Understanding Chemistry Unit 6 Worksheet 3: A Deep Dive into Acids, Bases, and Solutions

Welcome to a comprehensive exploration of the concepts typically found in a Chemistry Unit 6 Worksheet 3. This resource is designed to equip students with a thorough understanding of acids, bases, and the intricate world of solutions. We will meticulously break down the fundamental principles governing acid-base behavior, delve into the measurement and interpretation of solution properties like pH, and explore the practical applications of these concepts. Whether you're grappling with titration curves, buffer solutions, or the stoichiometry of acid-base reactions, this guide aims to provide clarity and confidence. By dissecting each key area, we ensure that students gain not only the knowledge but also the skills necessary to successfully complete their chemistry coursework related to Unit 6, Worksheet 3.

Navigating Chemistry Unit 6 Worksheet 3: Acids, Bases, and Solution Chemistry

Chemistry Unit 6 Worksheet 3 often serves as a pivotal point in a student's learning journey, solidifying their understanding of the fundamental principles of acids, bases, and the behavior of solutions. This unit typically delves into the defining characteristics of acidic and basic substances, their interactions, and the quantitative methods used to analyze them. Mastering the content presented in Chemistry Unit 6 Worksheet 3 is crucial for building a robust foundation in general chemistry and preparing for more advanced topics. This comprehensive guide is designed to demystify these complex concepts, providing clear explanations, illustrative examples, and strategic approaches to problem-solving. We will explore the theoretical underpinnings as well as the practical implications, ensuring a well-rounded comprehension of the material. Success with your Chemistry Unit 6 Worksheet 3 begins here.

The Foundational Concepts of Acids and Bases

At its core, understanding Chemistry Unit 6 Worksheet 3 necessitates a firm grasp of what defines an acid and a base. Historically, definitions have evolved, but the most commonly used in introductory chemistry are the Arrhenius, Brønsted-Lowry, and Lewis theories. The Arrhenius theory defines acids as substances that increase the concentration of hydrogen ions (H^+) in aqueous solutions and bases as substances that increase the concentration of hydroxide ions (OH^-) in aqueous solutions. While useful, this definition is limited to aqueous systems.

The Brønsted-Lowry theory offers a broader perspective, defining an acid as a proton (H^+) donor and a base as a proton acceptor. This definition is particularly important for understanding reactions in non-aqueous solvents and the concept of conjugate acid-base pairs. When an acid donates a proton, it forms its conjugate base, and when a base accepts a proton, it forms its conjugate acid. This dynamic equilibrium is a cornerstone of acid-base chemistry.

The Lewis theory provides the most general definition, where an acid is an electron-pair acceptor, and a base is an electron-pair donor. This theory is essential for understanding reactions that do not involve proton transfer, such as the formation of coordinate covalent bonds. Familiarity with these definitions is paramount for successfully tackling problems on your Chemistry Unit 6 Worksheet 3.

Arrhenius Theory: The Aqueous Perspective

The Arrhenius definition, while foundational, is specific to how substances behave when dissolved in water. An Arrhenius acid, such as hydrochloric acid (HCl), dissociates in water to produce H^+ ions: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$. Conversely, an Arrhenius base, like sodium hydroxide ($NaOH$), dissociates to produce OH^- ions: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$.

This simple dissociation model helps explain the characteristic properties of acids (sour taste, ability to corrode metals) and bases (bitter taste, slippery feel). However, its limitation lies in its exclusive focus on aqueous solutions and the direct production of H^+ and OH^- ions, which are often present in combined forms like hydronium ions (H_3O^+).

Brønsted-Lowry Theory: Proton Transfer Mastery

The Brønsted-Lowry theory expands our understanding by focusing on the transfer of protons. For instance, when HCl dissolves in water, it donates a proton to a water molecule, forming hydronium ions (H_3O^+) and chloride ions (Cl^-): $HCl(aq) + H_2O(l) <=> H_3O^+(aq) + Cl^-(aq)$. In this reaction, HCl is the acid (proton donor), and water acts as a base (proton acceptor).

Similarly, ammonia (NH_3) acts as a base when it reacts with water: $NH_3(aq) + H_2O(l) <=> NH_4^+(aq) + OH^-(aq)$. Here, NH_3 is the base (proton acceptor), and water acts as an acid (proton donor). Understanding conjugate pairs is vital; in the first example, Cl^- is the conjugate base of HCl, and H_3O^+ is the conjugate acid of H_2O. Recognizing these pairs is a key skill for Chemistry Unit 6 Worksheet 3.

Lewis Theory: The Electron Pair Exchange

The Lewis theory broadens the scope of acid-base chemistry to include reactions where no protons are exchanged. A Lewis acid accepts an electron pair to form a covalent bond, while a Lewis base donates an electron pair. A classic example is the reaction between boron trifluoride (BF_3) and ammonia (NH_3). BF_3, with an incomplete octet, is electron-deficient and acts as a Lewis acid, accepting a pair of electrons from the lone pair on the nitrogen atom in NH_3, which acts as a Lewis base. This forms an adduct: $BF_3 + NH_3 \rightarrow F_3B\text{-}NH_3$.

This definition is powerful because it encompasses a wider range of chemical reactions and species that exhibit acidic or basic properties, even those that don't contain hydrogen. It's a more abstract but fundamental concept in understanding chemical reactivity.

The pH Scale: Quantifying Acidity and Basicity

A central theme in Chemistry Unit 6 Worksheet 3 is the pH scale, a convenient way to express the concentration of hydrogen ions (or hydronium ions) in an aqueous solution. The pH is defined as the negative logarithm (base 10) of the hydrogen ion concentration: $pH = -\log[H^+]$. The scale typically ranges from 0 to 14, with a pH of 7 considered neutral at 25°C.

Solutions with a pH less than 7 are acidic, meaning they have a higher concentration of H^+ ions than OH^- ions. Solutions with a pH greater than 7 are basic (alkaline), indicating a higher concentration of OH^- ions than H^+ ions. A tenfold change in $[H^+]$ concentration results in a one-unit change in pH, highlighting the logarithmic nature of the scale.

Understanding the relationship between pH, pOH, and the ion product constant of water (K_w) is also critical. $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C. From this, we derive $pK_w = pH + pOH = 14$. This equation allows us to calculate the pH of a solution if we know the concentration of either H^+ or OH^- ions, and vice versa. These calculations are fundamental to solving problems on Chemistry Unit 6 Worksheet 3.

Calculating pH and pOH

Accurate pH and pOH calculations are vital. For instance, if a solution has a hydrogen ion concentration of 1.0×10^{-4} M, its pH is calculated as: $pH = -\log(1.0 \times 10^{-4}) = 4.0$. If the pOH of a solution is 9.0, then the pH is $14.0 - 9.0 = 5.0$.

Conversely, if the pH of a solution is 2.5, we can find the $[H^+]$ concentration by rearranging the pH formula: $[H^+] = 10^{-pH} = 10^{-2.5} \approx 3.16 \times 10^{-3}$ M. Similarly, if the pH is 2.5, then the pOH is $14 - 2.5 = 11.5$, and the $[OH^-]$ concentration is $10^{-11.5} \approx 3.16 \times 10^{-12}$ M.

The Ion Product of Water (Kw)

Water itself undergoes autoionization to a small extent, producing both hydrogen ions and hydroxide ions: $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$. The equilibrium constant for this reaction is the ion product of water, K_w. Its value is temperature-dependent, but at 25°C, it is a constant 1.0×10^{-14}. This fundamental constant links the concentrations of H^+ and OH^- in any aqueous solution, whether acidic, basic, or neutral.

In a neutral solution, $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ M, leading to a pH of 7. In acidic solutions, $[H^+] > 1.0 \times 10^{-7}$ M and $[OH^-] < 1.0 \times 10^{-7}$ M. In basic solutions, $[H^+] < 1.0 \times 10^{-7}$ M and $[OH^-] > 1.0 \times 10^{-7}$ M. Mastering these relationships is crucial for Chemistry Unit 6 Worksheet 3.

Strong vs. Weak Acids and Bases

A critical distinction in acid-base chemistry, and a common topic on Chemistry Unit 6 Worksheet 3, is the difference between strong and weak acids and bases. This difference lies in their degree of dissociation or ionization in water.

Strong acids and bases dissociate completely in water, meaning that essentially all of the acid or base molecules react to form ions. For example, hydrochloric acid (HCl) is a strong acid; when dissolved in water, it completely ionizes into H^+ and Cl^- ions. Similarly, sodium hydroxide ($NaOH$) is a strong base and dissociates completely into Na^+ and OH^- ions.

Weak acids and bases, on the other hand, only partially dissociate in water. They establish an equilibrium between the undissociated molecules and their ions. For example, acetic acid (CH_3COOH) is a weak acid. When dissolved in water, only a fraction of the CH_3COOH molecules ionize to form H^+ and CH_3COO^- ions, while most remain as intact molecules. The extent of dissociation for weak acids and bases is quantified by their acid dissociation constant (K_a) and base dissociation constant (K_b), respectively.

Dissociation of Strong Acids and Bases

Strong acids like HCl, HBr, HI, HNO_3, H_2SO_4, and $HClO_4$ are characterized by their 100% ionization in water. This means that for a given molar concentration of a strong acid, the concentration of H^+ ions produced is equal to the initial concentration of the acid. For example, a 0.1 M solution of HCl will have $[H^+] = 0.1$ M, resulting in a pH of 1.

Similarly, strong bases like alkali metal hydroxides (e.g., $NaOH$, KOH) and some alkaline earth metal hydroxides (e.g., $Ca(OH)_2$, $Ba(OH)_2$) also dissociate completely. A 0.1 M solution of $NaOH$ will have $[OH^-] = 0.1$ M, leading to a pOH of 1 and a pH of 13.

The Role of Ka and Kb for Weak Acids and Bases

Weak acids and bases exist in equilibrium. For a weak acid HA, the dissociation reaction is $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$. The acid dissociation constant, K_a, is given by the expression: $K_a = ([H^+][A^-]) / [HA]$. A smaller K_a value indicates a weaker acid, meaning it dissociates less.

For a weak base B, the reaction with water is $B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$. The base dissociation constant, K_b, is given by: $K_b = ([BH^+][OH^-]) / [B]$. A smaller K_b value indicates a weaker base.

These constants are crucial for calculating the pH of solutions containing weak acids or bases and are frequently tested in Chemistry Unit 6 Worksheet 3. For conjugate acid-base pairs, the relationship $K_a K_b = K_w$ holds true.

Neutralization Reactions and Their Stoichiometry

Neutralization reactions occur when an acid reacts with a base, typically forming a salt and water. The fundamental principle is that the H^+ ions from the acid react with the OH^- ions from the base to form H_2O. The overall reaction is often a spectator ion displacement.

For example, the reaction between hydrochloric acid (HCl) and sodium hydroxide (NaOH) is: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$. The net ionic equation is $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. This reaction is exothermic, releasing heat.

Understanding the stoichiometry of these reactions is essential for quantitative analysis, especially in titration. The point at which the moles of acid are stoichiometrically equivalent to the moles of base is known as the equivalence point. For strong acid-strong base titrations, the pH at the equivalence point is 7. For titrations involving weak acids or bases, the pH at the equivalence point will deviate from 7.

The Equivalence Point in Titrations

The equivalence point is a theoretical point in a titration where the amount of titrant added is just enough to completely react with the analyte, according to the stoichiometry of the reaction. For a monoprotic acid and a monobasic base, the moles of H^+ from the acid will equal the moles of OH^- from the base at the equivalence point. This can be expressed as: $M_{acid} V_{acid} = M_{base} V_{base}$, where M represents molarity and V represents volume.

This equation is fundamental for calculating unknown concentrations when performing titrations. However, it's important to remember that it strictly applies only when the moles of reacting species are in a 1:1 ratio. For polyprotic acids or bases, the stoichiometry might be different, requiring adjustments to the calculation.

Salt Formation and Hydrolysis

When an acid and base react, they form a salt. The nature of the salt (acidic, basic, or neutral) depends on the strengths of the parent acid and base. Salts formed from the reaction of a strong acid and a strong base (e.g., NaCl from HCl and NaOH) are neutral because neither the cation nor the anion undergoes significant hydrolysis.

Salts formed from a strong acid and a weak base (e.g., NH_4Cl from NH_3 and HCl) will produce an acidic solution. The cation from the weak base (e.g., NH_4^+) will react with water to produce H^+ ions (hydrolysis): $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$.

Salts formed from a weak acid and a strong base (e.g., CH_3COONa from CH_3COOH and NaOH) will produce a basic solution. The anion from the weak acid (e.g., CH_3COO^-) will react with water to produce OH^- ions (hydrolysis): $CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$.

Titration: A Quantitative Analysis Tool

Titration is a cornerstone technique in analytical chemistry, widely used to determine the concentration of an unknown solution (the analyte) by reacting it with a solution of known concentration (the titrant). In the context of acid-base chemistry, an acid-base titration is performed by slowly adding a base of known concentration to a known volume of an acid of unknown concentration, or vice versa.

The process typically involves a burette to deliver the titrant accurately, a flask to contain the analyte, and an indicator that changes color at or near the equivalence point. The point at which the indicator changes color is called the endpoint. Ideally, the endpoint should coincide with the equivalence point.

Titrations are invaluable for practical applications, such as determining the acidity of vinegar, the concentration of household cleaners, or the purity of pharmaceutical compounds. Mastery of titration principles is a key outcome expected from completing a Chemistry Unit 6 Worksheet 3.

Performing an Acid-Base Titration

A typical acid-base titration procedure involves several steps. First, a precise volume of the analyte (e.g., an unknown concentration of HCl) is measured into a flask. Then, a few drops of an appropriate acid-base indicator are added. The titrant (e.g., a standard solution of NaOH) is placed in a burette and slowly added to the analyte, with constant swirling, until the indicator shows a permanent color change, signifying the endpoint.

The volume of titrant used to reach the endpoint is recorded. Using the known concentration of the titrant, the volume of analyte, and the stoichiometry of the reaction, the concentration of the analyte can be calculated. For instance, if 25.00 mL of an HCl solution requires 20.00 mL of 0.100 M NaOH to reach the endpoint, the concentration of HCl can be determined.

Choosing the Right Indicator

The selection of an appropriate acid-base indicator is crucial for accurate titration results. An indicator is a weak acid or base that changes color over a specific pH range. The ideal indicator for a titration is one whose color change range (transition range) brackets the pH of the solution at the equivalence point.

For a strong acid-strong base titration, the pH at the equivalence point is 7. Indicators like phenolphthalein (pH range 8.2-10) or bromothymol blue (pH range 6.0-7.6) can be used, although phenolphthalein might lead to a slightly delayed endpoint. For a weak acid-strong base titration, the equivalence point occurs at a pH greater than 7 due to the hydrolysis of the conjugate base of the weak acid; phenolphthalein is often the preferred indicator.

Conversely, for a strong acid-weak base titration, the equivalence point is at a pH less than 7; indicators like methyl orange (pH range 3.1-4.4) or methyl red (pH range 4.4-6.2) are more suitable.

Interpreting Titration Curves

A titration curve is a graph that plots the pH of the solution as a function of the volume of titrant added. These curves provide valuable information about the strength of the acid and base involved and the nature of the equivalence point. Understanding how to interpret these curves is a key skill for Chemistry Unit 6 Worksheet 3.

A typical titration curve for a strong acid with a strong base will show a gradual increase in pH as the base is added, followed by a very sharp increase in pH around the equivalence point, and then a leveling off at a high pH. The steepness of the pH change around the equivalence point reflects the strength of the acid and base.

Titration curves for weak acids or bases exhibit different shapes. For a weak acid titrated with a strong base, the curve starts at a lower pH than a strong acid, has a buffer region where the pH changes slowly, and then a less steep rise at the equivalence point, which will be above pH 7. Similarly, titrating a strong acid with a weak base will result in a curve starting at a higher pH and an equivalence point below pH 7.

Features of a Strong Acid-Strong Base Titration Curve

In a titration of a strong acid (e.g., HCl) with a strong base (e.g., NaOH), the initial pH is very low. As NaOH is added, OH- ions neutralize H+ ions, and the pH gradually increases. A significant portion of the titration occurs before the equivalence point. At the equivalence point, the moles of NaOH added exactly equal the moles of HCl present. Since NaCl is a neutral salt, the solution at the equivalence point is neutral, with a pH of 7. Beyond the equivalence point, adding more NaOH causes a rapid increase in pH, and the curve levels off at a high pH, determined by the excess OH- concentration.

Features of a Weak Acid-Strong Base Titration Curve

When a weak acid (e.g., acetic acid, CH3COOH) is titrated with a strong base (e.g., NaOH), the initial pH is higher than that of a strong acid. The first part of the curve, after the initial pH, is characterized by the formation of a buffer solution as the weak acid is converted to its conjugate base (acetate ion, CH3COO-). Within this buffer region, the pH changes relatively slowly with the addition of titrant. The midpoint of this buffer region, where half of the weak acid has been neutralized, is called the half-equivalence point, and at this point, pH = pKa of the weak acid.

At the equivalence point, all the weak acid has been converted to its conjugate base. Since the conjugate base of a weak acid is a weak base, it will hydrolyze water, producing OH- ions, making the solution basic. Therefore, the pH at the equivalence point for a weak acid-strong base titration is always greater than 7.

Features of a Strong Acid-Weak Base Titration Curve

Titrating a strong acid (e.g., HCl) with a weak base (e.g., ammonia, NH3) yields a titration curve with an initial high pH. As the strong base is added, it neutralizes the strong acid, and the pH decreases. A buffer region is established as the weak base is converted to its conjugate acid (e.g., NH4+). At the half-equivalence point, pH = pKa of the conjugate acid of the weak base.

At the equivalence point, all the weak base has reacted with the strong acid to form its conjugate acid. Since the conjugate acid of a weak base is a weak acid, it will hydrolyze water, producing H^+ ions, making the solution acidic. Thus, the pH at the equivalence point for a strong acid-weak base titration is always less than 7.

Buffer Solutions: Resisting pH Change

Buffer solutions are critical components of biological systems and chemical processes, playing a vital role in maintaining a stable pH. A buffer solution is a mixture that resists changes in pH upon the addition of small amounts of acid or base, or upon dilution. This resistance to pH change is a key concept often reinforced in Chemistry Unit 6 Worksheet 3.

Buffer solutions typically consist of a weak acid and its conjugate base (e.g., acetic acid and sodium acetate) or a weak base and its conjugate acid (e.g., ammonia and ammonium chloride). The key to buffering capacity lies in the presence of both the acidic component and the basic component in significant concentrations. When a small amount of strong acid is added, the basic component of the buffer reacts with it. When a small amount of strong base is added, the acidic component of the buffer reacts with it, thus minimizing the change in pH.

The pH of a buffer solution can be calculated using the Henderson-Hasselbalch equation: pH = pKa + log([conjugate base]/[weak acid]). This equation demonstrates how the pH of a buffer is related to the pKa of the weak acid and the ratio of the concentrations of the conjugate base and the weak acid.

How Buffers Work

Consider a buffer made from a weak acid (HA) and its conjugate base (A^-). If a strong acid (H^+) is added, the conjugate base (A^-) reacts with the added H^+ to form more weak acid (HA): $A^- + H^+ \Leftrightarrow HA$. The net effect is that the concentration of H^+ ions does not increase significantly, as the strong acid is converted into a weak acid. If a strong base (OH^-) is added, the weak acid (HA) reacts with the OH^- to form its conjugate base (A^-) and water: $HA + OH^- \Leftrightarrow A^- + H_2O$. The net effect is that the concentration of OH^- ions does not increase significantly, as the strong base is neutralized by the weak acid.

The effectiveness of a buffer is measured by its buffer capacity, which is the amount of acid or base the buffer can neutralize before the pH begins to change significantly. Buffer capacity is greatest when the concentrations of the weak acid and its conjugate base are equal, i.e., when pH = pKa.

The Henderson-Hasselbalch Equation

The Henderson-Hasselbalch equation is a powerful tool for calculating the pH of buffer solutions and understanding the relationship between pH, pKa, and the ratio of conjugate base to weak acid. It is derived from the Ka expression for a weak acid: Ka = [H+][A-]/[HA]. By taking the negative logarithm of both sides and rearranging, we arrive at: pH = pKa + log([A-]/[HA]).

This equation is invaluable for preparing buffer solutions of a specific pH. By choosing a weak acid with a pKa close to the desired pH and adjusting the ratio of conjugate base to weak acid, one can create an effective buffer. For example, to prepare a buffer with a pH of 4.74 using acetic acid (pKa = 4.76), one would mix equal concentrations of acetic acid and acetate ions.

Common Mistakes and How to Avoid Them in Chemistry Unit 6 Worksheet 3

While tackling Chemistry Unit 6 Worksheet 3, students often encounter common pitfalls. Awareness of these potential errors can significantly improve accuracy and understanding. One frequent mistake is confusing strong and weak acids/bases and their implications for dissociation and pH. Another common error involves misapplying the stoichiometry in neutralization reactions or titrations, especially with polyprotic acids or bases.

Calculation errors, particularly with logarithms and exponents when dealing with pH and pOH, are also prevalent. Students may also struggle with correctly identifying conjugate acid-base pairs or incorrectly applying the Henderson-Hasselbalch equation, perhaps by switching the positions of the acid and base components or misinterpreting the role of pKa.

To avoid these mistakes, it is essential to carefully review the definitions of acids, bases, strong vs. weak behavior, and the specific context of each problem. Practicing a variety of problems, checking units, and understanding the underlying principles rather than just memorizing formulas can lead to greater success.

- Confusing strong and weak acids/bases: Always remember that strong acids/bases dissociate completely, while weak ones do not.
- Stoichiometry errors: Double-check the mole ratios in neutralization reactions and titrations, especially for polyprotic species.

- Logarithm and exponent mistakes: Be meticulous when calculating pH, pOH, and ion concentrations. Use a calculator carefully.
- Incorrect use of Henderson-Hasselbalch: Ensure you are using the correct pKa and that the ratio of conjugate base to weak acid is correctly placed in the logarithm term.
- Misinterpreting titration curves: Pay close attention to the starting pH, the shape of the curve, and the pH at the equivalence point.
- Forgetting Kw: Remember that $[H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C, and pH + pOH = 14.

Practice Problems and Solutions

To truly master the concepts in Chemistry Unit 6 Worksheet 3, consistent practice is key. Below are a few example problems that cover the core topics discussed. Working through these will reinforce your understanding and prepare you for assessments.

Problem 1: pH Calculation

Calculate the pH of a 0.050 M solution of sulfuric acid (H_2SO_4). Sulfuric acid is a strong diprotic acid, meaning it dissociates completely for the first proton, but the second dissociation is that of a weak acid. However, for simplicity in introductory problems, it's often treated as a strong monoprotic acid, meaning each mole of H_2SO_4 produces 2 moles of H^+.

Solution 1:

Assuming complete dissociation of both protons for simplicity:
$[H^+] = 2\ [H_2SO_4] = 2\ 0.050\ M = 0.100\ M$
$pH = -\log[H^+] = -\log(0.100) = 1.00$

Problem 2: Titration Stoichiometry

A 20.0 mL sample of an unknown monoprotic acid solution required 25.0 mL of 0.150 M sodium hydroxide (NaOH) to reach the equivalence point. What is the molarity of the acid solution?

Solution 2:

At the equivalence point, moles of acid = moles of base.
Moles of NaOH = Molarity Volume = 0.150 mol/L 0.0250 L = 0.00375 mol
Since the acid is monoprotic and NaOH is a monobasic base, the mole ratio is 1:1.
Therefore, moles of acid = 0.00375 mol.
Molarity of acid = Moles of acid / Volume of acid = 0.00375 mol / 0.0200 L = 0.188 M

Problem 3: Buffer pH Calculation

Calculate the pH of a buffer solution that contains 0.20 M acetic acid (CH_3COOH) and 0.30 M sodium acetate (CH_3COONa). The pKa of acetic acid is 4.76.

Solution 3:

Using the Henderson-Hasselbalch equation:
pH = pKa + log([conjugate base]/[weak acid])
pH = 4.76 + log([CH_3COONa]/[CH_3COOH])
pH = 4.76 + log(0.30 M / 0.20 M)
pH = 4.76 + log(1.5)
pH = 4.76 + 0.176
pH = 4.94

Conclusion: Mastering Your Chemistry Unit 6 Worksheet 3

Successfully navigating the complexities of acid-base chemistry and solution properties is essential for a strong foundation in chemistry, and your Chemistry Unit 6 Worksheet 3 is a critical step in this process. By thoroughly understanding the definitions of acids and bases, the nuances of the pH scale, the distinctions between strong and weak electrolytes, and the principles of neutralization and titration, you are well-equipped to tackle any challenge. The ability to interpret titration curves and understand the buffering capacity of solutions further solidifies your analytical skills. Remember to practice diligently, pay attention to detail in your calculations, and always strive for conceptual clarity. With a solid grasp of these topics, you can confidently approach your Chemistry Unit 6 Worksheet 3 and excel in your chemistry studies.