

how far an electrolyte splits, and what that split does to pH



Strong vs Weak Electrolytes

Strong electrolyte

- ionises completely, or almost completely
- strong acids, strong bases and salts
- e.g. HCl, HNO₃, H₂SO₄, NaOH, KOH

Weak electrolyte

- ionises only partially, an equilibrium survives between ions and unionised molecules
- weak acids and weak bases
- e.g. CH₃COOH, HCOOH, HF, H₂S, NH₄OH

Degree of Dissociation *

fraction of total moles that dissociates at equilibrium $\alpha = \frac{\text{moles dissociated}}{\text{total moles}}$

percent dissociation $\% \text{ dissociation} = \alpha \times 100$

equilibrium concentration of either ion for a 1:1 electrolyte, initial concentration c $[\text{ion}] = \alpha c$

Acids, Bases and pH *

Arrhenius

acid: gives H⁺ in water;
base: gives OH⁻. Limited to water, and fails for NH₃, Na₂CO₃ (no OH group)

Bronsted-Lowry

acid donates a proton, base accepts one; conjugate pairs differ by one proton (HCl/Cl⁻, H₂O/H₃O⁺)

Lewis

acid accepts an electron pair, base donates one; broadest, covers species with no proton

every Bronsted base is also a Lewis base, since accepting a proton means donating an electron pair to it. Not every Bronsted acid is a Lewis acid though, since a Bronsted acid like HCl does not itself accept an electron pair, only the H⁺ it releases does

Amphoteric Nature of Water

- water acts as a base toward HCl: it accepts a proton, forming H₃O⁺ and Cl⁻
- water acts as an acid toward NH₃: it donates a proton, forming OH⁻ and NH₄⁺
- this dual role, acid toward one substance and base toward another, is called amphoteric behaviour

Ionisation Constants

acid dissociation constant, HA to H⁺ and A⁻

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

base dissociation constant, BOH to B⁺ and OH⁻

$$K_b = \frac{[\text{B}^+][\text{OH}^-]}{[\text{BOH}]}$$

a bigger K_a or K_b means a stronger weak acid or base, more of it sits as ions once equilibrium is reached

