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Topic 8.1: Introduction to Acids and Bases

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UNIT 8 · ACIDS AND BASES

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The Brønsted-Lowry Definitions

A Brønsted-Lowry acid is a species that donates a proton, H^+, to another species. A Brønsted-Lowry base is a species that accepts a proton from another species.

The labels describe what a species does in the reaction in front of you, not what it is in isolation. NH_3 contains no OH^- and releases none, yet it is a base because it accepts a proton from water:

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Every Brønsted-Lowry reaction therefore has two proton transfers running at once: one species hands a proton over, and another takes it.

A Lewis acid accepts an electron pair and a Lewis base donates one, so BF_3, whose boron atom has an incomplete octet, is a Lewis acid even though it holds no proton to give. The questions in this topic use the proton definitions; Lewis labels appear only at recognition level.

Conjugate Pairs Differ by Exactly One Proton

When an acid donates its proton, what remains is its conjugate base. HCl donates H^+ and leaves Cl^-, so HCl and Cl^- are a conjugate acid-base pair.

When a base accepts a proton, what forms is its conjugate acid. NH_3 accepts H^+ and becomes NH_4^+, so NH_3 and NH_4^+ are a conjugate pair.

The two members of a pair differ by exactly one proton, and by exactly one unit of charge, because the proton carries charge +1. Removing H^+ from $H_2PO_4^-$ gives HPO_4^{2-}: one fewer hydrogen, one more unit of negative charge.

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That one-proton test is also how false pairs are caught. H_2SO_4 and SO_4^{2-} differ by two protons, so they are not a pair. The conjugate base of H_2SO_4 is HSO_4^-, and SO_4^{2-} is the conjugate base of HSO_4^-.

Two Pairs in Every Reaction

A complete proton-transfer equation contains two conjugate pairs, one for each side's transfer. In the reaction above, the pairs are NH_4^+/NH_3 and H_2O/OH^-.

The acid of the first pair, NH_4^+, is the protonated form of the base NH_3. The acid of the second pair, H_2O, is the protonated form of the base OH^-.

Labelling a reaction means naming four roles: the acid and its conjugate base on the reactant side's transfer, and the base and its conjugate acid. Each reactant belongs to one pair, and each product is the other member of one of those pairs.

Species That Play Both Roles

A species that carries a proton it can give and a lone pair that can take one is amphiprotic: it can act as an acid or as a base. Water is the standard example. It donates a proton to become OH^-, and accepts one to become H_3O^+.

HCO_3^- is also amphiprotic. Given a proton it becomes H_2CO_3; stripped of one it becomes CO_3^{2-}.

The role is decided by the partner. Against a stronger acid, an amphiprotic species accepts the proton and plays the base. Against a stronger base, it donates and plays the acid.

$HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$

Here HCO_3^- is the acid and water is the base, the reverse of water's role with NH_3.

Water Ionizes Itself

In pure water, one water molecule occasionally donates a proton to another:

$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

This is the autoionization of water, and it is an equilibrium that lies far to the left.

Its equilibrium constant is Kw, the ion-product constant:

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C

Because the product is fixed at a given temperature, the two concentrations are locked together: raise one and the other must fall.

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In pure water the two ions form in equal amounts, so $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol L^{-1} at 25 °C. That equality is what neutral means. Acidity is then read on the pH scale:

$pH = -\log[H_3O^+]$

That gives pH 7 for the neutral point at 25 °C, with one pH unit per factor of ten in $[H_3O^+]$.

Strong, Weak, and the Conjugate Seesaw

A strong acid ionizes completely in water: essentially every HCl molecule transfers its proton, and no HCl is left:

$HCl + H_2O \rightarrow Cl^- + H_3O^+$

The common strong acids are HCl, HBr, HI, HNO_3, $HClO_4$ and H_2SO_4 (first proton only).

A weak acid ionizes only partially, so most of it stays as molecules:

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

That reaction lies well to the left. HF and H_2CO_3 are weak in the same way.

Conjugate strengths sit on a seesaw. The stronger the acid, the weaker its conjugate base. Cl^-, left behind by the strong acid HCl, has almost no pull on a proton and is effectively not a base in water at all. The conjugate base of a weak acid holds its proton loosely but still measurably, so F^- and CH_3COO^- do accept protons back.

Worked Examples

Example 1: Labelling a reaction.

$HF + NH_3 \rightleftharpoons F^- + NH_4^+$

HF loses H^+, so HF is the acid and F^- is its conjugate base. NH_3 gains H^+, so NH_3 is the base and NH_4^+ is its conjugate acid. Both pairs check out: HF/F^- and NH_4^+/NH_3 each differ by exactly one proton.

Example 2: Using Kw. A solution at 25 °C has $[H_3O^+] = 2.5 \times 10^{-4}$ mol L^{-1}. Then $[OH^-] = Kw \div [H_3O^+] = (1.0 \times 10^{-14}) \div (2.5 \times 10^{-4}) = 4.0 \times 10^{-11}$ mol L^{-1}. The product of the two answers returns 1.0×10^{-14}, and $[H_3O^+] > [OH^-]$, so the solution is acidic.

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Example 3: An equilibrium constant for a weak acid. A 0.20 mol L^{-1} solution of a weak acid HA reaches equilibrium with $[H_3O^+] = 2.0 \times 10^{-3}$ mol L^{-1}. The ionization produces equal $[H_3O^+]$ and $[A^-]$, so $Ka \approx [H_3O^+][A^-] \div [HA] = (2.0 \times 10^{-3})^2 \div 0.20 = 2.0 \times 10^{-5}$. Only 1.0 percent of the acid has ionized, which is what weak means in numbers.

Example 4: Writing conjugates down a ladder. For the sulfuric acid family: the conjugate base of H_2SO_4 is HSO_4^-, and the conjugate base of HSO_4^- is SO_4^{2-}. Each step removes one proton and adds one unit of negative charge. The conjugate acid of SO_4^{2-}, running the ladder upward, is HSO_4^-, not H_2SO_4.

Traps That Cost Marks in Topic 8.1

"H_2SO_4 and SO_4^{2-} are a conjugate pair." They differ by two protons. A pair differs by exactly one; H_2SO_4 pairs with HSO_4^-.

"NH_3 is an acid because its formula contains hydrogen." Formulas do not assign roles, behavior does. In water NH_3 accepts a proton, which makes it the base, and its conjugate acid is NH_4^+.

"Water is the acid in every reaction it joins." Water is amphiprotic. With NH_3 it donates and is the acid; with HCO_3^- it accepts and is the base.

"The conjugate base of a strong acid is a strong base." Strengths run in opposite directions across a pair. Cl^- is such a weak base that it is a spectator in water.

"Strong means concentrated." Strength is the fraction ionized; concentration is how much solute is present. A very dilute HCl solution is still a strong acid, fully ionized, and concentrated CH_3COOH is still weak.

"Neutral always means pH 7." Neutral means $[H_3O^+] = [OH^-]$. That gives pH 7 only at 25 °C, where $Kw = 1.0 \times 10^{-14}$; Kw grows with temperature, so the neutral pH falls below 7 in hot water while the solution stays neutral.

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Key Relationships to Remember

- A Brønsted-Lowry acid donates a proton; a Brønsted-Lowry base accepts one
- Roles are assigned from the reaction shown: the species that loses H^+ is the acid, the one that gains it is the base
- An acid and its conjugate base differ by exactly one proton and one unit of charge; so do a base and its conjugate acid
- Every proton-transfer equation contains two conjugate pairs
- Amphiprotic species such as H_2O, HCO_3^- and $H_2PO_4^-$ can donate or accept a proton, and the partner decides the role
- Autoionization: $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
- $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C
- Neutral means $[H_3O^+] = [OH^-]$; at 25 °C each is 1.0×10^{-7} mol L^{-1} and the pH is 7
- The stronger the acid, the weaker its conjugate base; a strong acid ionizes completely and a weak acid only partially

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