

Reactivity 3.1 · Proton transfer reactions

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Reactivity 3.1, Proton transfer reactions, for the course first assessed in 2025.

Prerequisites: amount of substance and concentration from Structure 1.4, equilibrium and K from Reactivity 2.3, and logarithms to base 10. You will use K_w and K_a or pK_a values from the data booklet.

Brønsted–Lowry acids and bases

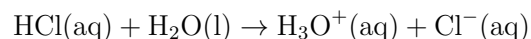
■ What do vinegar, stomach acid and cleaning ammonia have in common?

Vinegar and stomach acid both turn blue litmus red, and ammonia solution turns red litmus blue. One idea explains all three: in every acid–base reaction a **proton**, an H^+ ion, moves from one particle to another.

■ Acids donate, bases accept

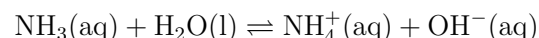
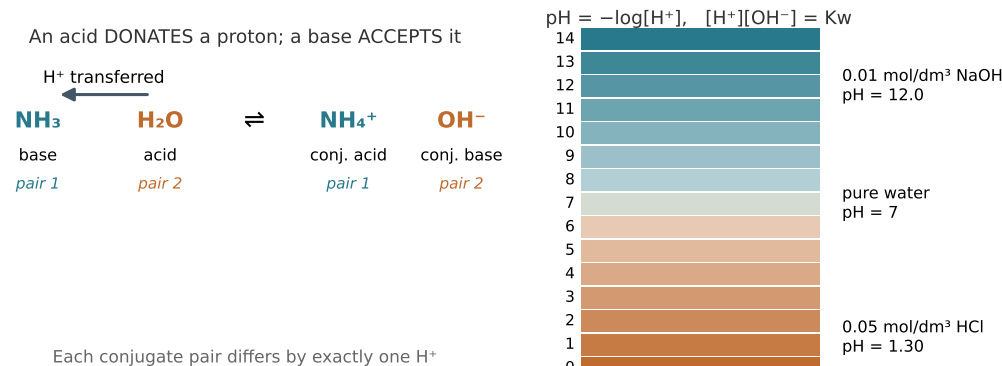
- A **Brønsted–Lowry acid** is a proton donor.
- A **Brønsted–Lowry base** is a proton acceptor. A base must have a lone pair to accept the proton.

In water a proton never exists on its own: it bonds to a water molecule to form the oxonium ion, H_3O^+ . Writing $H^+(aq)$ is a short form for $H_3O^+(aq)$.



■ Conjugate acid–base pairs

When an acid donates a proton, what is left can accept a proton back: it is a base. A **conjugate acid–base pair** is two species that differ by one proton.



A Brønsted–Lowry acid donates H^+ and a base accepts it; $pH = -\log[H^+]$

Here NH_3 (base) and NH_4^+ (its conjugate acid) are one pair; H_2O (acid) and OH^- (its conjugate base) are the other.

Water acts as a base with HCl and as an acid with NH_3 . A species that can both donate and accept a proton is **amphiprotic**. Other examples are HCO_3^- , HSO_4^- and $H_2PO_4^-$.

pH and the ionic product of water

■ The pH scale

The concentration of H^+ in solutions ranges from about 1 to 10^{-14} mol dm⁻³. The **pH** scale turns this into simple numbers:

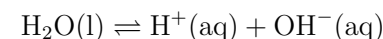
$$pH = -\log_{10}[H^+] \quad [H^+] = 10^{-pH}$$

- A decrease of one pH unit means $[H^+]$ is **ten times** larger.
- pH has no units. Give it to the same number of decimal places as the significant figures in $[H^+]$.

Worked example: $[H^+] = 0.050$ mol dm⁻³ gives $pH = -\log(0.050) = 1.30$. A solution of pH 3.40 has $[H^+] = 10^{-3.40} = 4.0 \times 10^{-4}$ mol dm⁻³.

■ Water ionizes a little

Pure water contains a tiny amount of ions:



The **ionic product of water**, K_w , is

$$K_w = [H^+][OH^-] = 1.00 \times 10^{-14} \text{ at } 298 \text{ K}$$

In pure water $[H^+] = [OH^-] = 1.00 \times 10^{-7}$, so the pH is 7.00. A solution is **neutral** when $[H^+] = [OH^-]$, acidic when $[H^+] > [OH^-]$ and alkaline when $[H^+] < [OH^-]$.

The ionization of water is endothermic, so K_w increases with temperature. At 50 °C pure water has a pH of about 6.6, but it is still neutral, because $[H^+]$ and $[OH^-]$ are still equal.

■ **pH of strong acids and bases**

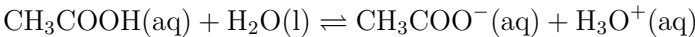
A strong acid or strong alkali is fully ionized, so the concentration of ions comes straight from the concentration of the solution.

- 0.020 mol dm⁻³ HCl: $[H^+] = 0.020$, pH = 1.70.
- 0.010 mol dm⁻³ NaOH: $[OH^-] = 0.010$, so $[H^+] = K_w/0.010 = 1.0 \times 10^{-12}$, pH = 12.00.

Strong and weak acids and bases

■ **Full and partial ionization**

- A **strong acid** ionizes completely in water: HCl, HNO₃, H₂SO₄ (first proton). A **strong base** ionizes or dissociates completely: group 1 hydroxides and Ba(OH)₂.
- A **weak acid** ionizes only partly; the equilibrium lies to the left: ethanoic acid and other carboxylic acids, carbonic acid. A **weak base** accepts protons only partly: ammonia and amines.



In 0.10 mol dm⁻³ ethanoic acid only about 1 molecule in 80 is ionized at any moment. Do not confuse **strong** with **concentrated**. Strength is about how far the acid ionizes; concentration is how much acid is dissolved in each dm³. A dilute solution of a strong acid can have a higher pH than a concentrated solution of a weak acid.

The stronger an acid, the weaker its conjugate base: Cl⁻ has almost no tendency to accept a proton, while CH₃COO⁻ accepts one easily.

■ **Telling them apart**

Compare solutions of the **same concentration** at the same temperature:

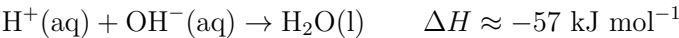
Test	0.10 mol dm ⁻³ HCl (strong)	0.10 mol dm ⁻³ CH ₃ COOH (weak)
pH	1.0	about 2.9
electrical conductivity	high (many ions)	low (few ions)
rate with magnesium or CaCO ₃	fast bubbling	slow bubbling

The total volume of gas, and the amount of alkali needed to neutralize each solution, are the **same**: as the weak acid reacts, its equilibrium shifts and more molecules ionize.

Neutralization and pH curves

■ **Neutralization**

An acid and a base react to form a salt and water. For any strong acid and strong base the reaction is the same:

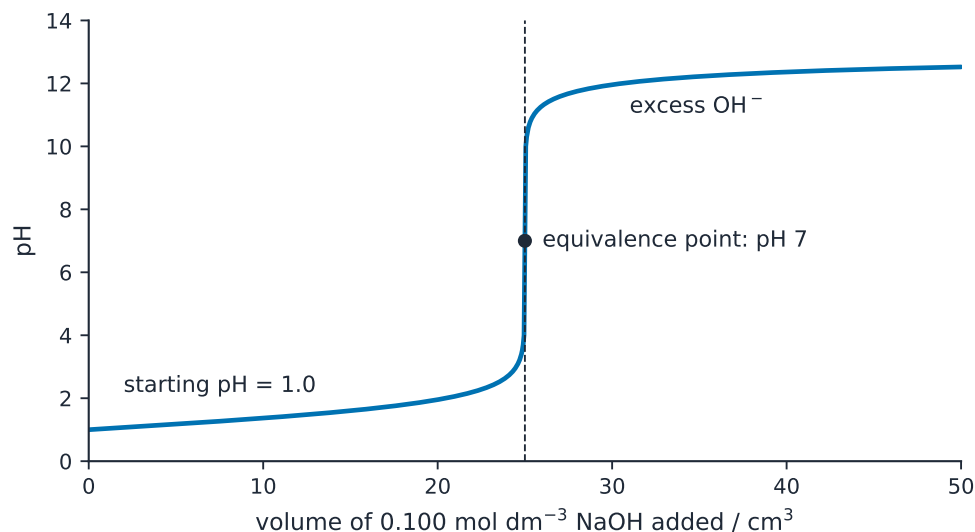


Neutralization is exothermic. Other reactions of acids:

- with a metal carbonate or hydrogencarbonate: salt, water and carbon dioxide;
- with a reactive metal: salt and hydrogen;
- with a metal oxide or hydroxide: salt and water.

■ **The pH curve of a strong acid and a strong base**

A **pH curve** shows how the pH changes as one solution is added to another in a titration.



The pH changes very little until almost all the acid is used up, then jumps by several units

- The **equivalence point** is where the amounts of acid and base match the equation. For a strong acid and strong base it is at pH 7.
- Near the equivalence point one drop of alkali changes the pH by several units. This steep section is why an indicator can show the end point sharply.
- The starting pH comes from the acid concentration, and the final pH approaches the pH of the alkali.

Practice

1. Identify the two conjugate acid–base pairs in $\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{NO}_3^- + \text{H}_3\text{O}^+$.
2. Write equations to show that HCO_3^- is amphoteric.
3. Calculate the pH of: (a) $0.0025 \text{ mol dm}^{-3} \text{ HCl}$, (b) $0.050 \text{ mol dm}^{-3} \text{ KOH}$.
4. A solution has pH 4.70. Calculate $[\text{H}^+]$ and $[\text{OH}^-]$.
5. Describe two experiments that distinguish 0.10 mol dm^{-3} solutions of a strong and a weak acid, with the expected results.
6. Explain why pure water at 60°C has a pH below 7 but is still neutral.
7. Sketch the pH curve when $0.100 \text{ mol dm}^{-3} \text{ NaOH}$ is added to 20.0 cm^3 of $0.100 \text{ mol dm}^{-3} \text{ HCl}$, marking the start pH and the equivalence point.

Answers.

1. HNO_3 (acid) and NO_3^- (conjugate base); H_3O^+ (acid) and H_2O (conjugate base).
2. As an acid: $\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{CO}_3^{2-} + \text{H}_3\text{O}^+$. As a base: $\text{HCO}_3^- + \text{H}_3\text{O}^+ \rightleftharpoons \text{H}_2\text{CO}_3 + \text{H}_2\text{O}$.

3. (a) $\text{pH} = -\log(0.0025) = 2.60$. (b) $[\text{H}^+] = 1.00 \times 10^{-14}/0.050 = 2.0 \times 10^{-13}$, $\text{pH} = 12.70$.
4. $[\text{H}^+] = 10^{-4.70} = 2.0 \times 10^{-5} \text{ mol dm}^{-3}$; $[\text{OH}^-] = 1.00 \times 10^{-14}/2.0 \times 10^{-5} = 5.0 \times 10^{-10} \text{ mol dm}^{-3}$.
5. Any two: measure the pH (strong about 1, weak about 3); measure the conductivity (strong much higher); add magnesium ribbon or calcium carbonate (strong bubbles faster). Use the same concentration, volume and temperature.
6. Ionization of water is endothermic, so at higher temperature the equilibrium shifts right and $[\text{H}^+]$ rises above 10^{-7} , giving pH below 7. But $[\text{H}^+] = [\text{OH}^-]$ still, so the water is neutral.
7. Start at pH 1.0; slow rise; vertical jump from about pH 3 to 11 at 20.0 cm^3 with the equivalence point at pH 7; then levelling off near pH 12–13.

Exam tips

- Define acids and bases as proton **donors** and **acceptors**; a base needs a **lone pair**.
- Conjugate pairs differ by exactly **one** H^+ ; always pair each acid with its own conjugate base.
- Give pH to as many **decimal places** as $[\text{H}^+]$ has significant figures.
- Compare strong and weak acids at the **same concentration**; never use “strong” to mean “concentrated”.
- Neutral means $[\text{H}^+] = [\text{OH}^-]$, which is pH 7 only at 298 K.