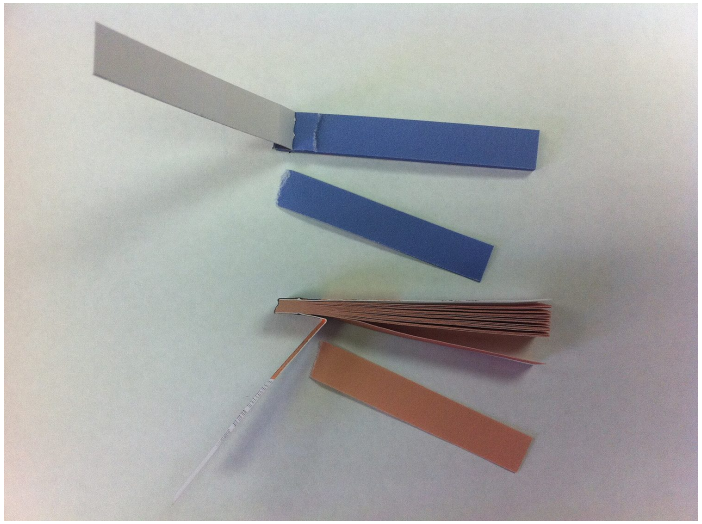
Acids and Bases

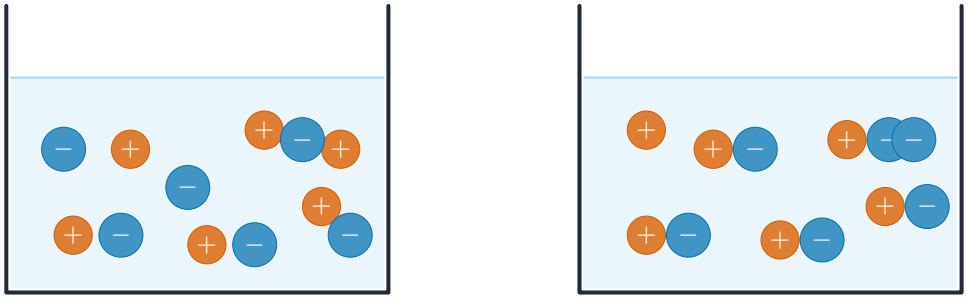
AP Chemistry

Introduction to Acids and Bases



Litmus paper: acid–base indicators change colour with H^+ concentration (pH)

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



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

pH test strips: acid–base strength and concentration set the pH you measure

The Autoionization of Water

Even pure water conducts a tiny current, because water molecules react with each other in a reaction called **autoionization**:

$$2H_2O \rightleftharpoons H_3O^+ + OH^-.$$

One water molecule passes a proton to another, making a **hydronium ion** (H_3O^+, the ion we loosely write as H^+) and a hydroxide ion. This equilibrium has its own constant, the **ion-product constant of water**:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

at 25 °C. Defining $\text{pOH} = -\log[OH^-]$ and taking $-\log$ of the K_w expression gives the rule the rest of this topic leans on:

$$\text{pH} + \text{pOH} = \text{p}K_w = 14.0$$

(again at 25 °C). Know one of pH or pOH and you get the other by subtracting from 14.

In pure water the two ions form in equal numbers, so $[H_3O^+] = [OH^-]$ and $\text{pH} = \text{pOH} = 7.0$. This is what **neutral** means. Acidic means $[H_3O^+] > [OH^-]$ (pH below 7); basic is the reverse.

Worked example. A solution has $[H_3O^+] = 2.0 \times 10^{-3}$ M. Find $[OH^-]$. Because K_w always holds in water,

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} = 5.0 \times 10^{-12} \text{ M}.$$

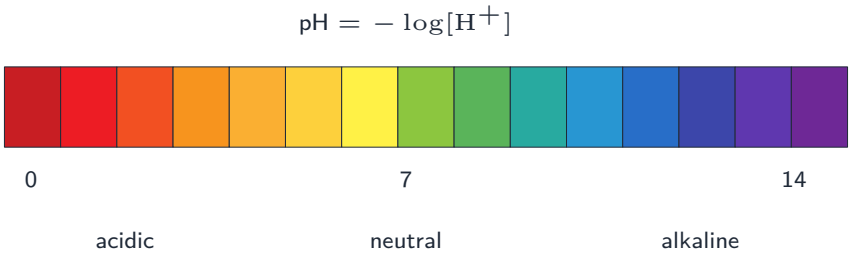
The hydroxide concentration is far smaller than the hydronium, confirming the solution is acidic.

K_w is **temperature dependent**: autoionization is endothermic, so heating water shifts it forward and raises K_w. At 50 °C, $K_w > 1.0 \times 10^{-14}$, so neutral water has $\text{pH} = \text{pOH} < 7.0$. It is still neutral, because the ion concentrations are still equal – neutral means equal ions, not a pH of exactly 7.

pH strips: water autoionizes so pure water is pH 7; acids and bases shift H+ and OH-

pH and pOH of Strong Acids and Bases

The **pH** pH scale measures acidity: $\text{pH} = -\log[H^+]$, with $\text{pH} + \text{pOH} = 14$ at 25 °C. A **strong acid** or **strong base** 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, $\text{pOH} = 2.0$, so $\text{pH} = 14 - 2.0 = 12.0$.

Weak Acid and Base Equilibria

A **weak acid** only **partly** dissociates, described by an equilibrium constant K_a (larger K_a = stronger weak acid); a weak base has K_b. For any conjugate acid–base pair the two are linked through water, $K_a \times K_b = K_w$, so $\text{p}K_a + \text{p}K_b = 14$ – a stronger acid always has a weaker conjugate base. 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).

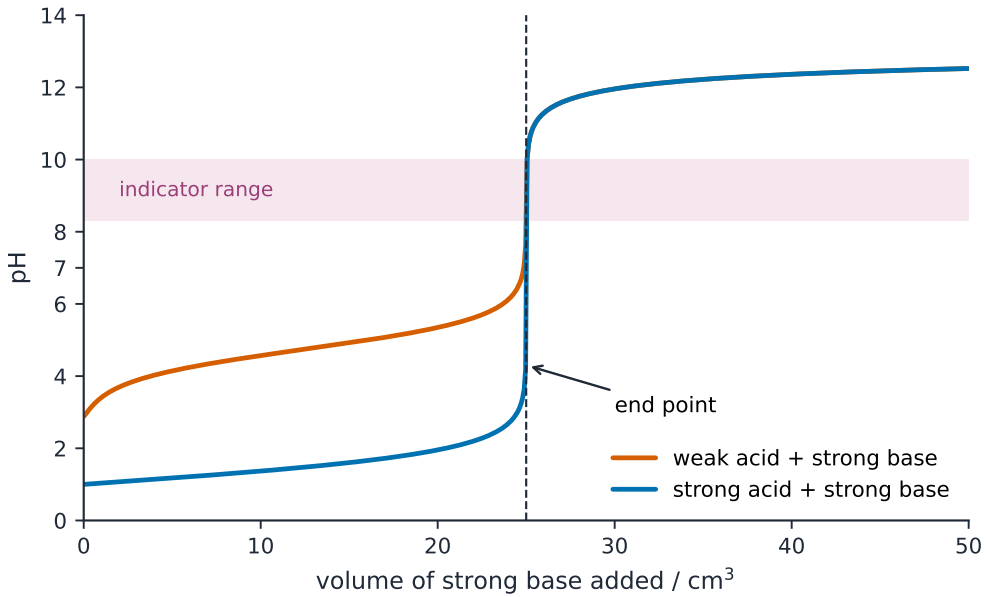
Acid-Base Reactions and Buffers

A **buffer** 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** plots pH as base (or acid) is added. Key points: the **equivalence point** (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).

An **acid-base indicator** is itself a weak acid whose protonated and deprotonated forms have **different colours**, so its colour responds to pH. You pick an indicator whose colour change happens near the titration's equivalence-point pH, so the colour flips just as the reaction completes.



A titration curve has a steep jump near the equivalence point

Molecular Structure of Acids and Bases

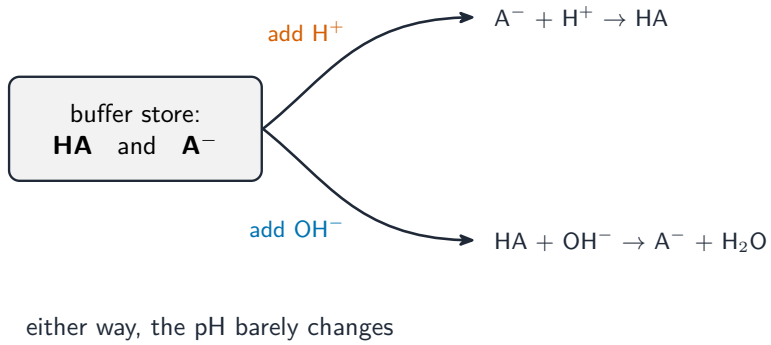
Strength has structural roots. An acid is stronger when its **conjugate base is more stable** – for example, more electronegative atoms or resonance spreading the negative charge stabilize it. For oxoacids, more oxygen atoms on the central atom means a stronger acid.

pH and pKa

p$K_a = -\log K_a$: 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 (pK_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 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

- **Neutral means equal ions, not pH 7.** In pure water $[H_3O^+] = [OH^-]$; since K_w rises with temperature, neutral pH drifts below 7 above 25 °C. Use $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ (25 °C) to convert between $[OH^-]$ and $[H_3O^+]$.
- $\mathrm{pH} = -\log[H^+]$ is **logarithmic** – each unit is a ten-fold change in $[H^+]$.
- 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 $\mathrm{pH} = \mathrm{p}K_a$.
- A smaller $\mathrm{p}K_a$ means a stronger acid.