

Strong and weak acids, pH and buffer solutions

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Lesson

Quick answer: A Brønsted-Lowry acid donates a proton and a base accepts one. Strong acids ionise fully, so $\text{pH} = -\log[\text{H}^+]$; weak acids ionise partly, so $[\text{H}^+] \approx \sqrt{(\text{K}_a \times c)}$. A buffer, made of a weak acid and its conjugate base, resists pH change, and its pH is $\text{pK}_a + \log([\text{A}^-]/[\text{HA}])$.

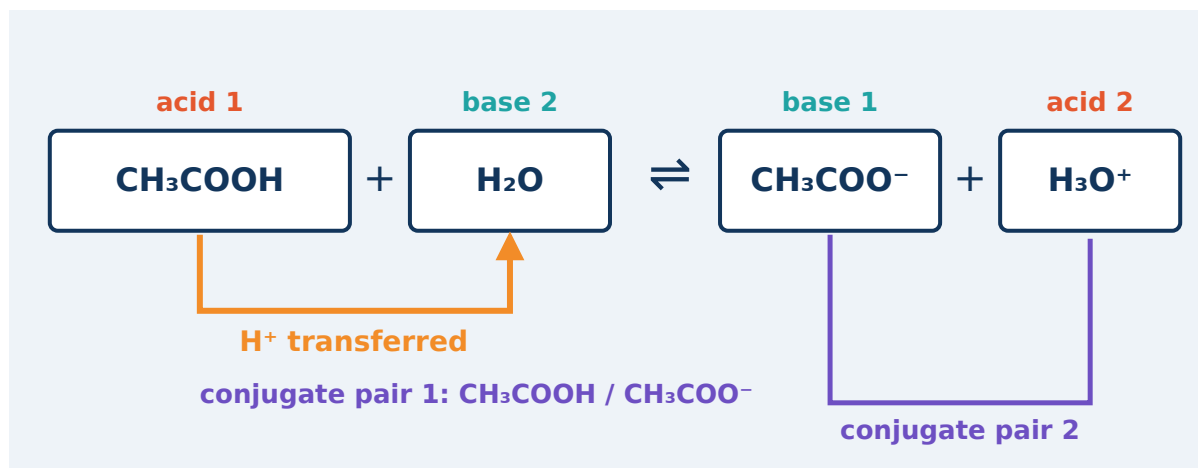
Level: Chemistry grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB R3.1

What do acid, base and conjugate pair mean?

Chemists describe acid-base reactions as proton transfer. A proton here is an H^+ ion, which in water is attached to a water molecule as H_3O^+ and is usually written simply as $\text{H}^+(\text{aq})$.

Definition

A Brønsted-Lowry acid is a proton donor. A Brønsted-Lowry base is a proton acceptor. When an acid loses H^+ it becomes its conjugate base; when a base gains H^+ it becomes its conjugate acid.



The acid loses H^+ and becomes its conjugate base; the base gains H^+ and becomes its conjugate acid.

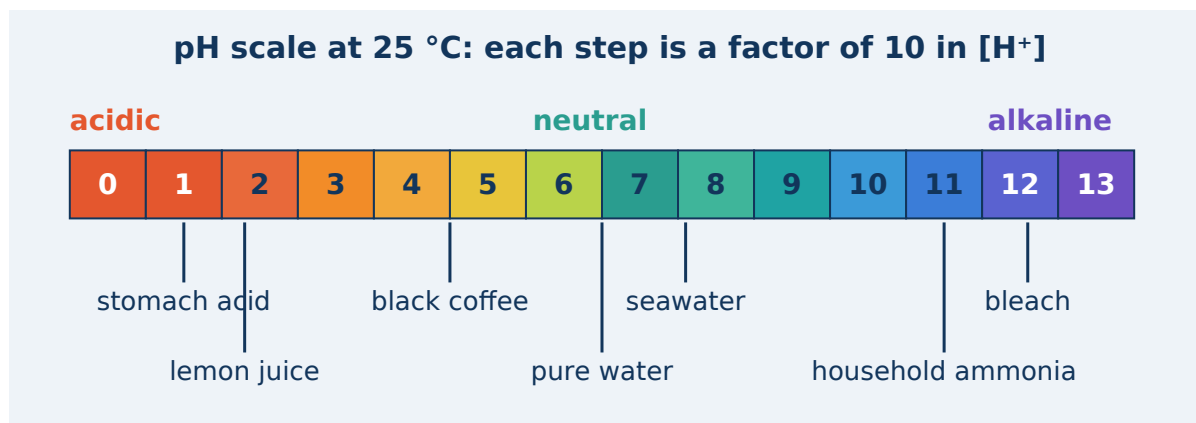
Strong acids such as HCl and HNO_3 ionise almost completely in water. Weak acids such as ethanoic acid ionise only slightly, so most of the acid stays as molecules in an equilibrium: $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$.

How do you calculate pH and pOH?

The pH scale turns very small concentrations into manageable numbers. Water itself ionises a little, and its equilibrium constant is the ionic product of water, K_w .

Formulae

$\text{pH} = -\log_{10}[\text{H}^+]$ and $[\text{H}^+] = 10^{-\text{pH}}$
 $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ mol}^2/\text{dm}^6$ at 25°C
 $\text{pH} + \text{pOH} = 14$ at 25°C
 $[\text{H}^+]$ is the hydrogen ion concentration in mol/dm^3 .



A change of one pH unit means $[\text{H}^+]$ changes tenfold.

Worked example 1

Aarav in Pune dilutes hydrochloric acid until it is $0.0080 \text{ mol}/\text{dm}^3$. Find the pH.
 HCl is strong, so $[\text{H}^+] = 0.0080 \text{ mol}/\text{dm}^3$.
 $\text{pH} = -\log(0.0080) = \mathbf{2.10}$

How do you find the pH of a weak acid?

A weak acid HA sets up the equilibrium $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$. Its strength is measured by the acid dissociation constant K_a .

Formulae

$K_a = [\text{H}^+][\text{A}^-] / [\text{HA}]$ and $\text{p}K_a = -\log_{10} K_a$
 For a weak acid of concentration c : $[\text{H}^+] \approx \sqrt{(K_a \times c)}$
 K_a is in mol/dm^3 ; c is the starting concentration in mol/dm^3 .

The shortcut assumes two things: the acid ionises so little that $[\text{HA}] \approx c$, and $[\text{H}^+] = [\text{A}^-]$ because water's own H^+ is negligible. It is reliable when less than about 5 % of the acid has ionised. A smaller $\text{p}K_a$ means a stronger acid.

Acid	$K_a (\text{mol}/\text{dm}^3)$	$\text{p}K_a$
Chloroethanoic acid	1.4×10^{-3}	2.85
Methanoic acid	1.6×10^{-4}	3.80
Benzoic acid	6.3×10^{-5}	4.20
Ethanoic acid	1.8×10^{-5}	4.74
Hydrocyanic acid	4.9×10^{-10}	9.31

Worked example 2

Find the pH of 0.100 mol/dm³ ethanoic acid ($K_a = 1.8 \times 10^{-5}$ mol/dm³).

$$[H^+] = \sqrt{(1.8 \times 10^{-5} \times 0.100)} = \sqrt{(1.8 \times 10^{-6})} = 1.34 \times 10^{-3} \text{ mol/dm}^3$$

$$\text{pH} = -\log(1.34 \times 10^{-3}) = \mathbf{2.87}$$

Check: $1.34 \times 10^{-3} \div 0.100 = 1.3\%$ ionised, so the approximation holds.

What is a buffer and how does it work?

Definition

A buffer is a solution that resists a change in pH when small amounts of acid or alkali are added. It contains a weak acid HA together with a large amount of its conjugate base A⁻ (or a weak base with its conjugate acid).

Added H⁺ is removed by A⁻, forming HA. Added OH⁻ is removed by HA, forming A⁻ and water. Because both HA and A⁻ are plentiful, their ratio changes only slightly, so the pH hardly moves. The pH of a buffer follows the Henderson-Hasselbalch equation.

Formula

$$\text{pH} = \text{p}K_a + \log_{10}([A^-] / [HA])$$

[A⁻] and [HA] are the concentrations of the conjugate base and the weak acid in mol/dm³.

In one solution the mole ratio can be used instead.

Buffer: HA and A⁻ both present

HA reservoir (neutralises OH⁻)

A⁻ reservoir (neutralises H⁺)

$$\text{pH} = \text{p}K_a + \log ([A^-] / [HA])$$

add H⁺: A⁻ + H⁺ → HA

add OH⁻: HA + OH⁻ → A⁻ + H₂O

Each reservoir uses up one kind of added ion, so the ratio [A⁻]/[HA] barely moves.

Worked example 3

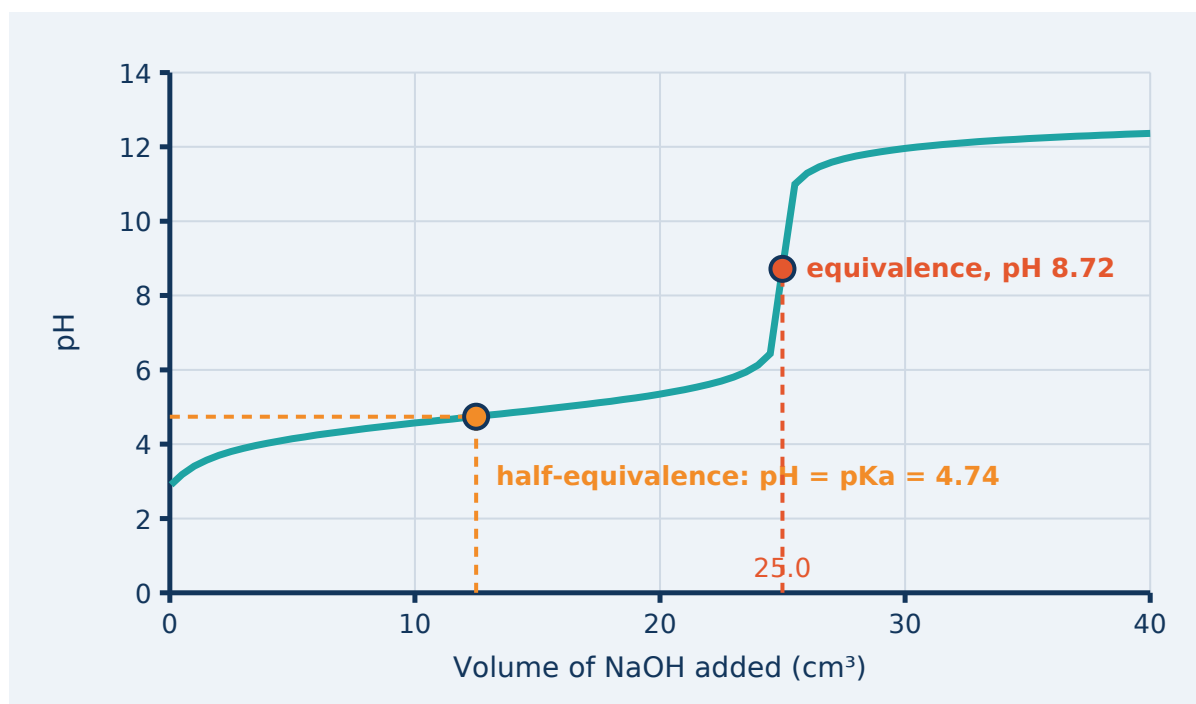
A buffer holds 0.20 mol/dm³ ethanoic acid and 0.10 mol/dm³ sodium ethanoate.

$$\text{pH} = 4.74 + \log(0.10 \div 0.20) = 4.74 - 0.30 = \mathbf{4.44}$$

What does a titration curve of a weak acid show?

When a weak acid is titrated with a strong alkali, the curve begins at a higher pH than a strong acid of the same concentration. A buffer region appears while HA and A⁻ coexist. Exactly half-

way to the equivalence point, $[HA] = [A^-]$, so $\log(1) = 0$ and $pH = pK_a$. At equivalence only the weak base A^- remains, so the pH is above 7.



Weak acid against strong base: half-equivalence gives pK_a and the equivalence point lies above pH 7.

Common mistake

Do not use $[H^+] = \sqrt{(K_a \times c)}$ for a strong acid, and do not trust it for very dilute solutions. Also, a buffer needs both HA and A^- present: adding exactly equal moles of a weak acid and a strong alkali gives only a salt solution, not a buffer.

Lina's tip

Whenever a question says half-neutralised, equal moles of acid and conjugate base, or half-equivalence point, write $pH = pK_a$ straight away. The logarithm term is zero.

Key points to remember

- $pH = -\log_{10}[H^+]$; at 25 °C $pH + pOH = 14$ and $K_w = 1.0 \times 10^{-14}$.
- Weak acid: $K_a = \frac{[H^+][A^-]}{[HA]}$; $[H^+] \approx \sqrt{(K_a \times c)}$ when ionisation is under about 5 %.
- Buffer: $pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$; A^- removes added H^+ and HA removes added OH^- .
- Half-equivalence point of a weak acid titration: $pH = pK_a$.
- At the equivalence point of a weak acid with a strong base the pH is above 7.

Strong and weak acids, pH and buffer solutions: frequently asked questions

What is the difference between a strong and a weak acid?

A strong acid ionises completely in water, so $[H^+]$ equals its concentration. A weak acid ionises only partly and reaches an equilibrium described by K_a . Strength is about how much it ionises, not how concentrated it is.

How do you calculate the pH of a weak acid?

Write $K_a = [H^+][A^-]/[HA]$, assume $[H^+] = [A^-]$ and $[HA] \approx c$, so $[H^+] = \sqrt{K_a \times c}$. Then $pH = -\log[H^+]$. Check that less than 5 % of the acid has ionised.

How does a buffer resist pH change?

It contains a weak acid and its conjugate base. The base reacts with added H^+ and the acid reacts with added OH^- , so the ratio $[A^-]/[HA]$ changes only a little and the pH stays almost constant.

What is the Henderson-Hasselbalch equation?

It is $pH = pK_a + \log([A^-]/[HA])$. It gives the pH of a buffer from the pK_a of the weak acid and the ratio of conjugate base to acid. It works best for ratios between 0.1 and 10.

Why is the equivalence point of a weak acid titration above pH 7?

At equivalence all the weak acid has become its conjugate base A^- . A^- is a weak base that reacts with water to form OH^- , so the solution is slightly alkaline.