Reactivity 3.1 · Proton transfer reactions · supported focus

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

Supported HL/SL core focus. First assessment 2025; current brief acquired; full Chemistry guide not acquired. Remaining guide, assessment and practical requirements retain their recorded holds.

Prerequisites: chemical equations and reversible reactions from Structure and Reactivity 2, logarithms, and the mole. Each lesson develops its own method before independent transfer.

These are original or explicitly fictional teaching examples, not actual measurements or completed assessed learner investigations.

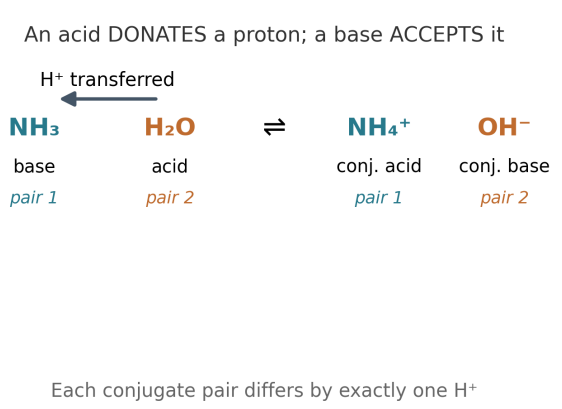
Proton transfer reactions

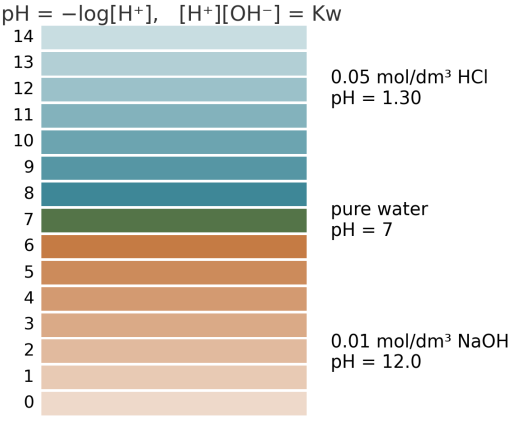
What would explain this observation?

- Two bottles both labelled "0.1 mol/dm^3 acid" give very different pH readings: 1.0 and 2.9. Same concentration, different acidity – how?
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

Build the model

- In the Brønsted–Lowry model, an acid is a **proton donor** and a base is a proton acceptor. Every such reaction links a **conjugate acid–base pair** differing by exactly one H^+: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ pairs NH_3 with NH_4^+ and H_2O with OH^-.
- A **strong acid** dissociates almost completely in water; a **weak acid** only partially, so the same concentration gives a different $[H^+]$. Water itself ionises slightly: $Kw = [H^+][OH^-] = 1.0 \times 10^{-14}$ mol^2/dm^6 at 298 K, and $pH = -\log[H^+]$.
- proton donor: A species that gives an H^+ to another species; conjugate acid–base pair: Two species differing by exactly one proton; strong acid: An acid that dissociates almost completely in water; weak acid: An acid that only partially dissociates in water.





Choose evidence that can test it

- Compare equal-concentration solutions: a strong acid conducts better and shows a lower pH than a weak acid of the same concentration, because more of its molecules have dissociated. Measure pH with a calibrated meter, not by estimating from the label.
- Adding the conjugate base of a weak acid shifts its equilibrium toward the undissociated form – evidence that the dissociation is a genuine equilibrium, not a one-way reaction.

Work from known quantities

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- Known: HCl is a strong acid at 0.05 mol/dm^3, so $[H^+] = 0.05$ mol/dm^3. $pH = -\log[H^+]$. Substitution: $pH = -\log(0.05) = 1.30$. For 0.01 mol/dm^3 NaOH, a strong base, $[OH^-] = 0.01$; using Kw, $[H^+] = Kw/[OH^-] = 1.0 \times 10^{-14} / 0.01 = 1.0 \times 10^{-12}$, so pH = 12.0.

Example:

A strong-acid solution has $[H^+] = 2.5 \times 10^{-4}$ mol/dm^3. Find its pH using the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: $pH = -\log(2.5 \times 10^{-4}) = 3.60$. pH is dimensionless; state the input concentration with its unit.)

Check the conclusion and its limits

- $pH = -\log[H^+]$ in this form assumes dilute solutions where concentration approximates activity; at high concentration the approximation weakens. $Kw = 1.0 \times 10^{-14}$ holds at 298 K and changes with temperature.

- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

A strong acid is the same thing as a concentrated acid. This claim is false: strength is the degree of dissociation and concentration is the amount per volume; a concentrated weak acid and a dilute strong acid both exist.

Key:

Proton transfer reactions: acids donate protons and bases accept them; every transfer links a conjugate pair differing by one H^+. Strength (degree of dissociation) is not concentration, and $pH = -\log[H^+]$ with $Kw = [H^+][OH^-]$.