Answer on Question #67225, Chemistry / General Chemistry

Commercial chlorine bleach, for example Clorox, consists of a 1.108 M aqueous solution of the salt sodium hypochlorite, NaClO. Determine the pH of this salt solution. Hypochlorous acid, HClO, has a K_a=2.9x10^{-8}. Is your numerical answer consistent with your understanding of what makes a salt acidic, basic, or neutral? Yes or No? Explain.

Solution:

$$NaClO \leftrightarrow Na^+ + ClO^-$$

Firstly we should take to attention that hypochlorite-ion ClO^- is a weak base:

$$ClO^- + H_2O \leftrightarrow HClO + OH^-$$

$$K_b(HClO) = \frac{K_w}{K_a} = \frac{10^{-14}}{2.9\times10^{-8}} = 3.4\times10^{-7}$$

Now we need to calculate the degree of dissociation:

$$\alpha = \sqrt{\frac{K_b}{C_B}}$$

$$\alpha = \sqrt{\frac{3.4\times10^{-7}}{1.108}} = 5.539\times10^{-4} = 0.00055$$

So $[ClO^-] = c_{ClO^-}$

Because α ≤ 0.05 we can use the next formula:

$$[H^+] = \frac{K_w}{\sqrt{K_{bC_B}}} = \frac{10^{-14}}{\sqrt{3.4\times10^{-7}\times1.108}} = 1.6\times10^{-11}$$

$$pH = -lg[H^+] = -\lg(1.6\times10^{-11}) = 10.8$$

Answer: 10.8