pH

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1 pH

pH refers to the negative log base of the proton concentration of a solution. pH and proton concentration are related by the following equations.

$$pH = -log_{10}([H^+])$$

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

2 Strong Acid

$$HA \rightarrow H^+ + A^-$$

In a strong acid there is full ionization which means that all of the left hand side goes to the right hand side. This means that the proton concentration is equal to the concentration of the acid.

$$[H^+] = [HA]$$

3 Weak Acid

In a weak acid there is partial deionization which means that a disassociation constant, pK_a is brought in. This relates to the proportion of the solution which has ionized.

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

$$pK_a = -log_{10}(Ka)$$

3.1 pH Calculations

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

$$[H^+] = [A^-]$$

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

For small $[H^+]$
so
And hence the proton concentration can be calculated from the weak acid concentration and the disassociation constant (which will depend on temperature).

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

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

4 Ionic Product of Water

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

$$K_w = [H_2O]K_c = [H^+][OH^-]$$

This reaction is endothermic, and therefore the hotter it is the more protons will be released. In neutral water[1],

$$[H^+] = [OH^-]$$

$$K_w = [H^+]^2 \text{ (so } mol^2 dm^{-6}\text{)}$$

5 pH of a Base

The pH of a base can be calculated using the ionic product of water. As in a base the A^- concentration, $[A^-]$ can be known from the molarity of the base, it can be used to calculate the proton concentration.

$$K_w = [H^+][OH^-]$$

$$[H^+] = \frac{K_w}{[OH^-]}$$

6 Buffer Solution Calculations

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

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

$$pH = pK_a + log \frac{[A^-]}{[HA]}$$

[1]At 25 C, $K_w = 1 \cdot 10^{-14} mol^2 dm^{-6}$