

chapter 19 acids bases salts answers

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Chapter 19 Acids, Bases, and Salts: Answers and Insights

When you dive into the world of chemistry, one of the most pivotal chapters in any high school or introductory college textbook is Chapter 19, which focuses on acids, bases, and salts. These concepts form the backbone of countless reactions in biology, industry, and everyday life. Yet, students often find themselves puzzled by seemingly simple questions: What exactly defines an acid? How do you calculate pH? Why do some salts behave like acids or bases? This article serves as a comprehensive guide, answering the most common questions from Chapter 19 while weaving in the essential theory, practical examples, and problem solving strategies you'll need to master this material.

Why Chapter 19 Matters in Chemistry

Acids, bases, and salts are not just abstract terms; they are the language that describes how substances interact with water and with each other. From the acidity of your stomach to the corrosion of metal, from the taste of lemon to the formulation of fertilizers, the principles in Chapter 19 govern the behavior of countless systems. Understanding these concepts allows chemists to:

- Predict reaction outcomes in both laboratory and industrial settings.
- Design effective buffers for biological and pharmaceutical applications.
- Control corrosion and material degradation.
- Develop new materials such as ion exchange resins and advanced polymers.

Because of this wide-ranging impact, a solid grasp of acids, bases, and salts is essential for anyone pursuing a career in science, engineering, medicine, or environmental studies.

Key Concepts Covered in Chapter 19

Chapter 19 typically introduces the following core ideas:

1. Acid–Base Definitions – Arrhenius, Brønsted–Lowry, and Lewis theories.
2. Neutralization Reactions – How acids and bases combine to form salts and water.
3. pH, pOH, and the pH Scale – Quantifying acidity and basicity.
4. Salt Formation and Hydrolysis – How salts can act as acids or bases.
5. Buffer Systems – Maintaining stable pH in biological and chemical contexts.
6. Common Equilibrium Calculations – Using K_a , K_b , and K_w .

Below, we'll unpack each of these topics and provide clear, step by step answers to the most frequently asked questions.

1. Acid–Base Definitions: From Arrhenius to Lewis

Arrhenius Theory

Arrhenius was the first to formalize the idea of acids and bases in aqueous solutions. According to this definition:

- Acid – A substance that increases the concentration of hydrogen ions (H^+) in water.
- Base – A substance that increases the concentration of hydroxide ions (OH^-) in water.

While simple, this definition is limited because it only applies to aqueous solutions and does not account for non aqueous acids or bases.

Brønsted–Lowry Theory

Brønsted and Lowry expanded the concept by focusing on proton transfer:

- Acid – A proton donor.
- Base – A proton acceptor.

This definition works in any medium, including gas phase and non aqueous solutions.

Lewis Theory

The most general definition comes from Lewis, who defined acids and bases based on electron pair interactions:

- Lewis Acid – An electron pair acceptor.
- Lewis Base – An electron pair donor.

Under this framework, even molecules that don't participate in proton transfer (e.g., BF_3) qualify as acids.

Common Questions and Answers

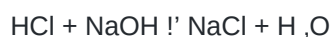
1. Q: Why do some textbooks still use Arrhenius definitions?
2. A: They are easier for beginners to grasp and provide a solid foundation before moving to more complex theories.
3. Q: Can a single compound be both an acid and a base?
4. A: Yes, amphiprotic substances (like water and bicarbonate ions) can donate and accept protons.

2. Neutralization Reactions: The Acid–Base Dance

Neutralization is the process by which an acid reacts with a base to produce a salt and water. The general equation is:

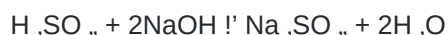


Example:



Stoichiometry of Neutralization

Neutralization reactions are typically **1:1** in terms of moles of H^+ and OH^- . However, polyprotic acids or bases may require stoichiometric adjustments. For instance, sulfuric acid (H_2SO_4) can donate two protons:



Common Questions and Answers

1. Q: What happens if the acid and base are not in stoichiometric balance?
2. A: The excess reagent determines the pH of the resulting solution.
3. Q: Can a neutralization reaction produce more than one salt?
4. A: Yes, if multiple acidic or basic sites are involved, you may get mixed salts.

3. pH, pOH, and the pH Scale

Defining pH

pH is a logarithmic measure of hydrogen ion concentration:

$$\text{pH} = -\log [\text{H}^+]$$

Similarly, pOH measures hydroxide ion concentration:

$$\text{pOH} = -\log [\text{OH}^-]$$

In water at 25 °C, the product of $[\text{H}^+]$ and $[\text{OH}^-]$ is constant:

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

pH Scale Ranges

- 0–7: Acidic (lower pH = stronger acid)
- 7: Neutral (pure water)
- 7–14: Basic (higher pH = stronger base)

Calculating pH from Concentration

For a strong acid like HCl, the concentration of H^+ equals the concentration of the acid:

$$\text{pH} = -\log (c)$$

For weak acids, you must solve the equilibrium expression using K_a .

Common Questions and Answers

1. Q: Why is pH a logarithmic scale?
2. A: Because hydrogen ion concentrations vary over many orders of magnitude; the log scale compresses this range into a manageable scale.
3. Q: Can pH be negative?
4. A: Yes, very strong acids can produce pH values below 0.

4. Salt Formation and Hydrolysis

What Makes a Salt?

A salt is an ionic compound formed by the neutralization of an acid and a base. Common examples include NaCl, CaCO_3 , and NH_4Cl .

Salt Hydrolysis

Many salts are not neutral. Their constituent ions can undergo hydrolysis, reacting with water to produce acidic or basic solutions. For example:

- $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$ (acidic)
- $\text{ClO}^- + \text{H}_2\text{O} \rightleftharpoons \text{ClO}_2^-$ (neutral)
- $\text{Na}^+ + \text{H}_2\text{O} \rightleftharpoons \text{Na}^+$ (neutral)

The extent of hydrolysis depends on the K_a and K_b values of the ions.

Common Questions and Answers

1. Q: Why does NaCl remain neutral in solution?
2. A: Both Na^+ and Cl^- are conjugate ions of a strong base (NaOH) and a strong acid (HCl), respectively; neither undergoes hydrolysis.
3. Q: How can a salt be basic?
4. A: Salts from weak acids and strong bases (e.g., NaCH_3COO) have basic cations or anions that hydrolyze to produce OH^- .

5. Buffer Systems: Keeping pH Stable

What Is a Buffer?

A buffer is a solution that resists changes in pH upon addition of small amounts of acid or base. It typically consists of a weak acid and its conjugate base or a weak base and its conjugate acid.

Buffer Capacity and Henderson–Hasselbalch Equation

The Henderson–Hasselbalch equation relates pH to the ratio of conjugate base to acid:

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

Buffer capacity (β) is a measure of how much acid or base a buffer can absorb per unit change in pH:

$$\beta = \frac{d[\text{A}^-]}{d\text{pH}}$$

Examples of Biological Buffers

- Phosphate buffer ($\text{H}_2\text{PO}_4^- / \text{HPO}_4^{2-}$)
- Bicarbonate buffer ($\text{H}_2\text{CO}_3^* / \text{HCO}_3^-$)
- Protein buffers (e.g., amino acid side chains)

Common Questions and Answers

1. Q: Why is the ratio of conjugate base to acid important?
2. A: It determines the direction and extent of proton exchange, thus controlling the pH.
3. Q: Can a buffer become saturated?
4. A: Yes, once the buffer components are consumed, the solution's pH will shift significantly.

6. Equilibrium Calculations: Ka, Kb, and Kw

Acid Dissociation Constant (Ka)

Ka quantifies the strength of a weak acid:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

A larger Ka indicates a stronger acid.

Base Dissociation Constant (Kb)

Kb quantifies the strength of a weak base:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

Relating Ka and Kb

For a conjugate acid–base pair:

$$K_a \times K_b = K_w = 1.0 \times 10^{-14}$$

Solving Weak Acid/Base Equilibria

Steps:

1. Write the dissociation equation.
2. Set up the equilibrium expression.
3. Assume initial concentration (c) and let x be the change.
4. Approximate if x ≪ c.

5. Solve for x and compute $[H^+]$ or $[OH^-]$.
6. Calculate pH or pOH.

Common Questions and Answers

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