$$

resulting in a slightly basic solution. Salts may be divided into four general categories, based upon whether the cation, the anion, or both undergo hydrolysis; these categories are listed and described in Table 1.

Type	Characteristics	Examples
Neutral salts	Result from a reaction between a strong base and a strong acid No measurable hydrolysis occurs Solution is neutral	NaCl KBr KNO_3
Basic salts	Result from a reaction between a strong base and a weak acid Only the anion hydrolyzes Solution is basic.	KClO $KC_2H_3O_2$ $NaC_4H_7O_2$
Acidic salts	Result from a reaction between a weak base and a strong acid Only the cation hydrolyzes Solution is acidic.	NH_4Cl NH_4NO_3
Salts that could be acidic, basic, or neutral	Result from a reaction between a weak base and a weak acid Both ions hydrolyze Solution pH depends on the extent of hydrolysis of each ion.	NH_4ClO $NH_4C_2H_3O_2$

Table 1: Four types and examples of salts

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Procedure

NOTE*: Distilled water (not deionized) will be used for all experimental procedures.*

A. Strong Acid Solutions – Preparation by serial dilution and pH Measurement

A1. Label four small test tubes “**A**”, “**B**”, “**C**”, and “**D**”.

A2. Obtain 15 mL of 0.10 *M* HCl in a clean, dry beaker.

A3. Fill test tube “**A**” to 1/3 capacity with 0.10 *M* HCl solution.

A4. Record the exact concentration on the data sheet (0.10 M HCl may be 0.098 or 0.12 M).

First dilution

A5. Using a volumetric pipette, transfer 5.00 mL of 0.10 *M* HCl solution into a clean, dry 100-mL graduated cylinder.

A6. Dilute the solution in the graduated cylinder to exactly 50.0 mL with distilled water.

A7. Mix the solution with a glass-stirring rod.

A8. Fill test tube “**B**” to 1/3 capacity with this new solution.

Second dilution

A9. Using a volumetric pipette, transfer exactly 5.00 mL of the solution prepared in the first dilution (**A6**) to a second 100-mL graduated cylinder.

A10. Dilute the solution in the graduated cylinder to exactly 50.0 mL with distilled water.

A11. Mix the solution with a glass-stirring rod.

A12. Fill test tube “**C**” to 1/3 capacity with this new solution.

Third Dilution

A13. Using a volumetric pipette, transfer exactly 5.00 mL of the solution prepared in the first dilution (**A10**) to a second 100-mL graduated cylinder.

A14. Dilute the solution in the graduated cylinder to exactly 50.0 mL with distilled water.

A15. Mix the solution with a glass-stirring rod.

A16. Fill test tube “**C**” to 1/3 capacity with this new solution.

Measurement

A17. Rinse the pH electrode with distilled water into a waste beaker and lower the electrode into the solution in test tube “**A**” until the glass bulb is completely beneath the solution surface.

A18. Read and record the solution pH.

A19. Repeat the above procedure using the solutions in the remaining test tubes, making sure to rinse with distilled water between measurements.

A20. Transfer all HCl solutions and rinses into the “Discarded Solutions” beaker.

A21. Rinse the electrode with distilled water and store it in provided solution.

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B. Weak Acid Solutions - Preparation by serial dilution and pH Measurement

***NOTE**: in preparing the dilutions, be sure to rinse the transfer pipette between solutions.*

B1. Label four small test tubes "**1**", "**2**", "**3**", and "**4**."

B2. Obtain 15 mL of 0.25 *M* CH_3COOH in a clean, dry beaker.

B3. Fill test tube "**1**" to 1/3 capacity with 0.25 *M* CH_3COOH solution.

B4. Record the exact concentration on the data sheet.

First dilution

B5. Using a volumetric pipette, transfer 5.00 mL of 0.25 *M* CH_3COOH solution into a clean, dry 100-mL graduated cylinder.

B6. Dilute the solution in the graduated cylinder to exactly 50.0 mL with distilled water.

B7. Mix the solution with a glass-stirring rod.

B8. Fill test tube "**2**" to 1/3 capacity with this new solution.

Second dilution

B9. Using a volumetric pipette, transfer exactly 5.00 mL of the solution prepared in the first dilution (**B6**) to a second 100-mL graduated cylinder.

B10. Dilute the solution in the graduated cylinder to 50.0 mL with distilled water.

B11. Mix the solution with a glass-stirring rod.

B12. Fill test tube "**3**" to 1/3 capacity with this second new solution.

Third Dilution

B13. Using a volumetric pipette, transfer exactly 5.00 mL of the solution a prepared in the second dilution (**B10**) to a 100-mL graduated cylinder.

B14. Dilute the solution in the 100-mL graduated cylinder to 50.0 mL with distilled water.

B15. Fill test tube "**4**" to 1/3 capacity with this third new solution.

Measurement

B16. Measure the pH of the solutions in test tubes **1-4**, rinsing the electrode with distilled water between measurements.

B17. Dispose of all solutions used in this part to the proper waste container.

C. Salt Solutions – pH Measurement

C1. Obtain four clean small test tubes. Label them "$NaCH_3CO_2$", "NH_4Cl", "NaCl", and "$NaHCO_3$".

C2. Fill each test tube 1/3 of the way filled with the appropriate salt solution.

C3. Measure and record the pH of the resulting solution with the pH meter, making sure to rinse the electrode between measurements.

C4. These solutions may all be poured down the drain with running water.

Disposal

Dispose of all solutions into the appropriate waste containers.

Clean-Up

Wash all glassware with soap then rinse 3 times with tap water, and once with distilled water.

Clean your work area with water and dry with paper towels. Wash your hands before leaving the laboratory.

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Calculations and Helpful Hints

Perform the following calculations, record the results on the Data Sheets, and ***show all work CLEARLY on extra paper****.*

A. Hydrochloric Acid Solution and pH

- Calculate the theoretical pH of each HCl solution

B. Acetic Acid Solution and pH

- Calculate the concentration of each acetic acid solution using the dilution equation ($M_1V_1 = M_2V_2$)
- Calculate the theoretical pH of each acetic acid solution using K_a = 1.8x10^{-5}.
- Calculate an experimental value of K_a for acetic acid using the measured pH values for each solution.
- Calculate average value and the percent error of the experimentally determined K_a.

C. pH *of Salt Solutions*

- Classify each salt as acidic, basic or neutral based on the measured pH values.

A pH Study of Acids and Salt Solutions
Data Sheet

Name: ______________________________ **Lab Partner**______________________________

A. Hydrochloric Acid Solution and pH

Solution	Concentration *moles/L*	Measured pH	Theoretical pH
A			
B			
C			
D			

B. Acetic Acid Solution and pH

Solution	Concentration *moles/L*	Measured pH	Theoretical pH	Calculated K_a
1				
2				
3				
4				
			Average K_a	

C. pH of Salt Solutions

Solution	Concentration *moles/L*	Measured pH	Salt Type	Predicted Salt Type (from pre-lab)
$NaCH_3CO_2$				
NH_4Cl				
NaCl				
$NaHCO_3$				

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Post-lab Assignment

Name: ____

1. Explain why a 0.10 molar hydrochloric acid solution has a different pH value compared to a 0.10 molar acetic acid solution. Use chemical reactions to support your answer.

2. Calculate the percent error in the value of K_a calculated in Part B, and explain any sources of error in this measurement.

A pH Study of Acids and Salt Solutions
Pre-lab Assignment

Name: ______________________________________

1. Calculate each of the following values for a solution whose $[H_3O^+] = 2.1 \times 10^{-4}$ M

   pH = ____________ pOH = ___________ $[OH^-]$ = ____________

2. Calculate each of the following values for a solution having a pH = 6.21

   $[H_3O^+]$ = ____________ pOH = ___________ $[OH^-]$ = ____________

3. Predict whether the following salts are acidic, neutral or basic when dissolved in water. Write balanced chemical equations that support your answer.

Salt	**$NaCH_3CO_2$**	**NH_4Cl**	**NaCl**	**$NaHCO_3$**
Nature				