

Strong and weak acids, bases and pH

AP Chemistry · Lesson

Quick answer: A strong acid ionises completely, so its pH comes straight from its concentration. A weak acid ionises only partly and is described by K_a : for $HA \rightleftharpoons H^+ + A^-$, $K_a = \frac{[H^+][A^-]}{[HA]}$. Solve for $[H^+]$ with an ICE table, then take $pH = -\log[H^+]$. Conjugate pairs obey $K_a \times K_b = K_w$.

Level: AP Chemistry · **Curriculum:** AP Chem Unit 8

What is pH and how do you calculate it?

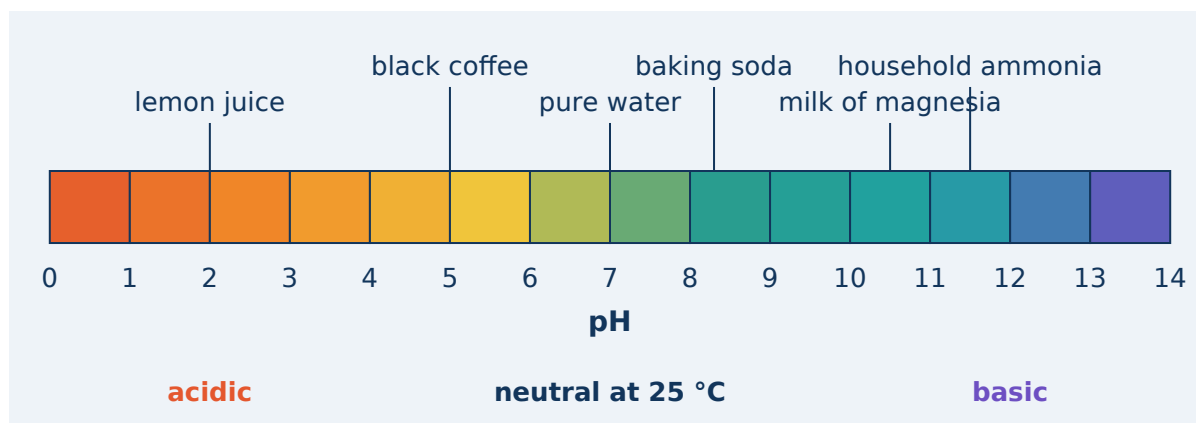
Definition

pH = $-\log_{10}[H_3O^+]$, where $[H_3O^+]$ is the hydronium concentration in mol/dm³. A Brønsted-Lowry **acid** donates a proton and a **base** accepts one. When an acid HA loses a proton, the leftover A^- is its **conjugate base**.

Water ionises slightly, $2H_2O \rightleftharpoons H_3O^+ + OH^-$, with an equilibrium constant called the ion-product of water. At 25 °C it is fixed, so pH and pOH always add to 14.

Formulas

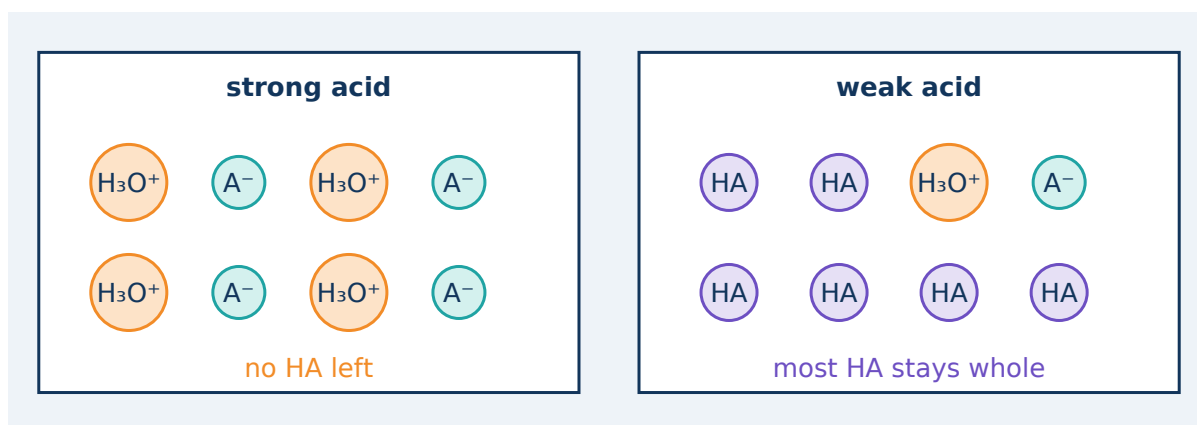
$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C; $pH + pOH = 14.00$; $pH = -\log_{10}[H_3O^+]$; $[H_3O^+] = 10^{-pH}$. Concentrations are in mol/dm³; K_w and K_a have no unit in AP work.



The pH scale at 25 °C. Each step down the scale means ten times more H_3O^+ .

How do strong and weak acids differ?

A strong acid ionises completely, so $[H_3O^+]$ equals the acid concentration (for a monoprotic acid). A weak acid ionises only partly and reaches an equilibrium described by its acid dissociation constant K_a . A larger K_a means a stronger weak acid, and a smaller pK_a ($= -\log_{10}K_a$) means the same thing.



Particle view of the same number of acid molecules in water.

Type	Examples	How to get $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$
strong acids	HCl, HBr, HI, HNO_3 , HClO_4 , H_2SO_4 (first step)	equal to the acid concentration
strong bases	group 1 hydroxides, $\text{Ca}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$, $\text{Ba}(\text{OH})_2$	$[\text{OH}^-] = \text{concentration} \times \text{number of OH}^-$
weak acids	CH_3COOH , HF, HCN, HCO_3^-	solve with K_a and an ICE table
weak bases	NH_3 , CH_3NH_2 , CH_3COO^-	solve with K_b and an ICE table

Worked example 1

Find the pH of $0.0250 \text{ mol/dm}^3 \text{ HNO}_3$ and of $0.0040 \text{ mol/dm}^3 \text{ Ba}(\text{OH})_2$. HNO_3 is strong, so $[\text{H}_3\text{O}^+] = 0.0250$ and $\text{pH} = -\log_{10}(0.0250) = \mathbf{1.60}$. $\text{Ba}(\text{OH})_2$ releases two OH^- , so $[\text{OH}^-] = 2 \times 0.0040 = 0.0080 \text{ mol/dm}^3$, $\text{pOH} = 2.10$ and $\text{pH} = 14.00 - 2.10 = \mathbf{11.90}$.

How do you find the pH of a weak acid?

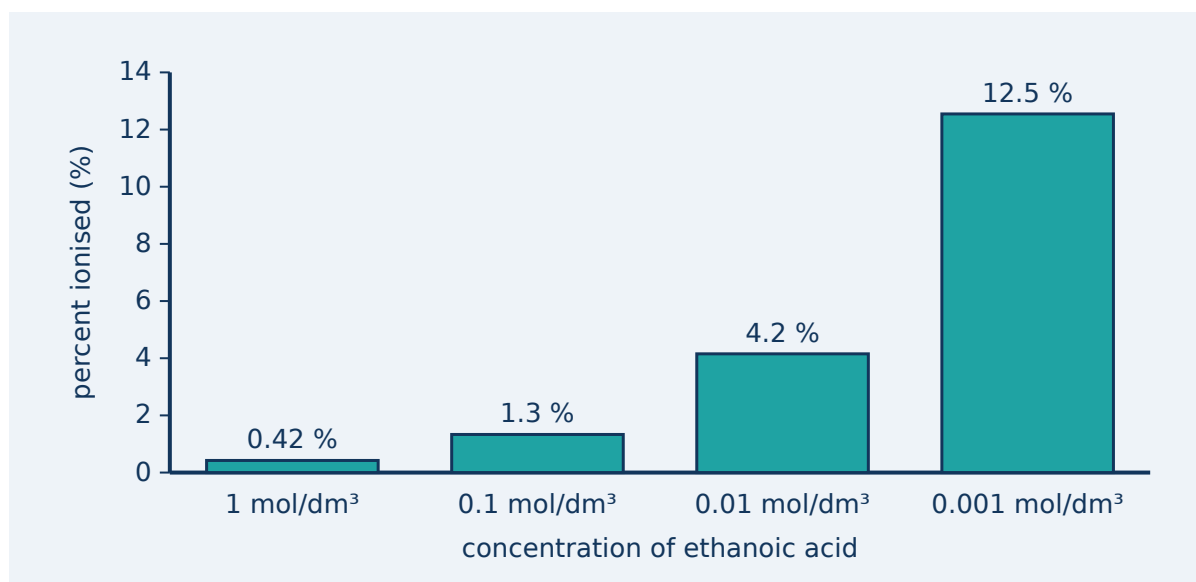
Write the K_a expression, set up an ICE table (initial, change, equilibrium) and let x be the amount of HA that ionises. For $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$, $K_a = x^2 / (C - x)$. When K_a is small compared with the concentration C , x is tiny and $C - x \approx C$, so $x = \sqrt{K_a \times C}$.

Key rule

The approximation is acceptable when x is less than 5 % of C . Always check it after calculating x . If x/C exceeds 5 %, solve the quadratic $x^2 + K_a \cdot x - K_a \cdot C = 0$ instead.

Worked example 2

Find the pH of 0.150 mol/dm^3 ethanoic acid, $K_a = 1.8 \times 10^{-5}$. $K_a = x^2/0.150$, so $x^2 = 2.7 \times 10^{-6}$ and $x = 1.64 \times 10^{-3} \text{ mol/dm}^3 = [\text{H}_3\text{O}^+]$. The check: 1.1 % of 0.150, below 5 %, so the approximation holds. $\text{pH} = -\log_{10}(1.643 \times 10^{-3}) = \mathbf{2.78}$.



More dilute weak acid, larger percentage ionised ($K_a = 1.8 \times 10^{-5}$).

How do weak bases and conjugate pairs work?

A weak base such as ammonia reacts with water: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$, with $K_b = [\text{NH}_4^+][\text{OH}^-]/[\text{NH}_3]$. The method matches the acid case, except that x is $[\text{OH}^-]$ and you convert through pOH . A conjugate acid-base pair obeys one neat link.

Formula

$K_a \times K_b = K_w$, so $\text{p}K_a + \text{p}K_b = 14.00$ for a conjugate pair. K_a is the constant of the acid, K_b that of its conjugate base. Both are dimensionless in AP calculations.

Worked example 3

Find the pH of 0.20 mol/dm³ NH_3 , $K_b = 1.8 \times 10^{-5}$. $x = [\text{OH}^-] = \sqrt{(1.8 \times 10^{-5} \times 0.20)} = 1.90 \times 10^{-3}$ mol/dm³. $\text{pOH} = 2.72$, so $\text{pH} = 14.00 - 2.72 = \mathbf{11.28}$.

Common mistake

Students write $\text{pH} = -\log_{10}(0.020)$ for 0.020 mol/dm³ of a weak acid. A weak acid never gives $[\text{H}_3\text{O}^+]$ equal to its concentration. Equally, never forget that $\text{Ba}(\text{OH})_2$ gives two OH^- per formula unit.

Lina's tip

Before you start, ask one question: strong or weak? Strong means no ICE table. Weak means write K_a , call the unknown x , and check the 5 % rule at the end.

Key points to remember

- $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$; $\text{pH} + \text{pOH} = 14.00$ at 25 °C ($K_w = 1.0 \times 10^{-14}$).
- Strong acid or base: ionises fully, so $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$ follows from the concentration.
- Weak acid: $K_a = x^2/(C - x)$, and $x \approx \sqrt{(K_a \times C)}$ if x under 5 % of C .

- Conjugate pair: $K_a \times K_b = K_w$, so $pK_a + pK_b = 14.00$.
- Weak acids ionise to a larger percentage as they are diluted.

Strong and weak acids, bases and pH: frequently asked questions

What is the difference between pH and pOH?

pH measures hydronium concentration and pOH measures hydroxide concentration. At 25 °C they always add to 14.00, so a solution with pOH 3.20 has pH 10.80.

When can you use the square-root shortcut for a weak acid?

Use $x = \sqrt{K_a \times C}$ when the answer x is less than 5 % of the starting concentration C . If it is bigger, solve the full quadratic instead.

Why is a 0.10 mol/dm³ weak acid less acidic than 0.10 mol/dm³ HCl?

The weak acid ionises only a small fraction of its molecules, so far fewer H_3O^+ ions are present. The two solutions still neutralise the same amount of base because the unionised acid keeps supplying protons.

How do you get K_b from K_a ?

For a conjugate pair, $K_b = K_w/K_a$. For example, F^- has $K_b = 1.0 \times 10^{-14} / 6.8 \times 10^{-4} = 1.5 \times 10^{-11}$.

Is pure water always at pH 7?

Only at 25 °C. K_w increases with temperature, so pure water at 50 °C has pH about 6.63 yet is still neutral because $[H_3O^+] = [OH^-]$.