Question #79159, Chemistry / General Chemistry

Suppose you are performing an experiment that requires a constant pH of 7.00.Calculate the ratio of conjugate base to weak acid required to prepare a buffer with pH= 7.00.

Solution

To answer this question we should use Henderson-Hasselbalch equation:

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

Where pH is the concentration of $[H^+]$

pK$_a$ is the acid dissociation constant

$\frac{[A^-]}{[HA]}$ is the ratio of the concentrations of the conjugate base and starting acid.

So we know pH =7, and we need to choose conjugate base and weak acid to use it's pK$_a$ in the formula above.

We could consider phosphate buffer:

$H_2PO_4^- + H_2O \leftrightarrow HPO_4^{2-} + H_3O^+$

pK_a ($H_2PO_4^-$)= 7.2

Find ratio $\frac{[A^-]}{[HA]}$

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

$$7 = 7.2 + log\left(\frac{[A^-]}{[HA]}\right)$$

$$log\left(\frac{[A^-]}{[HA]}\right) = -0.2$$

$$\frac{[A^-]}{[HA]} = 10^{-0.2}$$

$$\frac{[A^-]}{[HA]} = 0.63$$

Answer: $\frac{[A^-]}{[HA]} = 0.63$ (for phosphate buffer)