

Acids, bases and pH

Bronsted-Lowry definitions

Acid: proton donor

Base: proton acceptor

monobasic $\rightarrow 1 \text{ H}^+$

dibasic $\rightarrow 2 \text{ H}^+$

tribasic $\rightarrow 3 \text{ H}^+$

Arrhenius model

Acids dissociate and release H^+ ions in aqueous solution

Alkalis dissociate and release OH^- ions in aqueous solution

H^+ ions are neutralised by OH^- ions to form water $[\text{H}^+ + \text{OH}^-]_{\text{aq}} \rightarrow \text{H}_2\text{O}$

An acid is a proton donor



• Stronger acid donates proton so other acid acts as a base

• Each side of arrow needs:

① An acid and a base

② Different numbers



$$\text{pH} = -\log [\text{H}^+]$$

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

pH of strong acids

K_a - acid dissociation constant

Higher $K_a \rightarrow$ Stronger acid

Lower $\text{p}K_a \rightarrow$ Stronger acid



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

Assume full dissociation so initial $[\text{HA}] = [\text{H}^+]$ when in solution and $[\text{H}^+] = [\text{A}^-]$ so $K_a = [\text{H}^+]^2$

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

$$K_a = 10^{-\text{p}K_a}$$

A Level Chemistry Revision Notes

C. Boes



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