

Chapter 6 • Ionic Equilibrium

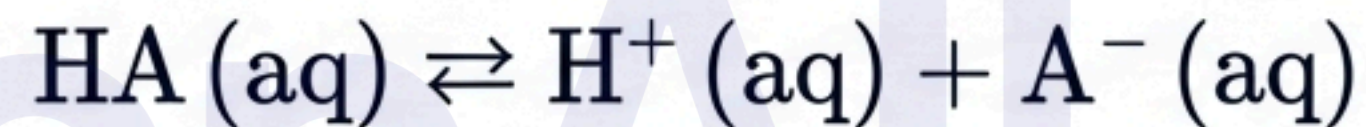
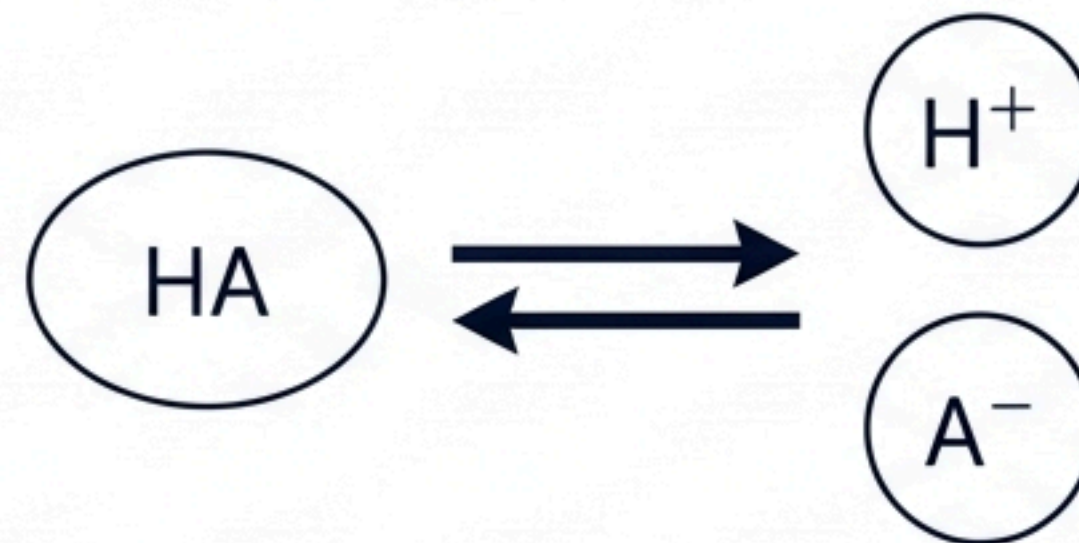
Measuring Acid and Base Strength

Mastering K_a , K_b , and the Mathematics of Equilibrium

Concept Explainer

The Reality of Weak Acids

Unlike strong acids that dissociate completely, **weak acids only partially ionize** in water. They reach a state of dynamic equilibrium where both the forward and reverse reactions happen at the exact same rate.



The Key Players:

- HA: Undissociated acid molecules (The majority)
- H⁺: Hydrogen ions (The source of acidity)
- A⁻: Conjugate base ions

Deriving the Acid Ionisation Constant (Ka)

By applying the Law of Chemical Equilibrium to the dissociation of a weak acid, we derive Ka.


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

Concentration of the products (ions).

Concentration of the un-ionized reactant at equilibrium.

**Note: Water (H_2O) is a pure liquid, so its concentration is constant and incorporated directly into Ka.*



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$$E=mc^2$$



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Decoding the Ka Value

At a given temperature, K_a is the ultimate measure of an acid's inherent strength.

Larger K_a Value

Numerator $[H^+][A^-]$ is large. The equilibrium is shifted to the right.

The Result:

→ **Stronger Acid**

Smaller K_a Value

Denominator $[HA]$ is large. The acid prefers to stay undissociated.

The Result:

→ **Weaker Acid**

The pKa Flip

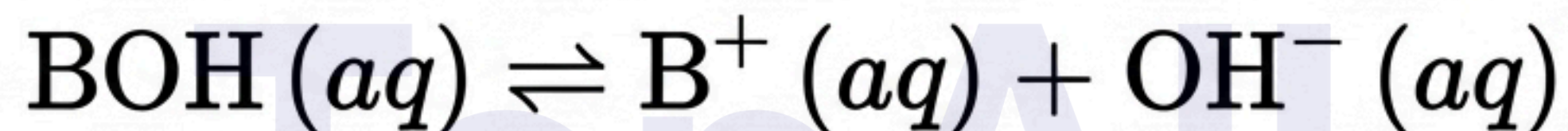
Because K_a values are tiny (e.g., 1.8×10^{-5}), we use a logarithmic scale:

$$pK_a = -\log(K_a)$$

**Crucial Rule: Lower
pKa = Stronger Acid.**

The Reality of Weak Bases

Just like weak acids, weak bases (like ammonia or organic amines) do not fully dissociate. They reach a dynamic equilibrium in water, yielding a small amount of hydroxide ions.



(Alternatively: $\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^-$)

The Key Players:

- BOH: Undissociated base molecules
- B^+ : Conjugate acid ions
- OH^- : Hydroxyl ions (The source of basicity)

Defining the Base Ionisation Constant (K_b)

The equilibrium constant for base ionization mirrors the acid constant, focusing on hydroxide production.

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

Concentration of
ionized products.

Concentration of
un-ionized base.

The Rule of Strength:

- Higher K_b = More OH⁻ produced = Stronger Base
- Lower pK_b = Stronger Base (where pK_b = -log(K_b))

Do Not Confuse Concentration with Strength

Examiners frequently test if you know the difference between a dilute strong acid and a concentrated weak acid.

The Trap

A 0.0001 M solution of HCl is a weak acid
~~because the H^+ concentration is very low.~~



The Reality

Strength is determined only by K_a / K_b (inherent tendency to ionize). Concentration (c) is merely how much you put in the beaker.

- **HCl is ALWAYS a strong acid (High K_a),** even if highly diluted.
- **Acetic acid is ALWAYS a weak acid (Low K_a),** even if highly concentrated.

Ionic Equilibrium Snapshot

Acids

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

$$K_a = \frac{c\alpha^2}{1 - \alpha}$$

where α is degree of ionization

$$\text{p}K_a = -\log(K_a)$$

Bases

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

$$K_b = \frac{c\alpha^2}{1 - \alpha}$$

$$\text{p}K_b = -\log(K_b)$$

The Conjugate Relationship (At 298K):

$$K_a \times K_b = K_w = 1.0 \times 10^{-14}$$

$$\text{p}K_a + \text{p}K_b = \text{p}K_w = 14$$

✓ **Key Takeaway:** For a conjugate acid-base pair, a stronger acid yields a much weaker conjugate base, and vice versa!



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