

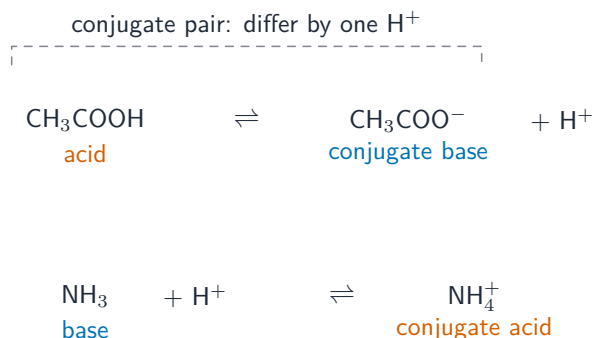
Equilibria

A-Level Chemistry

Conjugate acids and bases

When a Brønsted acid loses an H^+ , what is left is its **conjugate base**. When a base gains an H^+ , it becomes its **conjugate acid**. The two species that differ by just one H^+ form a **conjugate acid–base pair**.

For example, in $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$, the pair is CH_3COOH (acid) and CH_3COO^- (its conjugate base).



Conjugate pairs differ by one proton: an acid loses H^+ to give its conjugate base; a base gains H^+ to give its conjugate acid

pH and the equilibrium constants

- **pH** measures acidity: $\text{pH} = -\log[\text{H}^+]$, so $[\text{H}^+] = 10^{-\text{pH}}$.
- the **acid dissociation constant** of a weak acid HA is $K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$, and $\text{p}K_a = -\log K_a$. A larger K_a (smaller $\text{p}K_a$) means a stronger acid.
- the **ionic product of water** is $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ at 298 K.

0 7 14

acidic neutral alkaline

A hand wearing a white glove holds a box of Hanna P-4786 pH Test Strips. The box is white with a red label that includes the Hanna logo, the product name "P-4786 pH Test Strips 0-14.0", the lot number "Lot 046H0004", and the text "100 color-fixed indicator sticks". The box is open, showing several white test strips inside. To the right of the box is a small white plastic container with a red lid.

- **strong acid** : fully ionised, so $[\text{H}^+]$ equals the acid concentration; then take $-\log$.
- **strong alkali** : find $[\text{OH}^-]$ from the concentration, then use $[\text{H}^+] = K_w/[\text{OH}^-]$.
- **weak acid** : only partly ionised, so use $[\text{H}^+] = \sqrt{K_a \times [\text{HA}]}$.

It is fully ionised, so $[\text{H}^+] = 0.050 \text{ mol dm}^{-3}$, giving $\text{pH} = -\log(0.050) = 1.30$.

Worked example. Find the pH of 0.10 mol dm^{-3} sodium hydroxide (a strong base). ($K_w = 1.0 \times 10^{-14}$.)

$[\text{OH}^-] = 0.10$, so $[\text{H}^+] = K_w/[\text{OH}^-] = (1.0 \times 10^{-14})/0.10 = 1.0 \times 10^{-13}$, giving $\text{pH} = 13.0$.

Worked example. Find the pH of 0.10 mol dm^{-3} ethanoic acid (a weak acid). ($K_a = 1.8 \times 10^{-5}$.)

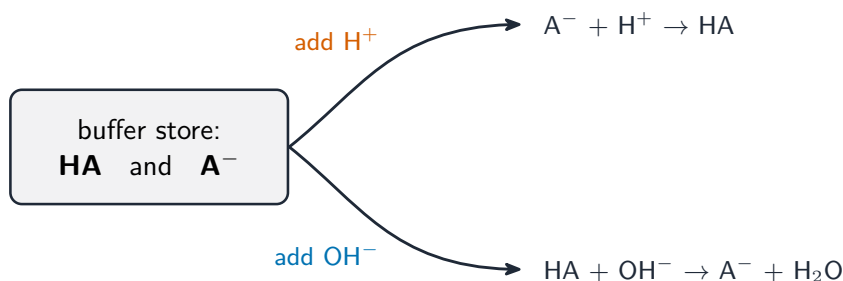
$$[\text{H}^+] = \sqrt{K_a \times [\text{HA}]} = \sqrt{(1.8 \times 10^{-5})(0.10)} = 1.3 \times 10^{-3}, \quad \text{pH} = -\log(1.3 \times 10^{-3}) = 2.87.$$

Buffer solutions

A **buffer solution** resists a change in pH when a small amount of acid or alkali is added. You make one from a weak acid and its conjugate base (for example ethanoic acid and sodium ethanoate).

It works because the mixture holds a store of both partners:

- added H^+ is removed by the conjugate base: $\text{CH}_3\text{COO}^- + \text{H}^+ \rightarrow \text{CH}_3\text{COOH}$.
- added OH^- is removed by the weak acid: $\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$.



either way, the pH barely changes

A buffer holds a store of a weak acid and its conjugate base: added H^+ is mopped up by A^- and added OH^- by HA , so the pH barely changes

To find the pH, put the concentrations of the acid and its salt into the K_a expression. Buffers are important in living things – for example, HCO_3^- keeps the pH of blood close to 7.4.

Solubility product

For a salt that barely dissolves, the **solubility product** (K_{sp}) is the product of the ion concentrations in a saturated solution, each raised to the power of its number in the formula:

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] \quad K_{\text{sp}} = [\text{Ca}^{2+}][\text{F}^-]^2$$

You can find K_{sp} from the solubility, or the solubility from K_{sp} .



Stalactites grow as dissolved calcium carbonate slowly comes back out of solution – a real solubility equilibrium

■ The common ion effect

The **common ion effect** is the way a salt becomes **less** soluble in a solution that already contains one of its ions. The extra ion pushes the dissolving equilibrium back (Le Chatelier), so less salt dissolves. You can calculate the new solubility using K_{sp} and the concentration of the common ion.

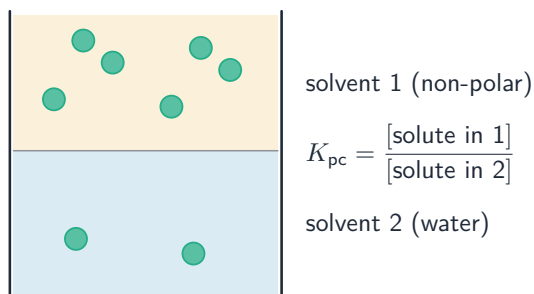
Partition coefficients

When a solute is shaken with two solvents that do not mix, it spreads between them. The **partition coefficient** (K_{pc}) is the ratio of its concentrations in the two layers (at constant temperature):

$$K_{\text{pc}} = \frac{[\text{solute in solvent 1}]}{[\text{solute in solvent 2}]}$$

This works when the **solute** is in the same physical state in both solvents. The value depends on the **polarity** of the solute and of each **solvent** : a non-

polar solute dissolves more in the non-polar solvent, while a polar solute prefers the polar solvent.



A solute shaken with two immiscible solvents spreads between them; the partition coefficient is the ratio of its concentrations in the two layers

Exam tips

- $\text{pH} = -\log[\text{H}^+]$; for a **strong acid** $[\text{H}^+] = \text{concentration}$, for a **weak acid** use $[\text{H}^+] = \sqrt{K_a[\text{HA}]}$.
- For a base, get $[\text{H}^+]$ from $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$.
- For a **buffer** use $[\text{H}^+] = K_a \times [\text{acid}]/[\text{salt}]$ and explain how it removes added H^+/OH^- .
- Write the K_{sp} expression with the correct powers; the number of decimal places in a pH equals the significant figures of $[\text{H}^+]$.