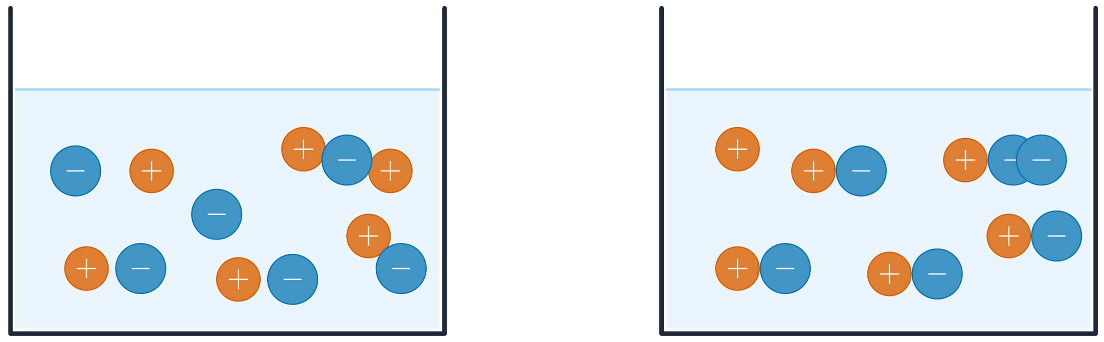
Acids and Bases

AP Chemistry

Introduction to Acids and Bases

By the **Brønsted–Lowry** definition, an **acid** 酸 is a proton (H^+) **donor** and a **base** 碱 is a proton **acceptor**. When an acid donates a proton it becomes its **conjugate base** 共轭碱; a base gaining a proton becomes its **conjugate acid** 共轭酸. Water is **amphoteric** –it can act as either.



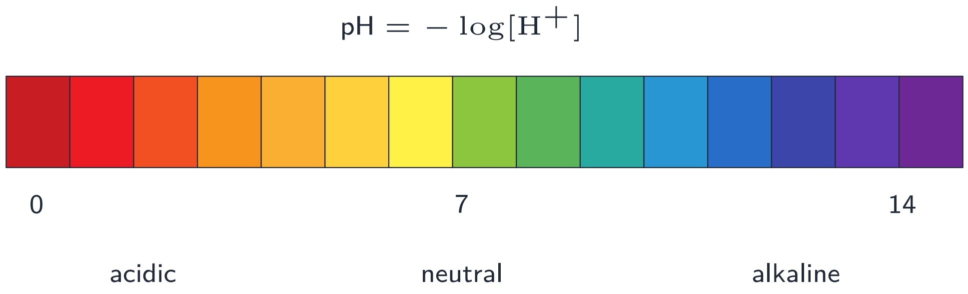
strong acid
fully dissociated: many ions

weak acid
mostly molecules: few ions

A strong acid is fully dissociated; a weak acid is only partly dissociated

pH and pOH of Strong Acids and Bases

The **pH** pH 值 scale measures acidity: $\mathrm{pH} = -\log[\mathrm{H}^+]$, with $\mathrm{pH} + \mathrm{pOH} = 14$ at 25 °C. A **strong acid** 强酸 or **strong base** 强碱 dissociates **completely**, so its ion concentration equals its concentration –read pH directly. Lower pH means more acidic.



The pH scale: pH is the negative log of the hydrogen-ion concentration

Worked example. Find the pH of 0.010 M HCl. Because HCl is a strong acid it is fully dissociated, so $[H^+] = 0.010$ M and

$$\text{pH} = -\log(0.010) = 2.0.$$

For a strong **base** like 0.010 M NaOH, pOH $= 2.0$, so pH $= 14 - 2.0 = 12.0$.

Weak Acid and Base Equilibria

A **weak acid** 弱酸 only **partly** dissociates, described by an equilibrium constant K_a (larger K_a = stronger weak acid); a weak base has K_b. Because dissociation is small, find $[H^+]$ with an ICE table and the K_a expression, often using the small-x approximation.

Worked example. Find the pH of 0.10 M acetic acid, $K_a = 1.8 \times 10^{-5}$. Let $x = [H^+]$ at equilibrium; the ICE table gives $K_a = \dfrac{x^2}{0.10 - x} \approx \dfrac{x^2}{0.10}$ (small-x approximation), so

$$x = \sqrt{K_a \times 0.10} = \sqrt{1.8 \times 10^{-6}} = 1.3 \times 10^{-3}\ \text{M} \ \Rightarrow\ \text{pH} = -\log(1.3 \times 10^{-3}) = 2.9.$$

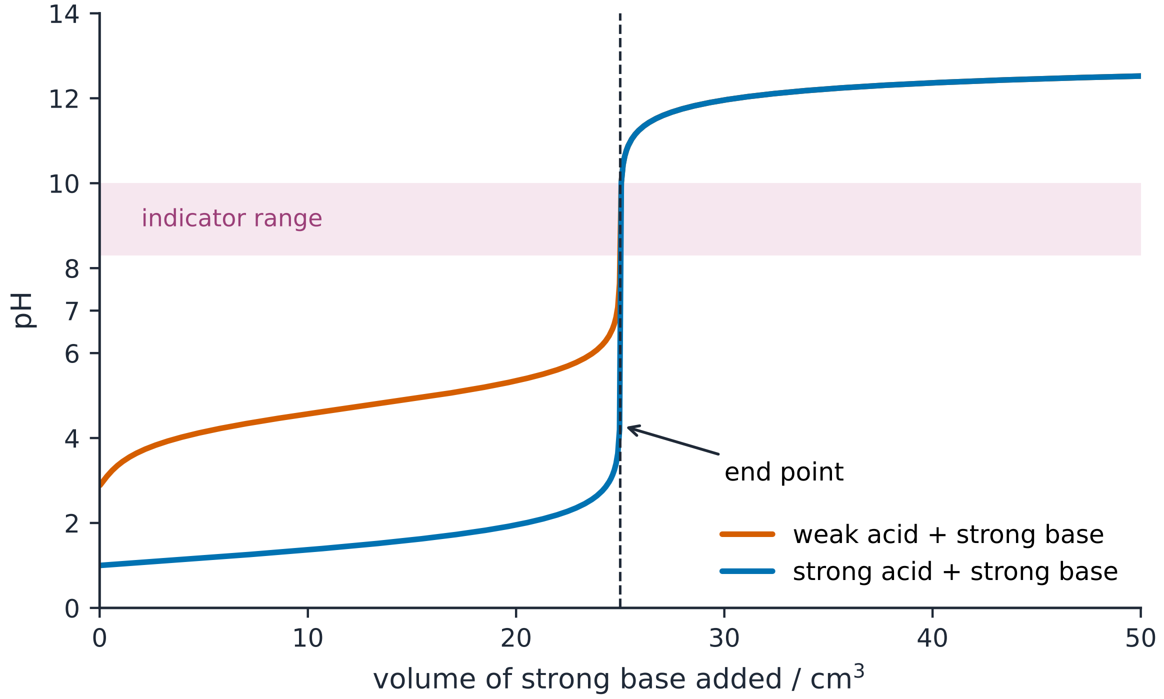
The weak acid is far less acidic than a strong acid of the same concentration (which would be pH 1.0).

Acid-Base Reactions and Buffers

A **buffer** 缓冲溶液 resists pH change. It contains a weak acid **and** its conjugate base (or a weak base and its conjugate acid) in comparable amounts. Added acid is neutralized by the conjugate base, and added base by the weak acid, so pH barely moves.

Acid-Base Titrations

A **titration curve** 滴定曲线 plots pH as base (or acid) is added. Key points: the **equivalence point** 等当点 (moles of acid = moles of base –a steep jump), and the **half-equivalence point**, where pH $=$ pK_a (half the weak acid is converted, so a buffer is at its center).



A titration curve has a steep jump near the equivalence point

Molecular Structure of Acids and Bases

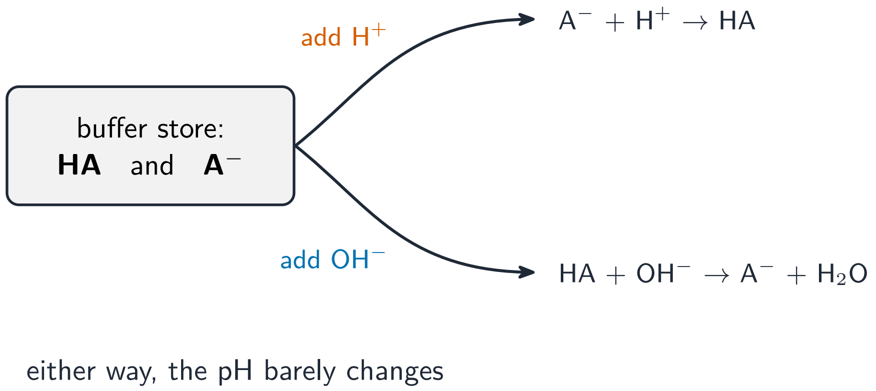
Strength has structural roots. An acid is stronger when its **conjugate base is more stable** –for example, more electronegative atoms or resonance spreading the negative charge stabilize it. For oxoacids, more oxygen atoms on the central atom means a stronger acid.

pH and pKa

$pK_a = -\log K_a$: a **smaller** pK_a means a **stronger** acid. Comparing pH to pK_a tells you the dominant form: below pK_a the acid form dominates; above it, the conjugate base form dominates.

Properties of Buffers

A buffer works best when the weak acid and conjugate base concentrations are **similar** (pH near pK_a). Choose a buffer whose pK_a is close to the target pH. Diluting a buffer barely changes its pH, because the acid-to-base **ratio** stays the same.



A buffer resists pH change by mopping up added acid or base

The Henderson-Hasselbalch Equation

For a buffer, the pH follows from the acid-to-base ratio:

$$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\mathrm{A}^-]}{[\mathrm{HA}]}.$$

When the concentrations are equal, the log is zero and pH $=$ pK_a. Use it to design a buffer or find its pH quickly.

Worked example. A buffer holds 0.20 M acetic acid (p$K_a = 4.74$) and 0.30 M acetate. Its pH is

$$\mathrm{pH} = 4.74 + \log \frac{0.30}{0.20} = 4.74 + \log(1.5) = 4.74 + 0.18 = 4.92.$$

A little more conjugate base than acid pushes the pH just above pK_a, as the equation predicts.

Buffer Capacity

Buffer capacity 缓冲容量 is how much acid or base a buffer can absorb before its pH changes sharply. It is greatest when the components are **concentrated** and in roughly **equal** amounts. Once one component is used up, the buffer fails.

pH and Solubility

The solubility of a salt with a basic anion **rises** in acidic solution: the added H^+ reacts with the anion, removing it from the solubility equilibrium and pulling more solid to dissolve (Le Chatelier applied to K_{sp}). So pH can control whether an ionic solid dissolves.

Exam tips

- $\mathrm{pH} = -\log[\mathrm{H}^+]$ is **logarithmic** —each unit is a ten-fold change in $[\mathrm{H}^+]$.

- A **strong** acid/base dissociates completely (so $[H^+]$ = concentration); a **weak** one only partly, needing K_a.
- A **buffer** contains a weak acid **and** its conjugate base together; it resists pH change by mopping up added acid or base.
- On a titration curve the **equivalence point** is the steep jump; the **half-equivalence point** has $\text{pH} = \text{p}K_a$.
- A smaller $\text{p}K_a$ means a stronger acid.