

## 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  $\text{H}^+$  ions in aqueous solution

Alkalis dissociate and release  $\text{OH}^-$  ions in aqueous solution

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

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

**Mike Smith**



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