A2 Physical Chemistry

Weak Acids

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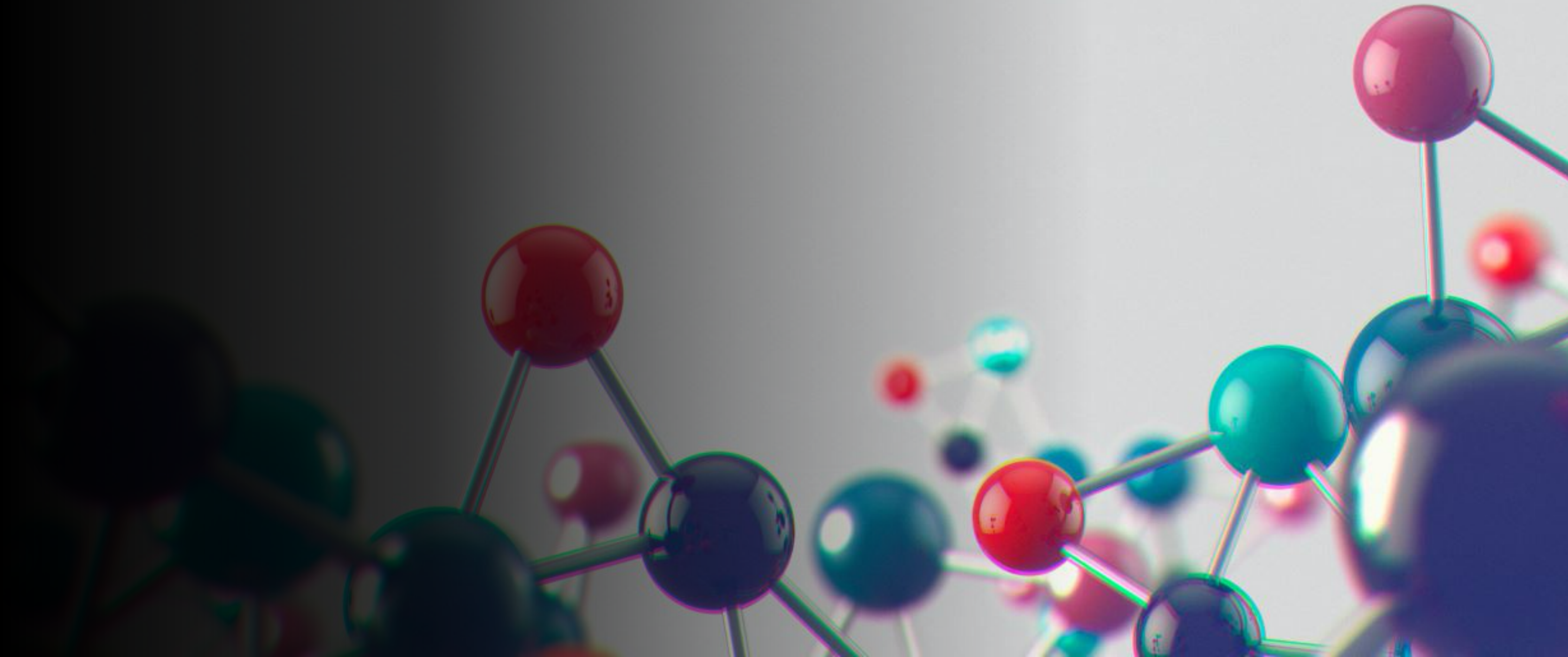


Weak Acids

Weak acids are only partially dissociated into ions when added to water.

All organic acids are weak acids.

$$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

Therefore there are many more H^+ ions in 1 dm^3 of 1 mol/dm^3 hydrochloric acid than there are in the same volume of 1 mol/dm^3 ethanoic acid.

K_a and pK_a

The acid dissociation constant (K_a) is the equilibrium constant for the ionisation of a weak acid.

$$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

$$K_a = \frac{[H^+_{(aq)}][CH_3COO^-_{(aq)}]}{[CH_3COOH_{(aq)}]}$$

Units: mol dm^{-3}

Ka – the bigger the value, the stronger the acid

$pK_a = -\log_{10}K_a$ $\qquad$ $K_a = 10^{-pKa}$

pKa – the smaller the value, the stronger the acid

Calculating pH of a weak acid

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

$$K_a = \frac{[H^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]}$$

$$K_a = \frac{[H^+_{(aq)}]^2}{[HA_{(aq)}]}$$

Approximations

$[H^+_{(aq)}] = [A^-_{(aq)}]$

$[HA_{(aq)}]$ = original concentration

$$[H^+_{(aq)}]^2 = K_a \times [HA_{(aq)}]$$

$$[H^+_{(aq)}] = \sqrt{K_a \times [HA_{(aq)}]}$$

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

Worked examples using K_a and pK_a

Example 1
Calculate the pH 0.250 mol dm^{-3} propanoic acid (pKa = 4.87).

$K_a = 10^{-pKa}$

$K_a = 10^{-4.87}$ $K_a = 1.35 \times 10^{-5}$

$$[H^+_{(aq)}] = \sqrt{K_a \times [HA_{(aq)}]}$$

$$[H^+_{(aq)}] = \sqrt{1.35 \times 10^{-5} \times 0.250}$$

$[H^+_{(aq)}] = 1.84 \times 10^{-3}$

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

pH = $-\log_{10}(1.84 \times 10^{-3})$

pH = 2.74

Example 2

Calculate the concentration of ethanoic acid with pH 4.97

(Ka = 1.74 x 10^{-5} mol dm^{-3})

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

$[H^+] = 10^{-4.97}$ $\quad$ $[H^+] = 1.07 \times 10^{-5}$

$[H^+_{(aq)}]^2 = K_a \times [HA_{(aq)}]$

$(1.07 \times 10^{-5})^2 = 1.74 \times 10^{-5} \times [HA_{(aq)}]$

$$[HA_{(aq)}] = \frac{(1.07 \times 10^{-5})^2}{1.74 \times 10^{-5}} = \frac{1.145 \times 10^{-10}}{1.74 \times 10^{-5}}$$

= 6.58 x 10^{-6} mol dm^{-3}

Example 3

Calculate the Ka value for phenyl ethanoic acid given that a 0.100 mol dm^{-3} solution has a pH of 2.66.

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

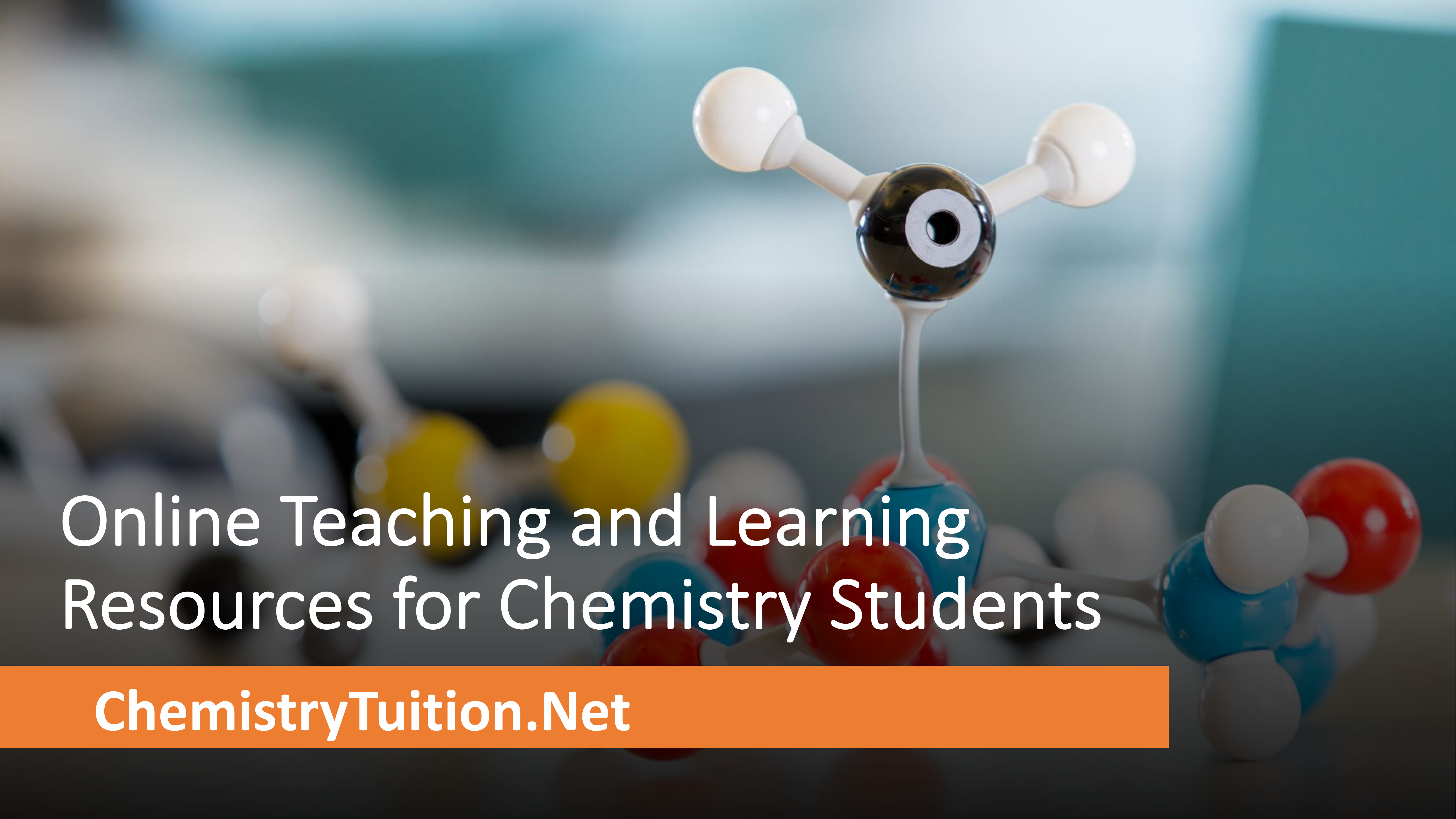
$[H^+] = 10^{-2.66}$ $\quad$ $[H^+] = 2.19 \times 10^{-3}$

$[H^+_{(aq)}]^2 = K_a \times [HA_{(aq)}]$

$(2.19 \times 10^{-3})^2 = K_a \times 0.100$

$$K_a = \frac{(2.19 \times 10^{-3})^2}{0.100} = \frac{4.7961 \times 10^{-6}}{0.100}$$

$= 4.79 \times 10^{-5}$ mol dm^{-3}



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