CHEM 1423
Chapters 18
Homework Solutions

TEXTBOOK HOMEWORK

18.15 Lower $[H^+]$ ‖ higher pH

(a) $Ka = 4x10^{-5}$: Lower Ka ‖ lower $[H^+]$

(b) pKa = 3.5: Higher pKa ‖ lower Ka ‖ lower $[H^+]$

(c) 0.01 M: lower concentration ‖ lower $[H^+]$

(d) Weaker Acid: lower $[H^+]$

(e) 0.01 M Base: Base has lower $[H^+]$

(f) pOH = 6: Has higher pH (8)

18.16 (a) $[OH^-] = 0.0111$ M ‖ $pOH = -\log(0.0111) = 1.95 \rightarrow pH = 14 - 1.95 \approx 12.05$ Basic

(b) $[H^+] = 1.35x10^{-3}$ M ‖ $pH = -\log(1.35\,x10^{-3}) = 2.87 \rightarrow pOH = 14 - 2.87 \approx 11.13$ Acidic

18.18 (a) pH = 9.85 ‖ pOH = 14 - 9.85 = 4.15 , $[H^+] = 10^{-pH} = 10^{-9.85} = 1.4x10^{-10}\,M$
$[\mathrm{OH}^-] = 10^{-pOH} = 10^{-4.15} = 7.1x10^{-5}\,M$

(b) pOH = 9.43 ‖ pH = 14 - 9.43 = 4.57 , $[H^+] = 10^{-pH} = 10^{-4.57} = 2.7x10^{-5}\,M$
$[\mathrm{OH}^-] = 10^{-pOH} = 10^{-9.43} = 3.7x10^{-10}\,M$

18.20 **Goal:** Calculate initial number of moles of H^+ and final number of moles of H^+. The difference is the number of moles of OH^- which must be added.

$[H^+]_{init} = 10^{-pH_{init}} = 10^{-4.52} = 3.020x10^{-5}\,M$

$n[H^+]_{init} = [\mathrm{H}^+]_{init}V = 3.020x10^{-5}\,M(5.60\,L) = 1.691x10^{-4}\,mol$

$[H^+]_{fin} = 10^{-pH_{fin}} = 10^{-5.25} = 5.623x10^{-6}\,M$

$n[H^+]_{fin} = [\mathrm{H}^+]_{fin}V = 5.623x10^{-6}\,M(5.60\,L) = 3.15x10^{-5}\,mol$

$\Delta\,\mathrm{n}[H^+] = n[H^+]_{fin} - n[H^+]_{init} = 3.149x10^{-5} - 1.691x10^{-4} = -1.38x10^{-4}\,mol$

In order to remove $1.38x10^{-4}$ moles of H^+ from the solution, we must **add** $1.38x10^{-4}$ moles of $[OH^-]$.

18.45 $HA \underset{}{\overset{K_a}{\rightleftarrows}} H^+ + A^-$

$[H^+] = [A^-] = x = 10^{-pH} = 10^{-4.88} = 1.32x10^{-5}$ M , [HA] = 0.035-x ≈ 0.035

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(1.32x10^{-5})^2}{0.035} = 4.97x10^{-9} \approx 5.0x10^{-9}$$

18.48 pKa = 2.87 ‖ $K_a = 10^{-pKa} = 10^{-2.87} = 1.35x10^{-3}$

$$ClCH_2COOH \overset{K_a}{\rightleftarrows} H^+ + ClCH_2COO^-$$

$[ClCH_2COO^-] = [H^+] = x$, $[ClCH_2COOH] = 1.25 - x \approx 1.25$ M**

$$K_a = 1.35x10^{-3} = \frac{[H^+][\mathrm{ClCH_2COO^-}]}{[\mathrm{ClCH_2COOH}]} = \frac{x^2}{1.25}$$

$$x^2 = 1.35x10^{-3}(1.25) = 1.69x10^{-3} \rightarrow x = 4.11x10^{-2} \approx 0.041M$$

$[H^+] = [ClCH_2COO^-] = x = 0.041$ M , $[ClCH_2COOH] = 1.25 - x = 1.21$ M ≈ 1.25 M

$$pH = -\log(0.041) \approx 1.4$$

18.50 $HA \overset{K_a}{\rightleftarrows} H^+ + A^-$ $[HA]_o = 0.20$ M , 3% dissociated

(a) **[H⁺] , [OH⁻], pH, pOH**

$$[H^+] = [A^-] = \frac{3}{100}\left[HA\right]_o = \frac{3}{100}\left(0.20M\right) = 0.006M$$

$$[\mathrm{HA}] = \frac{97}{100}\left[HA\right]_o = \frac{97}{100}\left(0.20M\right) = 0.194M$$

$$pH = -\log(0.006) = 2.22$$

$$pOH = 14 - 2.22 = 11.78$$

$$[OH^-] = 10^{-11.78} = 1.66x10^{-12}\ M \approx 1.7x10^{-12}\ M$$

(b) **K_a**

$$K_a = \frac{[H^+][\mathrm{A}^-]}{[HA]} = \frac{(0.006)(0.006)}{0.194} = 1.86x10^{-4} \approx 1.9x10^{-4}$$

18.54 For convenience, let's call $HC_9H_7O_4 \equiv$ HAsp: $HAsp \overset{K_a}{\rightleftarrows} H^+ + Asp^-$, $K_a = 3.6x10^{-4}$

$[H^+] = [Asp^-] = x$, [HAsp] = 0.018 - x ≈ 0.018

$$K_a = 3.6x10^{-4} = \frac{[H^+][\mathrm{Asp^-}]}{[\mathrm{HAsp}]} = \frac{x^2}{0.018}$$

$$x^2 = 3.6x10^{-4}(0.018) = 6.48x10^{-6} \rightarrow x = 2.55x10^{-3} = [H^+]$$

$$pH = -\log\left(2.55x10^{-3}\right) = 2.59 \approx 2.6$$

18.55 Formic Acid: HCOOH. For convenience, we'll call this HForm.

$K_a = 1.8x10^{-4}$. This value was looked up in a table. It would be given to you on a test.

$$HForm \xrightleftharpoons{K_a} H^+ + Form^-$$

$[H^+] = [Form^-] = x$, $[HForm] = 0.75 - x \approx 0.75$

$$K_a = 1.8x10^{-4} = \frac{[H^+][\text{Form}^-]}{[\text{HForm}]} = \frac{x^2}{0.75}$$

$$x^2 = 1.8x10^{-4}(0.75) = 1.35x10^{-4} \rightarrow x = 1.16x10^{-2} = [\text{Form}^-]$$

The concentration of the dissociated species is $[Form^-]$. Therefore, the percent dissociation is given by: $\%Dissoc = \frac{[\text{Form}^-]}{[HForm]_o} x100 = \frac{1.16x10^{-2}}{0.75} x100 = 1.55\% \approx 1.5\%$

18.65 (a) KCN: The salt dissociates completely to $K^+(aq)$ and $CN^-(aq)$.

KCN ‖ $K^+ + CN^-$(0.15 M). The K^+ is unimportant because it is the conjugate acid of the infiniely strong base, KOH, and does not react with water.

CN^- is the conjugate base of the weak acid, HCN, $K_a = 6.2x10^{-10}$ (Appendix C of the text). You would be given the value on a test.

Note: Because you have just put the conjugate base of a weak acid in solution, you know that your solution should have pH > 7.

CN^- reacts with water according to: $CN^- + H_2O \xrightleftharpoons{K_b} HCN + OH^-$

$[HCN] = [OH^-] = x$, $[CN^-] = 0.15 - x \approx 0.15$ M

We get $K_b(CN^+)$ from:

$$K_aK_b = K_w \rightarrow K_b(CN^-) = \frac{K_w}{K_a(HCN)} = \frac{1x10^{-14}}{6.2x10^{-10}} = 1.61x10^{-5}$$

$$K_b = \frac{[HCN][OH^-]}{[CN^-]} \rightarrow 1.61x10^{-5} = \frac{x \cdot x}{0.15}$$

$$x^2 = (0.15)(1.61x10^{-5}) = 2.42x10^{-6} \rightarrow x = 1.55x10^{-3} = [OH^-]$$

$$pOH = -\log([OH^-] = -\log(1.55\,x10^{-3}) = 2.81$$

$$pH = 14 - 2.81 = 11.19 \approx 11.2$$

(b) Triethylammonium Chloride, $(CH_3CH_2)_3NHCl$. Let's call this Et_3NHCl for convenience.

$Et_3NHCl \rightarrow Et_3NH^+(0.40\text{ M}) + Cl^-$. The Cl^- is unimportant because it is the conjugate base of the infinitely strong acid, HCl.

Et_3NH^+ is the conjugate acid of the weak base, Et_3N, $K_b = 5.2\text{x}10^{-4}$ (Appendix C of the text). You would be given this value on a test.

Note: Because you have just put the conjugate base of a weak acid in solution, you know that your solution should have pH > 7.

Et_3NH^+ dissociates according to: $Et_3NH^+ \xrightleftharpoons{K_x} H^+ + Et_3N$

$[H^+] = [Et_3N] = x$, $[Et_3NH^+] = 0.40 - x \approx 0.40$

We get $K_a(Et_3NH^+)$ from:

$$K_aK_b = K_w \rightarrow K_a(\mathrm{Et_3NH^+}) = \frac{K_w}{K_b(\mathrm{Et_3N})} = \frac{1x10^{-14}}{5.2x10^{-4}} = 1.92x10^{-11}$$

$$K_a = \frac{[H^+[\mathrm{Et}_3\,N]}{[\mathrm{Et_3NH^+}]} \rightarrow 1.92x10^{-11} = \frac{x \cdot x}{0.40}$$

$$x^2 = (0.40)(1.92x10^{-11}) = 7.69x10^{-12} \rightarrow x = 2.77x10^{-6} = [H^+]$$

$$\mathrm{pH} = -\log(2.77\,\mathrm{x}10^{-6}) = 5.56 \approx 5.6$$

18.67 NaClO dissociates completely into Na^+ + ClO^-. Because Na^+ is the conjugate acid of the infinitely strong base, NaOH, it does nothing in solution. On the other hand, ClO^- is the conjugate base of the weak acid, HClO. $K_a = 2.9x10^{-8}$ (Appendix C of the text). You would be given this value on a test.

Let's go ahead and get K_b, which we will need in the problem.

$$K_aK_b = K_w \rightarrow K_b(ClO^-) = \frac{K_w}{K_a(HClO)} = \frac{1x10^{-14}}{2.9x10^{-8}} = 3.45x10^{-7}$$

Strategy: (1) We will use the mass percent of NaClO and the density of the solution to calculate the initial concentration, [ClO$^-$]. (2) We will use the base equilibrium equation to determine [OH$^-$] and pOH/pH.

(1) [ClO$^-$]

Assume 1.0 L = 1000 mL of solution. Because d = 1 g/mL, we have 1,000 grams.
M(NaClO) = 74.5 g/mol.

$$mass(NaClO) = \frac{6.5}{100}(1,000\,g) = 65.0\,g$$

$$n(NaClO) = 65.0\,g \cdot \frac{1\,mol}{74.5\,g} = 0.87\,mol$$

Because NaClO dissociates completely to ClO^- (+ Na^+, which is unimportant), $n(ClO^-)$ = 0.87 mol. Therefore, $[ClO^-]_o$ = 0.87 mol/1.0 L = 0.87 M.

(2) Determine [OH$^-$], pOH, pH

ClO^- reacts with water according to: $ClO^- + H_2O \xrightleftharpoons{K_b} HClO + OH^-$

$[HClO] = [OH^-] = x$, $[ClO^-] = 0.87 - x \approx 0.87$ M

$$K_b = \frac{[HClO][OH^-]}{[ClO^-]} \rightarrow 3.45x10^{-7} = \frac{x \cdot x}{0.87}$$

$$x^2 = (0.87)(3.45x10^{-7}) = 3.00x10^{-7} \rightarrow x = 5.48x10^{-4} = [OH^-]$$

$$pOH = -\log([OH^-] = -\log(5.48\,x10^{-4}) = 3.26$$

$$pH = 14 - 3.26 = 10.74 \approx 10.7$$