Buffers and pH

1. What is the pH of a buffer that is 0.55 M in formic acid, HCOOH, and 0.63 M in sodium formate, NaHCOO?

   The K_a for formic acid is 1.8×10^{-4}, which corresponds to a pK_a of 3.75. The pH of the solution is

$$\mathrm{pH} = \mathrm{pK_a} + \log\frac{[\mathrm{HCOO^-}]}{[\mathrm{HCOOH}]} = 3.75 + \log\frac{0.63}{0.55} = 3.81$$

2. What is the ratio of hypobromite, BrO^-, to hypobromous acid, HBrO, in a buffer with a pH of 7.88?

   The K_a for hypobromous acid is 2.4×10^{-9}, which corresponds to a pK_a of 8.62. The solution's pH is

$$\mathrm{pH} = \mathrm{pK_a} + \log\frac{[\mathrm{OBr^-}]}{[\mathrm{HOBr}]} = 8.62 + \log\frac{[\mathrm{OBr^-}]}{[\mathrm{HOBr}]} = 7.88$$

   Solving for the ratio of conjugate weak base-to-conjugate weak acid gives

$$-0.74 = \log\frac{[\mathrm{OBr^-}]}{[\mathrm{HOBr}]} \quad \text{or} \quad \frac{[\mathrm{OBr^-}]}{[\mathrm{HOBr}]} = 0.18$$

3. Human blood contains two buffer systems, one based on phosphate species and one on carbonate species. If blood has a normal pH of 7.4, what are the principle phosphate and carbonate species present? What is the ratio between the two phosphate species? At the temperature of human blood, the K_a values for phosphoric acid are 1.3×10^{-2}, 2.3×10^{-7}, and 6×10^{-12}, respectively. The K_a values for carbonic acid are 8×10^{-7} and 1.6×10^{-10}.

   At the temperature of human blood, the pK_a values for phosphate are 1.89, 6.64, and 11.22 and the pK_a values for carbonic acid are 6.10 and 9.80. A pH of 7.40 falls within ±1 pH unit of phosphate's pK_{a2} of 6.63; thus, we expect there are significant amounts of both $H_2PO_4^-$ and HPO_4^{2-}. The relative abundance of these two species is

$$7.4 = 6.64 + \log\frac{[\mathrm{HPO_4^{2-}}]}{[\mathrm{H_2PO_4^-}]} \quad \text{or} \quad \frac{[\mathrm{HPO_4^{2-}}]}{[\mathrm{H_2PO_4^-}]} = 5.8$$

   A pH of 7.40 is more than 1 pH units above carbonic acid's pK_a of 6.10 and more than 1 pH units below its pK_a of 9.80; thus, the only important form of carbonic acid is HCO_3^-.