ANSWER KEY.

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8.9 Henderson-Hasselbalch Equation

AP Chemistry | ANSWER KEY

Show your work. Use pH = pK_a + log([A^-] / [HA]). Report pH to two decimal places.

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1. Calculate the pH of a buffer that is 0.200 *M* CH_3COOH and 0.300 *M* CH_3COONa. (K_a of CH_3COOH = 1.8×10^{-5})

pK_a = $-\log(1.8 \times 10^{-5})$ = 4.74

pH = pK_a + log([A^-] / [HA]) = 4.74 + log(0.300 / 0.200) = 4.74 + 0.176 = **4.92**

2. A chemist has 500.0 mL of 0.400 *M* NH_3 (K_b = 1.8×10^{-5}). What mass of NH_4Cl (53.49 g/mol) must be dissolved in it to make a buffer with pH 9.00? Assume the volume does not change.

pK_a of NH_4^+ = 14.00 − pK_b = 14.00 − 4.74 = 9.26

9.00 = 9.26 + log(0.400 / [NH_4^+]), so 0.400 / [NH_4^+] = $10^{-0.26}$ = 0.550 and [NH_4^+] = 0.728 *M*

mol NH_4Cl = (0.728 mol/L)(0.5000 L) = 0.364 mol; mass = (0.364 mol)(53.49 g/mol) = **19.5 g**

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3. A 1.00 L buffer contains 0.200 mol CH_3COOH and 0.300 mol CH_3COO^- (pK_a = 4.74). Assume no volume change.

(a) Calculate the pH after 0.0100 mol HCl is added.

(b) Calculate the pH after 0.0100 mol HCl is added to 1.00 L of pure water. Compare (a) and (b).

(a) HCl reacts with the base: CH_3COO^- = 0.300 − 0.0100 = 0.290 mol; CH_3COOH = 0.200 + 0.0100 = 0.210 mol

pH = 4.74 + log(0.290 / 0.210) = **4.88**

(b) $[H_3O^+]$ = 0.0100 *M*, so pH = **2.00**. The buffer moved from pH 4.92 to 4.88, while water moved from 7.00 to 2.00.

4. Blood plasma is buffered near pH 7.40 by $H_2PO_4^-$ and HPO_4^{2-} (pK_a = 7.21). What ratio $[HPO_4^{2-}] / [H_2PO_4^-]$ gives pH 7.40?

7.40 = 7.21 + log($[HPO_4^{2-}] / [H_2PO_4^-]$)

log(ratio) = 0.19, so ratio = $10^{0.19}$ = **1.5**

There is slightly more HPO_4^{2-} (the base form), as expected when pH is above pK_a.

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