A2 Physical Chemistry

Buffer Calculations

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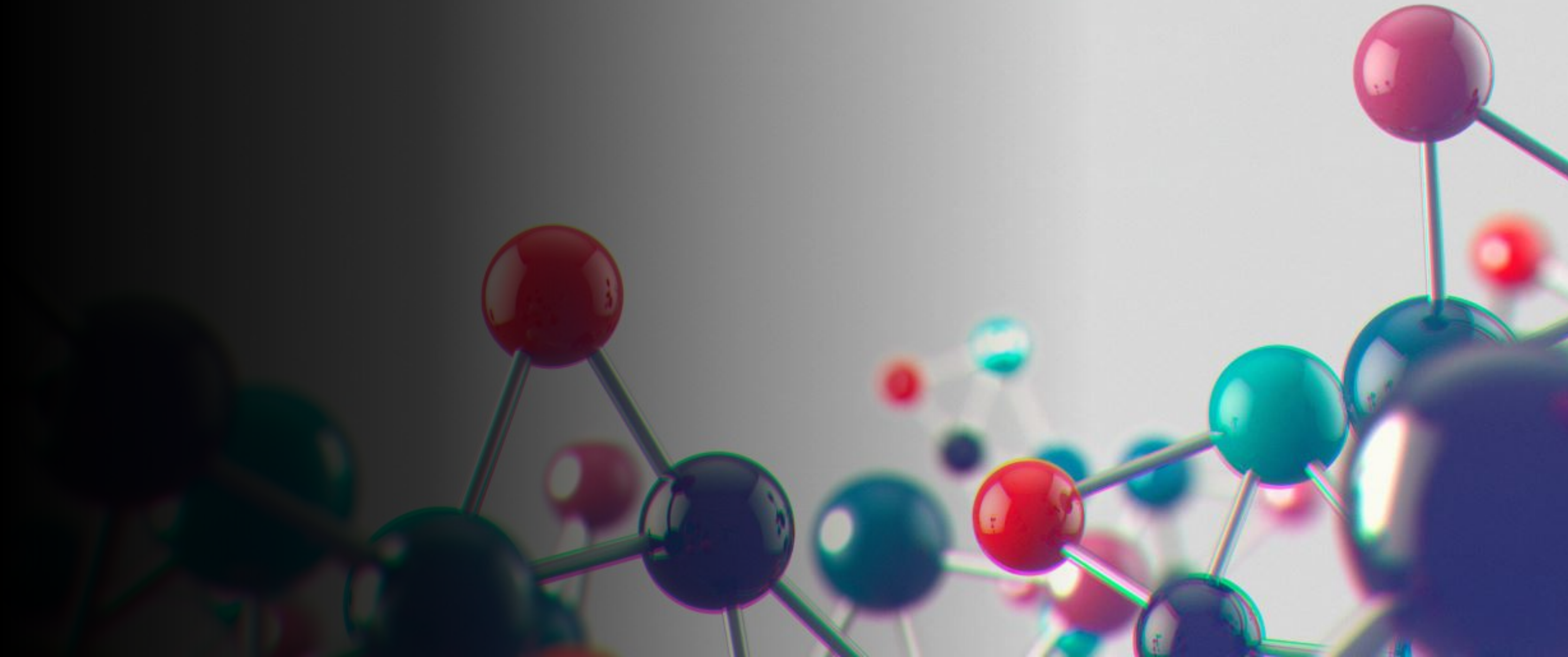

Acidic Buffers

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

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

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

Basic Buffers

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons HB^+_{(aq)} + OH^-_{(aq)}$$

$$K_b = \frac{[HB^+_{(aq)}][OH^-_{(aq)}]}{[B_{aq)}]}$$

$$\boxed{[OH^-_{(aq)}] = K_b \times \frac{[B^-_{(aq)}]}{[HB^+_{(aq)}]}}$$

A buffer solution was made by adding 2.05 g of sodium ethanoate to 0.500 dm^3 of 0.01 mol dm^{-3} ethanoic acid. Calculate the pH of this solution

(K_a for ethanoic acid = 1.74×10^{-5} mol dm^{-3}).

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

Moles of sodium ethanoate $= \frac{2.05}{82} =$ 0.025 moles

Conc of sodium ethanoate $= \frac{0.025}{0.500} =$ 0.050 mol dm^{-3}

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$[H^+_{(aq)}] = 3.48 \times 10^{-6}$ mol dm^{-3}

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

pH = $-\log(3.48 \times 10^{-6})$

pH = 5.46

A buffer solution was made by mixing 25.0 cm^3 of 1.00 mol dm^{-3} ethanoic acid with 25 cm^3 of 0.400 mol dm^{-3} sodium hydroxide.

(K_a for ethanoic acid = 1.74×10^{-5} mol dm^{-3}).

Find the pH of this buffer.

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Find the pH of this buffer.

Moles of Ethanoic acid = 0.025 x 1.00 = 0.025 moles

Moles of NaOH = 0.025 x 0.400 = 0.010 moles

$$CH_3COOH + NaOH \rightarrow CH_3COO^-Na^+ + H_2O$$

Moles of ethanoic acid remaining = 0.025 – 0.010 = 0.015 moles

Conc of ethanoic acid = $\frac{0.015}{0.050}$ = 0.300 mol dm^{-3}

Moles of sodium ethanoate formed = 0.010 moles

Conc of sodium ethanoate = $\frac{0.010}{0.050}$ = 0.200 mol dm^{-3}

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Conc of sodium ethanoate = 0.200 mol dm^{-3}

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

$$[H^+_{(aq)}] = 1.74\ 10^{-5} \times \frac{0.300}{0.200}$$

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

$[H^+_{(aq)}] = 2.61 \times 10^{-5}$ mol dm^{-3}

pH = -log[H^+]

pH = -log (2.61×10^{-5})

pH = 4.58

A buffer solution contains 0.20 mole of NH_3 and 0.60 mole NH_4Cl in 750 cm^3. Calculate the pH of this solution (K_b for ammonia = 1.8×10^{-5} mol dm^{-3}).

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$$[NH_{3(aq)}] = \frac{0.20}{0.750} = 0.267 \text{ mol dm}^{-3}$$

$$[NH_4Cl] = \frac{0.60}{0.750} = 0.800 \text{ mol dm}^{-3}$$

$$[OH^-_{(aq)}] = K_b \times \frac{[B_{(aq)}]}{[HB^+_{(aq)}]}$$

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$$[OH^-_{(aq)}] = 1.8 \times 10^{-5} \times \frac{0.267}{0.800}$$

$$[OH^-_{(aq)}] = 6.01 \times 10^{-6} \text{ mol dm}^{-3}$$

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

$$1 \times 10^{-14} = [H^+](6.01 \times 10^{-6})$$

$$[H^+] = 1.66 \times 10^{-9} \text{ mol dm}^{-3}$$

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

$$pH = -\log(1.66 \times 10^{-9})$$

$$pH = 8.78$$

Carbonic acid, H_2CO_3, is a weak Brønsted–Lowry acid formed when carbon dioxide dissolves in water. Healthy blood at a pH of 7.40 has a hydrogencarbonate : carbonic acid ratio of 10.5 : 1.

A patient is admitted to hospital. The patient's blood pH is measured as 7.20. Calculate the hydrogencarbonate : carbonic acid ratio in the patient's blood.

$[H^+_{(aq)}] = 10^{-pH}$ $[H^+_{(aq)}] = 10^{-7.40}$ $[H^+_{(aq)}] = 3.98 \times 10^{-8}$

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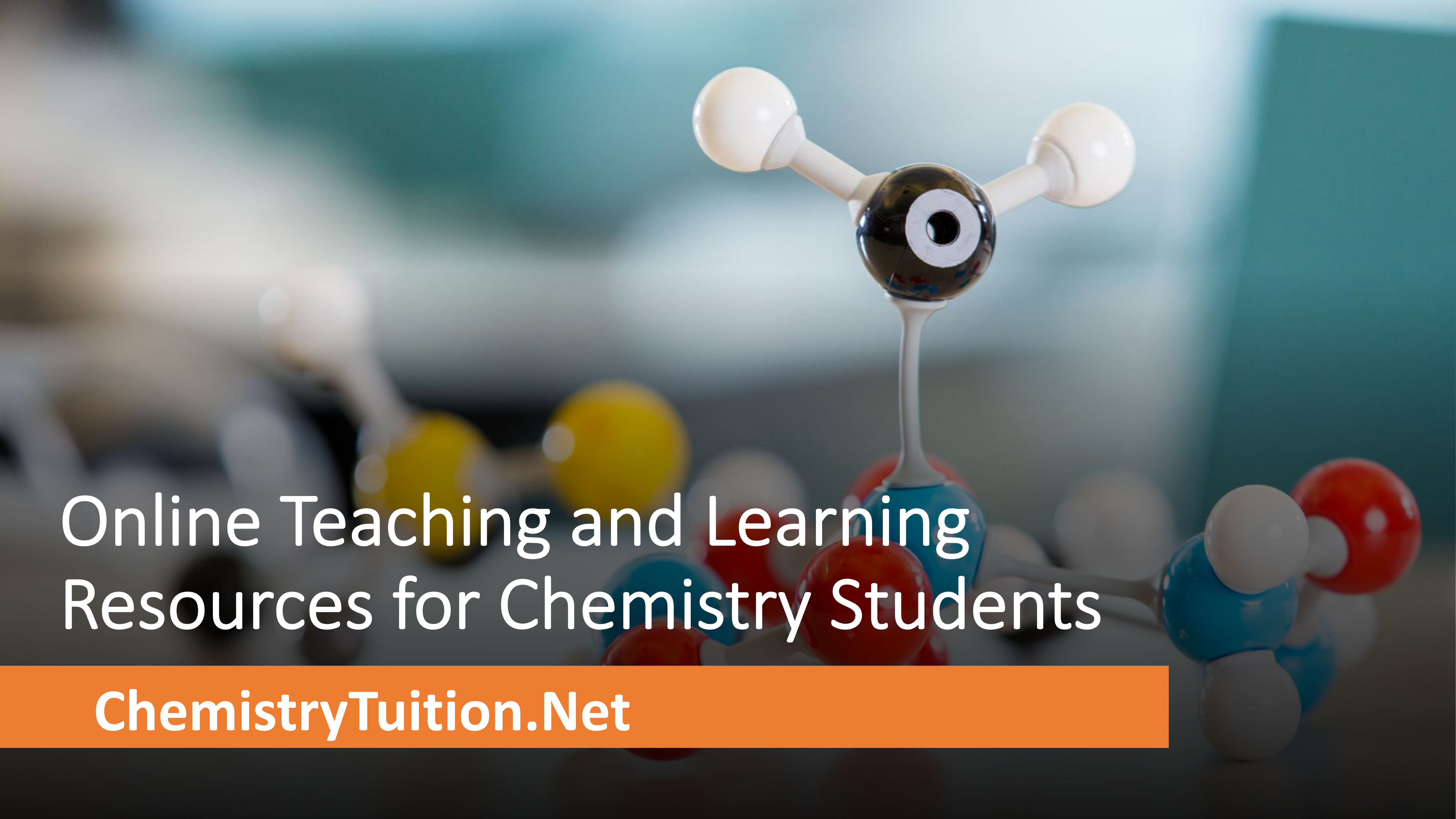

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