

Unit 8: Acids and Bases

Unit 8 covers the pH scale, strong and weak acids and bases, equilibrium constants K_a and K_b , acid-base reactions and salt hydrolysis, titration curves, the molecular structures that make acids strong or weak, buffers and the Henderson-Hasselbalch equation, buffer capacity, and the effect of pH on solubility.

AP Chemistry

Acids and Bases

About 13 minutes to read

How to use this guide

Read it in order the first time because the topics build on each other. The pH scale and K_w come first, then strong acids and bases, then the weak-acid equilibrium calculations that everything else depends on. Buffers and titrations only make sense once you can predict what happens when acids and bases mix.

After the first read, use the trap boxes and the comparison tables to review the distinctions the exam tests most often. Finish with the practice questions, then complete the recall check on the last page out loud and note any items you cannot explain yet.

What this unit is worth. Acids and Bases is about 11 to 15 percent of the AP Chemistry exam. It is also one of the most calculation-heavy units, and its ideas show up inside other units: equilibrium thinking from Unit 7, solubility from Unit 8 itself, and electrochemistry later. Being fluent with K_a , pH, and buffers pays off beyond this unit.

8.1 The pH Scale and Water

The **hydronium ion**, H_3O^+ , is the form hydrogen takes in water. You will see $\text{H}^+(\text{aq})$ and $\text{H}_3\text{O}^+(\text{aq})$ used interchangeably, but H_3O^+ is preferred because a bare proton does not exist on its own in solution. The **pH** is defined as $\text{pH} = -\log[\text{H}_3\text{O}^+]$, and **pOH** as $\text{pOH} = -\log[\text{OH}^-]$. A tenfold change in ion concentration moves the pH by exactly one unit.

Water reacts with itself in **autoionization**: $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$. The **ion product of water** is $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ at 25°C . Taking the negative log of both sides gives $\text{pH} + \text{pOH} = 14$ at 25°C . A **neutral solution** is one where $[\text{H}_3\text{O}^+] = [\text{OH}^-]$, which at 25°C means $\text{pH} = \text{pOH} = 7.0$.

Trap. pH 7 is neutral only at 25°C . K_w increases with temperature because autoionization is endothermic, so at higher temperatures neutral pH drops below 7. A question that changes the temperature is testing whether you know that neutrality means $[\text{H}_3\text{O}^+] = [\text{OH}^-]$, not $\text{pH} = 7$.

8.2 Strong Acids and Strong Bases

A **strong acid** ionizes completely in water, so $[\text{H}_3\text{O}^+]$ equals the initial acid concentration. There are six to memorize: HCl, HBr, HI, HClO_4 , HNO_3 , and H_2SO_4 (first proton only). A **strong base** dissociates completely: the Group 1 hydroxides (NaOH , KOH) and the Group 2 hydroxides ($\text{Ca}(\text{OH})_2$, $\text{Ba}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$). For Group 1 bases $[\text{OH}^-]$ equals the initial concentration, but Group 2 hydroxides release two hydroxides each, so $[\text{OH}^-]$ is double.

Example: 0.010 M HCl gives $[\text{H}_3\text{O}^+] = 0.010 \text{ M}$, so $\text{pH} = 2.00$. For 0.020 M $\text{Ca}(\text{OH})_2$, $[\text{OH}^-] = 0.040 \text{ M}$, $\text{pOH} = 1.40$, and $\text{pH} = 12.60$.

Trap. H_2SO_4 is only strong in its first ionization. The second proton comes off HSO_4^- , a weak acid with $K_a = 1.2 \times 10^{-2}$. Treating 0.10 M H_2SO_4 as giving 0.20 M H_3O^+ overstates the acidity.

8.3 Weak Acids, Weak Bases, and Equilibrium Constants

A **weak acid** only partially ionizes, so $[\text{H}_3\text{O}^+]$ is much less than the initial concentration. Its strength is measured by the **acid ionization constant** $K_a = [\text{H}_3\text{O}^+][\text{A}^-]/[\text{HA}]$, with $\text{p}K_a = -\log K_a$. A **weak base** partially produces hydroxide, with **base ionization constant** $K_b = [\text{OH}^-][\text{HB}^+]/[\text{B}]$ and $\text{p}K_b = -\log K_b$. For any **conjugate acid-base pair**, $K_a \times K_b = K_w$, or equivalently $\text{p}K_a + \text{p}K_b = 14$.

To find the pH of a weak acid solution, set up the equilibrium. For 0.10 M acetic acid ($K_a = 1.8 \times 10^{-5}$): $x^2 / 0.10 \approx 1.8 \times 10^{-5}$, so $x = [\text{H}_3\text{O}^+] = 1.3 \times 10^{-3}$ M and $\text{pH} = 2.87$. The **percent ionization** is x divided by the initial concentration, here about 1.3 percent. Percent ionization rises as the acid gets more dilute, which is why K_a is the reliable measure of strength, not the percent.

Trap. K_a belongs to the acid and K_b belongs to the base. When a question gives you K_a for NH_4^+ and asks about an NH_3 solution, convert first: $K_b = K_w / K_a$. Plugging K_a directly into a base calculation is one of the most common errors in this unit.

8.4 Acid-Base Reactions and Salt Hydrolysis

When a **strong acid and strong base** mix, $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow 2\text{H}_2\text{O}$ goes quantitatively, and the pH is set by whichever reactant is in excess. Equal moles give pH 7.

Mixing a **weak acid with a strong base** also goes quantitatively: $\text{HA} + \text{OH}^- \rightarrow \text{A}^- + \text{H}_2\text{O}$. With excess weak acid remaining, you have a buffer. With equimolar amounts, only the **conjugate base** A^- remains, and it reacts with water ($\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^-$), so the solution is slightly basic. The mirror image holds for **weak base plus strong acid**: $\text{B} + \text{H}_3\text{O}^+ \rightarrow \text{HB}^+ + \text{H}_2\text{O}$, and equimolar amounts leave the conjugate acid HB^+ , giving a slightly acidic solution. This reaction of a conjugate with water is called **salt hydrolysis**, and it is the reason these mixtures are not neutral. A **weak acid plus weak base** mixture is different: $\text{HA} + \text{B} \rightleftharpoons \text{A}^- + \text{HB}^+$ reaches equilibrium rather than reacting quantitatively.

Trap. "Neutralization" does not always mean pH 7. It means the acid and base consumed each other in stoichiometric amounts. The pH at that point depends on what is left: nothing (pH 7), a conjugate base (pH above 7), or a conjugate acid (pH below 7). Read what species remains before you predict the pH.

8.5 Acid-Base Titrations

A **titration curve** plots pH against the volume of titrant added. The **equivalence point** is where moles of titrant equal moles of analyte originally present. Its pH is set by whatever species dominates there: pH 7 for strong acid with strong base, above 7 for a weak acid titrated with strong base (conjugate base hydrolyzes), and below 7 for a weak base titrated with strong acid (conjugate acid hydrolyzes).

For a weak acid or weak base titration, the **half-equivalence point** is the landmark worth memorizing.

Halfway to equivalence, $[HA] = [A^-]$, and the Henderson-Hasselbalch equation reduces to $pH = pK_a$.

Reading the pH at the half-equivalence point off a titration curve is the standard way to determine an unknown acid's pK_a . The buffer region around it is the flattest part of the curve. **Polyprotic acid** titration curves show one equivalence point per acidic proton, so H_3PO_4 gives three, and each step has its own half-equivalence pH equal to that proton's pK_a .

Trap. The equivalence point and the endpoint are not the same thing. The equivalence point is the stoichiometric reality; the endpoint is where your indicator changes color. A good titration picks an indicator whose color change brackets the equivalence-point pH. Using phenolphthalein (changes near pH 9) for a strong acid–strong base titration is fine, but for a weak base–strong acid titration with equivalence near pH 5 it would overshoot badly.

8.6 Molecular Structure and Acid/Base Strength

Acid strength can be predicted from structure. The pattern to learn is that anything stabilizing the **conjugate base** makes the acid stronger, because a stable A^- means HA gives up its proton more willingly.

- **Binary acids (H-X).** Down a group, bond strength dominates: $HI > HBr > HCl > HF$, because the H-I bond is weakest and breaks most easily. Across a period, electronegativity dominates: $HF > H_2O > NH_3 > CH_4$.
- **Oxyacids.** More oxygens mean a stronger acid: $HClO < HClO_2 < HClO_3 < HClO_4$. The electronegative oxygens pull electron density away through **inductive effects**, stabilizing the anion.
- **Resonance. Carboxylic acids** ($-COOH$) are weak acids partly because the carboxylate anion is resonance-stabilized, with the negative charge spread over two oxygens.
- **Conjugates.** Strong acids have very weak conjugate bases (Cl^- , NO_3^-), and strong bases have very weak conjugate acids. Common **weak bases** are ammonia and other **nitrogenous bases**, plus **carboxylate ions**, the conjugates of carboxylic acids.

Trap. HF looks like it should be the strongest hydrohalic acid because fluorine is the most electronegative element. It is actually the weakest, because the H-F bond is so strong that HF barely ionizes. Down a group, bond strength beats electronegativity. Check which trend applies before you rank.

8.7 pH, pK_a , and Indicators

The relationship between **pH and pK_a** tells you which form of an acid predominates. When the solution pH is below the acid's pK_a , the protonated form HA predominates. When pH is above pK_a , the deprotonated form A^- predominates. At $pH = pK_a$ they are equal. This is the logic behind both buffers and titrations: moving the pH past a pK_a converts one form into the other.

An **acid-base indicator** is itself a weak acid whose protonated and deprotonated forms have different colors. It changes color over a range of about $pK_a \pm 1$. Choose an indicator whose pK_a sits near the titration's equivalence-point pH: phenolphthalein (around pH 8–10) for weak acid–strong base titrations, methyl orange (around pH 3–4) for strong acid–weak base titrations.

Trap. An indicator does not change color at exactly one pH. It changes across a range, roughly two pH units wide. A question asking for the best indicator wants the one whose range brackets the equivalence point, not one whose nominal pK_a merely touches it.

8.8 Buffers

A **buffer solution** contains large concentrations of both members of a conjugate acid-base pair, such as CH_3COOH and CH_3COO^- . Added base is consumed by the conjugate acid ($HA + OH^- \rightarrow A^- + H_2O$) and added acid is consumed by the conjugate base ($A^- + H_3O^+ \rightarrow HA + H_2O$), so the pH barely moves. A buffer needs both members present in appreciable amounts: a solution of HA alone, or A^- alone, is not a buffer.

8.9 The Henderson-Hasselbalch Equation

The **Henderson-Hasselbalch equation** gives a buffer's pH directly: $\text{pH} = \text{pK}_a + \log([\text{A}^-]/[\text{HA}])$. It comes from taking the negative log of the K_a expression and rearranging. When $[\text{A}^-] = [\text{HA}]$, the log term is zero and $\text{pH} = \text{pK}_a$, which is why equal-concentration buffers sit exactly at the acid's pK_a .

Example: a buffer with 0.20 M acetate and 0.10 M acetic acid ($\text{pK}_a = 4.74$) has $\text{pH} = 4.74 + \log(0.20/0.10) = 4.74 + 0.30 = 5.04$. A buffer resists pH change best near its pK_a , within about one pH unit either way. To prepare a buffer at a target pH, pick an acid whose pK_a is close to the target, then set the ratio to fine-tune.

Trap. Henderson-Hasselbalch only applies when both HA and A^- are present in significant amounts. Using it on a pure weak acid solution, or after enough strong base has been added to consume all the HA, gives a wrong answer. Check that you actually have a buffer before you reach for the equation.

8.10 Buffer Capacity

Buffer capacity is how much added acid or base a buffer can absorb before its pH shifts significantly. Raising the concentrations of both components while keeping the ratio fixed leaves the pH unchanged but increases capacity, because there are more moles of each member to do the neutralizing. A buffer with extra conjugate acid resists added base better, and one with extra conjugate base resists added acid better. Diluting a buffer does not change its pH (the ratio is unchanged) but it does lower its capacity.

8.11 pH and Solubility

A salt's solubility depends on pH when one of its ions comes from a weak acid, a weak base, or hydroxide. The reasoning is Le Châtelier's principle: lowering the pH protonates a basic anion and removes it from solution, pulling more solid into solution. Calcium carbonate dissolves better in acid because $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ consumes the carbonate. Salts of strong-acid anions, like AgCl, show no pH effect because Cl^- is too weak a base to be protonated.

Trap. "More soluble in acid" applies only to salts with basic anions (carbonates, sulfides, hydroxides, phosphates). A salt like BaSO_4 barely responds to pH because HSO_4^- is a relatively strong acid. Identify the anion's conjugate strength before you predict.

Confusions That Cost Points

PAIR	HOW TO KEEP THEM STRAIGHT
K_a vs K_b	K_a describes the acid ionizing; K_b describes the base. They are linked by $K_a \times K_b = K_w$ for a conjugate pair, but they are not interchangeable in a calculation.
Equivalence point vs half-equivalence point	Equivalence: moles titrant equal moles analyte; pH set by the remaining species. Half-equivalence: $[HA] = [A^-]$ and $pH = pK_a$, the flattest part of the curve.
Buffer pH vs buffer capacity	The ratio $[A^-]/[HA]$ sets the pH; the absolute concentrations set the capacity. Same ratio at higher concentration means the same pH with more capacity.
Strong acid vs concentrated acid	Strong means fully ionized (a K_a property). Concentrated means many moles per liter. A dilute strong acid can have a higher pH than a concentrated weak acid.
pH 7 vs neutral	Neutral means $[H_3O^+] = [OH^-]$. That gives pH 7 only at 25°C. At other temperatures K_w shifts and neutral pH moves with it.

Practice Questions

Original questions written for this guide in the style of the AP exam. Answers and explanations are on the next page, so complete the questions before checking them.

1. What is the pH of a 0.010 M solution of HNO_3 ?

- A. 1.00
- B. 2.00
- C. 3.00
- D. 12.00

2. A buffer must hold pH near 4.74. Which conjugate pair is the best choice?

- A. NH_4^+ / NH_3 ($pK_a = 9.26$)
- B. CH_3COOH / CH_3COO^- ($pK_a = 4.74$)
- C. H_2CO_3 / HCO_3^- ($pK_{a1} = 6.37$)
- D. HCl / Cl^-

3. 25.0 mL of 0.100 M acetic acid is titrated with 0.100 M NaOH. What is the pH at the equivalence point?

- A. 7.00
- B. Less than 7
- C. Greater than 7
- D. 4.74

4. Which salt shows the greatest increase in solubility as the pH is lowered?

- A. AgCl
- B. $BaSO_4$
- C. $CaCO_3$
- D. NaCl

Answer Key

- 1. B.** HNO_3 is a strong acid, so it ionizes completely: $[\text{H}_3\text{O}^+] = 0.010 \text{ M}$ and $\text{pH} = -\log(0.010) = 2.00$. A confuses pH with pOH arithmetic. C would be the pH of a 0.0010 M solution. D is the pH of 0.010 M NaOH , the strong base with the same concentration.
- 2. B.** A buffer works best within about one pH unit of the acid's pK_a , so match pK_a to the target pH. Acetic acid's pK_a of 4.74 is exactly the target. A buffers near pH 9.26, far from the target. C buffers near 6.37, outside the useful range. D is not a buffer at all: Cl^- is the conjugate of a strong acid and does not hydrolyze.
- 3. C.** At equivalence all the acetic acid has become acetate, CH_3COO^- , which hydrolyzes water to produce OH^- . The solution is basic, so pH is greater than 7. A assumes every neutralization lands at pH 7, which is true only for strong acid with strong base. B describes a weak base titrated with strong acid. D is the pH at the half-equivalence point, where $[\text{HA}] = [\text{A}^-]$.
- 4. C.** CaCO_3 has the basic anion CO_3^{2-} , which is protonated in acid ($\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$), pulling more solid into solution by Le Châtelier's principle. A and D have anions from strong acids (Cl^-) that are not protonated. B responds only slightly because HSO_4^- is a relatively strong acid.

One-Page Recall Check

- Write the definitions of pH and pOH, and derive $\text{pH} + \text{pOH} = 14$ from K_w .
- Explain why pH 7 is neutral only at 25°C.
- List the six strong acids and the strong bases, and state the $[\text{OH}^-]$ for 0.020 M $\text{Ba}(\text{OH})_2$.
- Write the K_a expression for a generic weak acid and calculate the pH of 0.10 M acetic acid ($K_a = 1.8 \times 10^{-5}$).
- State the $K_a \times K_b = K_w$ relationship and use it to find K_b for acetate.
- Predict whether the equivalence point is acidic, basic, or neutral for each of the four acid-base mixture types.
- Explain salt hydrolysis and why a solution of NaCH_3COO is basic.
- Sketch the three titration curve shapes and label the equivalence point, half-equivalence point, and buffer region on each.
- State what the pH equals at the half-equivalence point of a weak acid titration and why.
- Rank HF, HCl, HBr, HI by acid strength and justify the order.
- Explain why adding oxygens strengthens an oxyacid.
- State the Henderson-Hasselbalch equation and use it to find the pH of a 0.20 M / 0.10 M acetate buffer.
- Explain the difference between buffer pH and buffer capacity.
- Predict which salts become more soluble in acid and explain why with Le Châtelier's principle.

Study this unit in Rycal. Drill the 33 key terms for Acids and Bases at rycal.web.app/apchem, and work the AP-style questions there. If you have a test date, add it in the Test Planner. You can also start your next review with a Brain Dump, then check what you missed against this guide.

Key terms for this unit

pH, pOH, Hydronium ion (H_3O^+), Autoionization of water, Ion product of water (K_w), Neutral solution, Strong acid, Strong base, Weak acid, Acid ionization constant (K_a), Weak base, Base ionization constant (K_b), Percent ionization, $K_a \times K_b = K_w$, Strong acid–strong base neutralization, Weak acid + strong base mixtures, Weak base + strong acid mixtures, Weak acid + weak base mixtures, Salt hydrolysis, Titration curve, Equivalence point (acid-base titration), Half-equivalence point, Polyprotic acid titrations, Structural factors in acid/base strength, Carboxylic acids, Nitrogenous bases and carboxylate ions, Weak conjugates of strong acids and bases, pH vs. $\text{p}K_a$ (predominant form), Acid-base indicator, Buffer solution, Henderson-Hasselbalch equation, Buffer capacity, pH effect on solubility.

About this guide. Written for Rycal and aligned to the College Board AP Chemistry course framework, Unit 8. All questions and explanations are original Rycal writing. Rycal is independent and is not affiliated with or endorsed by the College Board.