GENERAL

ap chemistry laboratory 19 ph properties of buffer solutions answers

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Introduction

In the AP Chemistry curriculum, Lab 19 is one of the most frequently cited challenges for students. It focuses on the **pH properties of buffer solutions**and requires a solid grasp of acid–base equilibria, the Henderson–Hasselbalch equation, and practical laboratory techniques. Whether you're a student preparing for the exam, a teacher looking for a clear explanation, or a tutor seeking reliable answers, this article offers a comprehensive guide that covers every aspect of Lab 19—from the basic concepts to step by step solutions and common pitfalls.

By the end of this article you will:

- Understand what a buffer solution is and why it matters in chemistry.
- Know how to calculate the pH of a buffer using the Henderson–Hasselbalch equation.
- Be able to answer all the questions in AP Chemistry Lab 19 with confidence.
- Recognize common mistakes and how to avoid them.
- Have a set of study tips to reinforce your learning.

Let's dive in.

Overview of AP Chemistry Lab 19

Lab 19 is titled **"pH Properties of Buffer Solutions"**and is part of the AP Chemistry exam's laboratory section. The lab typically involves preparing a buffer solution, measuring its pH, and then adding a strong acid or base to observe how the buffer resists changes in pH. The questions test your ability to:

- Predict the pH of a buffer before and after addition of acid/base.
- Calculate buffer capacity.
- Explain the behavior of the buffer in terms of equilibrium shifts.
- Interpret experimental data and reconcile it with theory.

Because the questions are often open ended, students must not only calculate correct values but also provide clear, concise explanations that demonstrate a deep understanding of the underlying chemistry.

Understanding Buffer Solutions

What Is a Buffer?

A buffer is a solution that can resist changes in pH when small amounts of strong acid or base are added. It typically consists of a weak acid and its conjugate base (or a weak base and its conjugate acid). The presence of both species allows the buffer to neutralize added H z or OH { ions by shifting the equilibrium:

1. Weak acid !Ì Conjugate base + H z
2. Conjugate base + H z !Ì Weak acid

Because the equilibrium is maintained, the pH changes only slightly, which is the defining property of a buffer.

Key Terms

- pH – a measure of the hydrogen ion concentration in a solution.
- pKa – the negative logarithm of the acid dissociation constant (Ka) for a weak acid.
- Buffer Capacity (;"' b F†R Ö÷VçB öb 7G&öær 6–B ÷" & 6R F† B 6 â &R FFVB Fò uffer before its pH changes by one unit.
- Henderson–Hasselbalch Equation – a formula that relates pH, pKa, and the ratio of concentrations of conjugate base to weak acid.

The Henderson–Hasselbalch Equation

The Henderson–Hasselbalch equation is the cornerstone of buffer calculations:

pH = pKa + log([A {]/[HA])

Where:

- pH = the solution's pH.
- pKa = the negative logarithm of the weak acid's dissociation constant.
- [A {] = concentration of the conjugate base.
- [HA] = concentration of the weak acid.

By rearranging the equation, you can solve for any of the variables if the others are known.

Using the Equation in Lab 19

In Lab 19 you often start with a buffer composed of a weak acid (e.g., acetic acid) and its conjugate base (e.g., sodium acetate). You'll be given concentrations and volumes, from which you calculate the molar concentrations of HA and A {. Plugging these values into the Henderson–Hasselbalch equation yields the initial pH. After adding a strong acid or base, you recalculate the concentrations and use the equation again to find the new pH.

Step by Step Solutions for Lab 19 Questions

Below are detailed answers to the typical questions you'll encounter in AP Chemistry Lab 19. Each answer includes the necessary calculations, the reasoning behind the steps, and a short explanation of the results.

Question 1: Calculate the Initial pH of the Buffer

Given: 0.10 /M acetic acid (CH ƒCOOH) and 0.10 /M sodium acetate (CH ƒCOONa) in 250 /mL total volume. **pKa of acetic acid = 4.76**.

Solution:

1. Since the concentrations are equal, the ratio [A {]/[HA] = 1.
2. Apply Henderson–Hasselbalch: pH = 4.76 + log(1) = 4.76 + 0 = 4.76.
3. Answer: The initial pH is 4.76.

Explanation: When the concentrations of weak acid and conjugate base are equal, the pH equals the pKa of the acid because log(1) = 0.

Question 2: Determine the pH After Adding 10 /mL of 0.10 /M HCl

Given: 10 /mL of 0.10 /M HCl added to the buffer.

Solution:

1. Calculate moles of HCl added: 0.010 /L × 0.10 /mol/L = 0.001 /mol.
2. HCl reacts with the acetate ion (CH ƒCOO {) to form acetic acid: CH ƒCOO { + H z !’ CH ƒCOOH.
3. Subtract 0.001 /mol from acetate and add it to acetic acid.
4. New moles: HA = 0.025 /mol + 0.001 /mol = 0.026 /mol; A { = 0.025 /mol – 0.001 /mol = 0.024 /mol.
5. New concentrations (total volume 0.260 /L): [HA] = 0.026/0.260 = 0.10 /M; [A {] = 0.024/0.260 "H 0.0923 /M.
6. Apply Henderson–Hasselbalch: pH = 4.76 + log(0.0923/0.10) = 4.76 + log(0.923) "H 4.76 – 0.034 = 4.726.
7. Answer: The pH after adding 10 /mL of HCl is approximately 4.73.

Explanation: The buffer neutralizes the added acid, so the pH change is small. The slight drop reflects the reduction in conjugate base concentration.

Question 3: Calculate the Buffer Capacity (;"•

Definition: ;" Ò ”n / 9G ,Â v†W&R ”n is the moles of strong acid or base added and 9G , —2 F†R &W7VÇF–ær 6† ævR –â ,

Using data from Q2:

1. 9Fâ Ò ã öÖöÂ † FFVB „6Â’à
2. 9G , Ò Bãsb b Bãs2 Ò ã 2à
3. ;" Ò ã öÖöÂ ò ã 2 $, ã 332 öÖöÂ÷ , Væ—Bà
4. Answer: The buffer capacity is about 0.033 /mol per pH unit.

Explanation: A higher buffer capacity means the solution can resist larger pH changes. For a 0.10 M buffer, 0.033 /mol per pH unit is typical.

Question 4: Predict the pH After Adding 5 /mL of 0.10 /M NaOH

Given: 5 /mL of 0.10 /M NaOH added.

Solution:

1. Moles of NaOH: 0.005 /L × 0.10 /mol/L = 0.0005 /mol.
2. NaOH reacts with acetic acid: CH ƒCOOH + OH { !’ CH ƒCOO { + H ,O.
3. Subtract 0.0005 /mol from acetic acid and add to acetate.
4. New moles: HA = 0.025 /mol – 0.0005 /mol = 0.0245 /mol; A { = 0.025 /mol + 0.0005 /mol = 0.0255 /mol.
5. New concentrations (total volume 0.255 /L): [HA] = 0.0245/0.255 "H 0.0961 /M; [A {] = 0.0255/0.255 "H 0.1000 /M.
6. Henderson–Hasselbalch: pH = 4.76 + log(0.1000/0.0961) = 4.76 + log(1.040) "H 4.76 + 0.017 = 4.777.
7. Answer: The pH after adding 5 /mL of NaOH is approximately 4.78.

Explanation: Adding a base slightly increases the pH because it converts acetic acid into acetate, shifting the equilibrium toward the base.

Question 5: Explain the Buffer’s Resistance to pH Change

Answer: The buffer’s resistance arises from the reversible reaction between the weak acid and its conjugate base. When a strong acid is added, the conjugate base consumes H z ions, forming more weak acid. Conversely, when a strong base is added, the weak acid donates H z ions, producing more conjugate base. This dynamic equilibrium keeps the concentration of H z ions—and thus the pH—relatively stable.

Common Mistakes and How to Avoid Them

- Ignoring Volume Changes – Always recalculate concentrations after adding acid or base, as the total volume changes.
- Misapplying the Henderson–Hasselbalch equation – Remember that it requires concentrations (not moles)

and that the log term uses the ratio [A {]/[HA].

- Forgetting to convert pKa to Ka or vice versa – The equation uses pKa directly; if you have Ka, take the negative log.
- Overlooking the effect of ionic strength – In most AP labs, ionic strength is negligible, but it can slightly shift pKa values.
- Failing to explain the reasoning – Calculations alone won't score full points; always provide a brief explanation of the chemistry.

Study Tips for Lab 19

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