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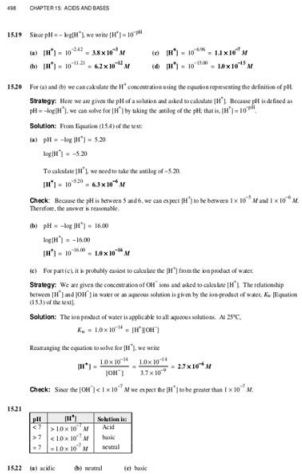

498 CHAPTER 15: ACIDS AND BASES

15.19 Since $pH = -\log[H^+]$, we write $[H^+] = 10^{-pH}$

(a) $[H^+] = 10^{-2.42} = 3.8 \times 10^{-3}\ M$ (c) $[H^+] = 10^{-6.06} = 1.1 \times 10^{-7}\ M$

(b) $[H^+] = 10^{-11.22} = 6.2 \times 10^{-12}\ M$ (d) $[H^+] = 10^{-15.00} = 1.0 \times 10^{-15}\ M$

15.20 For (a) and (b) we can calculate the H^+ concentration using the equation representing the definition of pH.

Strategy: Here we are given the pH of a solution and asked to calculate $[H^+]$. Because pH is defined as $pH = -\log[H^+]$, we can solve for $[H^+]$ by taking the antilog of the pH; that is, $[H^+] = 10^{-pH}$.

Solution: From Equation (15.4) of the text:

(a) $pH = -\log[H^+] = 5.20$

$\log[H^+] = -5.20$

To calculate $[H^+]$, we need to take the antilog of −5.20.

$[H^+] = 10^{-5.20} = 6.3 \times 10^{-6}\ M$

Check: Because the pH is between 5 and 6, we can expect $[H^+]$ to be between $1 \times 10^{-5}\ M$ and $1 \times 10^{-6}\ M$. Therefore, the answer is reasonable.

(b) $pH = -\log[H^+] = 16.00$

$\log[H^+] = -16.00$

$[H^+] = 10^{-16.00} = 1.0 \times 10^{-16}\ M$

(c) For part (c), it is probably easiest to calculate the $[H^+]$ from the ion product of water.

Strategy: We are given the concentration of OH^- ions and asked to calculate $[H^+]$. The relationship between $[H^+]$ and $[OH^-]$ in water or an aqueous solution is given by the ion-product of water, K_w [Equation (15.3) of the text].

Solution: The ion product of water is applicable to all aqueous solutions. At 25°C,

$K_w = 1.0 \times 10^{-14} = [H^+][OH^-]$

Rearranging the equation to solve for $[H^+]$, we write

$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{3.7 \times 10^{-9}} = 2.7 \times 10^{-6}\ M$

Check: Since the $[OH^-] < 1 \times 10^{-7}\ M$ we expect the $[H^+]$ to be greater than $1 \times 10^{-7}\ M$.

15.21

pH	$[H^+]$	Solution is:
< 7	$> 1.0 \times 10^{-7}\ M$	Acid
> 7	$< 1.0 \times 10^{-7}\ M$	basic
= 7	$= 1.0 \times 10^{-7}\ M$	neutral

15.22 (a) acidic (b) neutral (c) basic

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