

AP Chemistry – Acid Base (Kotz)

Homework Answers 3, 11, 15, 21, 27, 31, 35, 39, 41, 53, 59, 61, 67, 70, 99, 100

- 17.3 (a)  $\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$   
 acid A base B conjugate acid of B conjugate base of A
- (b)  $\text{HSO}_4^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$   
 acid A base B conjugate acid of B conjugate base of A
- (c)  $\text{H}_3\text{O}^+ + \text{F}^- \rightarrow \text{HF} + \text{H}_2\text{O}$   
 acid A base B conjugate acid of B conjugate base of A

17.11 HCl is a strong acid so  $[\text{H}_3\text{O}^+] = [\text{HCl}] = 0.0075 \text{ M}$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(0.0075) = 2.12$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.0075} = 1.3 \times 10^{-12} \text{ M}$$

- 17.15 (a) The strongest acid is  $\text{HCO}_2\text{H}$  (largest  $K_a$ ) and the weakest acid is  $\text{C}_6\text{H}_5\text{OH}$  (smallest  $K_a$ ).  
 (b) The strongest acid ( $\text{HCO}_2\text{H}$ ) has the weakest conjugate base.  
 (c) The weakest acid ( $\text{C}_6\text{H}_5\text{OH}$ ) has the strongest conjugate base.

17.21  $\text{Na}^+$ ,  $\text{CH}_3\text{CO}_2^-$ , and  $\text{Cl}^-$  are neutral ions.  $\text{S}^{2-}$ ,  $\text{PO}_4^{3-}$ , and  $\text{F}^-$  are basic ions.  $\text{H}_2\text{PO}_4^-$  and  $\text{Al}^{3+}$  are acidic ions. According to Table 17.3,  $\text{S}^{2-}$  is most basic. A solution of (a)  $\text{Na}_2\text{S}$  will have the highest pH. According to Table 17.3,  $\text{Al}(\text{H}_2\text{O})_6^{3+}$  is most acidic. A solution of (f)  $\text{AlCl}_3$  will have the lowest pH.

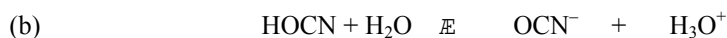
17.27 The acid with the smaller  $\text{p}K_a$  has the larger  $K_a$ . (b) 2-chlorobenzoic acid is the stronger acid

17.31  $K_a = 10^{-\text{p}K_a} = 10^{-9.80} = 1.6 \times 10^{-10}$        $K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.6 \times 10^{-10}} = 6.3 \times 10^{-5}$

- 17.35 (a) HBr is a stronger acid than  $\text{NH}_4^+$ , so the equilibrium lies predominantly to the left.  
 (b)  $\text{CH}_3\text{CO}_2\text{H}$  is a stronger acid than  $\text{HPO}_4^{2-}$ , so the equilibrium lies predominantly to the left.  
 (c)  $\text{Fe}(\text{H}_2\text{O})_6^{3+}$  is a stronger acid than  $\text{H}_2\text{CO}_3$ , so the equilibrium lies predominantly to the right.

17.39 (a)  $\text{CH}_3\text{CO}_2\text{H}(\text{aq}) + \text{HPO}_4^{2-}(\text{aq}) \rightleftharpoons \text{CH}_3\text{CO}_2^-(\text{aq}) + \text{H}_2\text{PO}_4^-(\text{aq})$   
 (b)  $\text{CH}_3\text{CO}_2\text{H}$  is a stronger acid than  $\text{H}_2\text{PO}_4^-$ . The equilibrium will lie predominantly to the right.  
 $K_a(\text{H}_2\text{PO}_4^-) > K_b(\text{CH}_3\text{CO}_2^-)$  The solution will be slightly acidic.

17.41 (a)  $[\text{H}_3\text{O}^+] = [\text{OCN}^-] = 10^{-\text{pH}} = 10^{-2.67} = 0.0021 \text{ M}$



<i>Initial</i> (M)	0.015	0	0
<i>Change</i> (M)	-0.0021	+0.0021	+0.0021
<i>Equilibrium</i> (M)	0.013	0.0021	0.0021

$$K_a = \frac{[\text{OCN}^-][\text{H}_3\text{O}^+]}{[\text{HOCN}]} = \frac{(0.0021)(0.0021)}{0.013} = 3.6 \times 10^{-4}$$

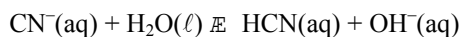
$$17.53 \quad K_b = 4.2 \times 10^{-4} = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{x^2}{0.25 - x}$$

Assume  $x$  is much smaller than 0.25

$$4.2 \times 10^{-4} = \frac{x^2}{0.25}$$

$$x = [\text{OH}^-] = 0.010 \text{ M} \quad \text{pOH} = -\log[\text{OH}^-] = 1.99 \quad \text{pH} = 14.00 - \text{pOH} = 12.01$$

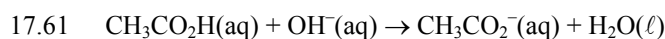
$$17.59 \quad [\text{CN}^-]_{\text{initial}} = \frac{10.8 \text{ g}}{0.500 \text{ L}} \cdot \frac{1 \text{ mol NaCN}}{49.01 \text{ g}} \cdot \frac{1 \text{ mol CN}^-}{1 \text{ mol NaCN}} = 0.441 \text{ M}$$



$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-10}} = 2.5 \times 10^{-5}$$

$$2.5 \times 10^{-5} = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} = \frac{x^2}{0.441 - x} \approx \frac{x^2}{0.441}$$

$$x = [\text{HCN}] = [\text{OH}^-] = 0.0033 \text{ M} \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = 3.0 \times 10^{-12} \quad [\text{CN}^-] = [\text{Na}^+] = 0.441 \text{ M}$$

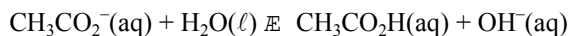


$$(0.0220 \text{ L CH}_3\text{CO}_2\text{H})(0.15 \text{ mol/L}) = 0.0033 \text{ mol CH}_3\text{CO}_2\text{H}$$

$$(0.0220 \text{ L NaOH})(0.15 \text{ mol/L}) = 0.0033 \text{ mol NaOH}$$

$$0.0033 \text{ mol CH}_3\text{CO}_2\text{H} \cdot \frac{1 \text{ mol CH}_3\text{CO}_2^-}{1 \text{ mol CH}_3\text{CO}_2\text{H}} = 0.0033 \text{ mol CH}_3\text{CO}_2^-$$

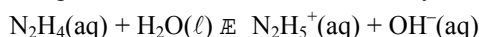
$$[\text{CH}_3\text{CO}_2^-] = \frac{0.0033 \text{ mol}}{(0.0220 + 0.0220) \text{ L}} = 0.075 \text{ M}$$



$$K_b = 5.6 \times 10^{-10} = \frac{[\text{CH}_3\text{CO}_2\text{H}][\text{OH}^-]}{[\text{CH}_3\text{CO}_2^-]} = \frac{x^2}{0.075 - x} \approx \frac{x^2}{0.075}$$

$$x = [\text{OH}^-] = 6.5 \times 10^{-6} \text{ M} \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = 1.5 \times 10^{-9} \quad \text{pH} = -\log[\text{H}_3\text{O}^+] = 8.81$$

17.67 The pH of the solution is determined by the first ionization of the base.



$$(a) \quad K_{b1} = 8.5 \times 10^{-7} = \frac{[\text{N}_2\text{H}_5^+][\text{OH}^-]}{[\text{N}_2\text{H}_4]} = \frac{x^2}{0.010 - x} \approx \frac{x^2}{0.010}$$

$$x = [\text{N}_2\text{H}_5^+] = [\text{OH}^-] = 9.2 \times 10^{-5} \text{ M} \quad [\text{N}_2\text{H}_6^{2+}] = K_{b2} = 8.9 \times 10^{-16} \text{ M}$$

$$(b) \quad \text{pOH} = -\log[\text{OH}^-] = 4.04 \quad \text{pH} = 14.00 - \text{pOH} = 9.96$$

17.70 (a)  $\text{BCl}_3$  is a Lewis acid.

(b)  $\text{H}_2\text{NNH}_2$  is a Lewis base.

(c)  $\text{Ag}^+$  is a Lewis acid,  $\text{NH}_3$  is a Lewis base

- 17.99 [A]<sub>1</sub> = initial concentration of weak acid in original solution  
 [A]<sub>2</sub> = initial concentration of weak acid in diluted solution  
 [H<sub>3</sub>O<sup>+</sup>]<sub>1</sub> = equilibrium hydronium ion concentration in original solution  
 [H<sub>3</sub>O<sup>+</sup>]<sub>2</sub> = equilibrium hydronium ion concentration in diluted solution

If the fraction ionized doubles, then [H<sub>3</sub>O<sup>+</sup>]<sub>2</sub> = 2 · [H<sub>3</sub>O<sup>+</sup>]<sub>1</sub>

$$\text{Because } K \approx \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]_{\text{initial}}}$$

$$\frac{[\text{H}_3\text{O}^+]_1^2}{[\text{A}]_1} = \frac{[\text{H}_3\text{O}^+]_2^2}{[\text{A}]_2}$$

$$\frac{[\text{H}_3\text{O}^+]_1^2}{[\text{A}]_1} = \frac{(2 \cdot [\text{H}_3\text{O}^+]_1)^2}{[\text{A}]_2}$$

$$\frac{[\text{A}]_2}{[\text{A}]_1} = 4$$

In order to double the percent ionization of the acid, you must dilute 100 mL of solution to 400 mL.

- 17.100 (a) BF<sub>3</sub> is a Lewis acid, (CH<sub>3</sub>)<sub>2</sub>O is a Lewis base

$$(b) P_{(\text{CH}_3)_2\text{OBF}_3} = \frac{nRT}{V} = \frac{\left(0.100 \text{ g} \cdot \frac{1 \text{ mol } (\text{CH}_3)_2\text{OBF}_3}{113.9 \text{ g}}\right) (0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})}{0.565 \text{ L}} = 0.0380 \text{ atm}$$

$$K_p = 0.17 = \frac{P_{\text{BF}_3} P_{(\text{CH}_3)_2\text{O}}}{P_{(\text{CH}_3)_2\text{OBF}_3}} = \frac{x^2}{0.0380 - x}$$

Solve using the quadratic equation

$$x = P_{\text{BF}_3} = P_{(\text{CH}_3)_2\text{O}} = 0.032 \text{ atm} \quad P_{(\text{CH}_3)_2\text{OBF}_3} = 0.0380 - x = 0.006 \text{ atm}$$

$$P_{\text{total}} = 0.032 \text{ atm} + 0.032 \text{ atm} + 0.006 \text{ atm} = 0.0700 \text{ atm}$$