

chapter 19 acids bases and salts answer key

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Introduction

Acids, bases, and salts are the building blocks of countless chemical reactions that shape our world. From the food we eat to the batteries that power our devices, the principles of acid–base chemistry are everywhere. If you're studying a chemistry textbook, you've probably encountered Chapter 19, which focuses on acids, bases, and salts. This article offers a thorough review of the chapter's core concepts, walks through typical problems, and provides a practical answer key to help you master the material. Whether you're a high school student tackling your first chemistry exam or a college student preparing for a midterm, the insights here will give you a solid foundation for success.

Chapter Overview: What You'll Learn

Chapter /19 typically covers the following topics:

- Definitions and Characteristics – Understanding what acids, bases, and salts are, along with their physical and chemical properties.
- Acid–Base Theories – The Arrhenius, Brønsted–Lowry, and Lewis definitions, and how they interrelate.
- pH and pOH Calculations – Using logarithmic scales to quantify acidity and basicity.
- Neutralization Reactions – Balancing equations and predicting products.
- Salts and Their Hydrolysis – How salts behave in aqueous solutions and the resulting pH.
- Buffer Systems – Designing solutions that resist pH changes.
- Stoichiometry and Titration – Practical calculations for determining concentrations.
- Applications and Safety – Real world uses of acids, bases, and salts, and lab safety considerations.

By the end of the chapter, you should be able to identify acids and bases, predict the outcomes of neutralization reactions, calculate pH values, and solve titration problems with confidence.

Key Concepts Explained

1. Acids, Bases, and Salts: Definitions

Acids are substances that donate protons (H^+) in aqueous solutions. Bases accept protons or donate hydroxide ions (OH^-). Salts are ionic compounds formed from the neutralization of an acid and a base. Understanding these definitions is crucial because they determine how each species behaves in a chemical reaction.

2. The Three Theories of Acidity and Basicity

1. Arrhenius Theory – Defines acids as H^+ donors and bases as OH^- donors in water.
2. Brønsted–Lowry Theory – Expands the concept to proton transfer, where acids donate and bases accept protons.
3. Lewis Theory – Focuses on electron pair donors and acceptors, broadening the scope to include reactions that don't involve H^+ directly.

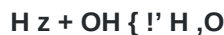
Each theory offers a different perspective but ultimately describes the same underlying chemistry.

3. pH, pOH, and the pH Scale

The pH scale ranges from 0 to 14, with 7 being neutral. Values below 7 indicate acidity, while values above 7 indicate basicity. The pOH scale is complementary, and the relationship $\text{pH} + \text{pOH} = 14$ holds at 25 °C. Calculating pH involves the negative logarithm of the hydrogen ion concentration: $\text{pH} = -\log[\text{H}^+]$

4. Neutralization Reactions

Neutralization occurs when an acid reacts with a base to form water and a salt:



The resulting salt depends on the acid and base used. For example, $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

5. Salt Hydrolysis

Salts can hydrolyze in water, producing acidic or basic solutions depending on the parent acid's strength. For instance, **NaCl** is neutral, while **NH₄Cl** forms an acidic solution because NH₄⁺ is the conjugate acid of a weak base.

6. Buffer Solutions

Buffers resist changes in pH when small amounts of acid or base are added. They typically consist of a weak acid and its conjugate base, or a weak base and its conjugate acid. The Henderson–Hasselbalch equation helps calculate buffer pH:

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

7. Titration and Stoichiometry

Titration determines the concentration of an unknown solution by reacting it with a solution of known concentration. The key steps are:

1. Choose an appropriate indicator.
2. Record the initial volume.
3. Add titrant until the endpoint.
4. Use stoichiometry to calculate the unknown concentration.

Common Problems and How to Solve Them

Problem 1: Calculate the pH of a 0.05 /M HCl Solution

Since HCl is a strong acid, it fully dissociates:

- $[\text{H}^+] = 0.05 \text{ /M}$
- $\text{pH} = -\log(0.05) \approx 1.30$

Problem 2: Find the pOH of a 0.10 /M NaOH Solution

NaOH is a strong base, fully dissociating into Na⁺ and OH⁻:

- $[\text{OH}^-] = 0.10 \text{ /M}$
- $\text{pOH} = -\log(0.10) = 1.00$
- $\text{pH} = 14 - 1.00 = 13.00$

Problem 3: Determine the pH of a 0.02 /M Acetic Acid (pK_a = 4.76)

Use the Henderson–Hasselbalch equation for weak acids:

- $\text{pH} = \text{pK}_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$
- For a pure acid, $[\text{A}^-] \approx [\text{HA}]$ (since dissociation is small). Thus, $\log(1) = 0$.
- $\text{pH} \approx 4.76$

Problem 4: Balance the Neutralization Reaction of H_2SO_4 and NaOH

Balanced equation:



Problem 5: Identify the Conjugate Base of H_2PO_4^-

Removing one H^+ from H_2PO_4^- gives HPO_4^{2-} , the conjugate base.

Sample Questions and Answers

Below are typical questions you might find in Chapter 19, along with detailed answers. These are designed to illustrate the application of concepts rather than provide direct textbook answers.

Question 1: A 0.10 M solution of HNO_3 is mixed with 0.10 M NaOH . What is the final pH?

Both reactants are strong. They neutralize completely, forming NaNO_3 and water. Since the amounts are equal, the solution is neutral.

- $\text{pH} = 7.00$

Question 2: Calculate the concentration of Na^+ in a solution formed by dissolving 10 g of NaCl in 250 mL of water.

Step 1: Moles of NaCl = mass / molar mass = $10 \text{ g} / 58.44 \text{ g/mol}$ = 0.171 mol.

Step 2: Volume = 0.250 L.

Step 3: $[\text{Na}^+] = 0.171 \text{ mol} / 0.250 \text{ L} = 0.684 \text{ M}$.

Question 3: A buffer solution contains 0.15 M H_2SO_4 and 0.10 M HSO_4^- . The pK_a of H_2SO_4 is 1.92. What is the pH?

Use Henderson–Hasselbalch:

$$\text{pH} = 1.92 + \log(0.10/0.15) = 1.92 + \log(0.666) = 1.92 - 0.176 = 1.74.$$

Question 4: A 0.05 M solution of NaOH is titrated with 0.05 M HCl . How many mL of HCl are needed to reach the endpoint?

Since concentrations are equal, the volumes must be equal.

- Volume needed = 100 mL (if starting with 100 mL of NaOH).

Question 5: Explain why NH_4OH is a weak base.

NH_3 accepts a proton to form NH_4^+ , but the equilibrium lies far to the left; only a small fraction of NH_3 molecules are protonated. Thus, NH_4OH (ammonium hydroxide) is weak.

Answer Key: A Step by Step Guide

Below is a concise answer key for common Chapter 19 questions. Use it to check your work and understand the reasoning behind each solution.

1. pH of 0.05 M HCl

$$\text{pH} = -\log(0.05) = 1.30.$$

2. pOH of 0.10 M NaOH

$\text{pOH} = -\log(0.10) = 1.00$; $\text{pH} = 13.00$.

3. pH of 0.02 /M Acetic Acid

$\text{pH} = 4.76$ (using $\text{pK}_a = 4.76$).

4. Balanced reaction of H_2SO_4 + NaOH

$\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$.

5. Conjugate base of H_3PO_4 {

HPO_4^{2-} {.

6. Final pH of equal 0.10 /M HNO_3 and NaOH

$\text{pH} = 7.00$.

7. $[\text{Na}^+]$ in 10 /g NaCl dissolved in 250 /mL water

0.684 /M.

8. pH of buffer 0.15 /M H_2SO_4 / 0.10 /M HSO_4^- {

$\text{pH} = 1.74$.

9. Volume of

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