Important ACID-BASE Equations and when to use them.

Equation 1

$$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}\ \mathrm{mol^2\ dm^{-6}}$$

K_w is known as the **ionic product of water** and it has a value of $10^{-14}\ \mathrm{mol^2\ dm^{-6}}$ (ONLY at 25 °C!)

Equation 2a

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

Equation 2b

$$[H^+] = \mathrm{antilog}_{10}\ (-\ pH)$$

10^x button on most calculators!

Equation 3a

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

Making the assumption that $[H^+] = [A^-]$ when HA has been dissolved in water!

Equation 3b

⚠ Caution

$$K_a = \frac{[H^+]^2}{[HA]}$$

Making the assumption that $[HA]$ at equilibrium is approx. equal to $[HA]$ before any dissociation.

Equation 3c

⚠ Caution

$$[H^+] = \sqrt{K_a \times [HA]}$$

⚠ Caution These equations (**3b and 3c**) should only be used when dealing with a simple solution of a weak acid that has been dissolved in water, i.e. no neutralisation has taken place!

Equation 4

$$pK_a = -\ \log_{10} K_a$$

Values of *K*a vary so greatly that dissociation of molecules is often quoted on the p*K*a scale. The **more acidic** a molecule, the **greater the *Ka*** and the **smaller the p*Ka***

For situations when **some neutralisation has taken place**, we need to use the following equation.

Salt concentration, which is also the conjugate base.

Equation 5
'Henderson-Hasselbalch' eqn!

$$pH = pK_a + \log_{10}\left(\frac{[A^-]}{[HA]}\right) \quad \text{OR} \quad pH = pK_a + \log_{10}\left(\frac{n_{A^-}}{n_{HA}}\right)$$

ratio of concentrations

ratio of moles

Remember $[A^-]/[HA]$ has the **same value** as n_{A^-}/n_{HA}

This equation has its limitations. See earlier!