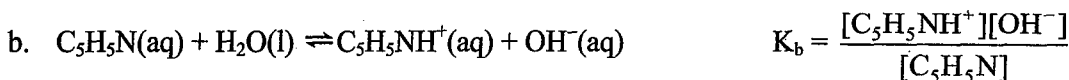
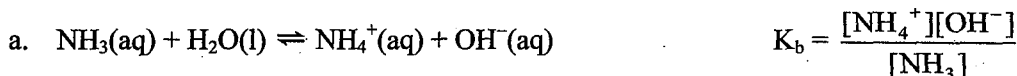


## Solutions of Bases

83. All  $K_b$  reactions refer to the base reacting with water to produce the conjugate acid of the base and  $\text{OH}^-$ .



85.  $\text{NO}_3^-$ : Because  $\text{HNO}_3$  is a strong acid,  $\text{NO}_3^-$  is a terrible base ( $K_b \ll K_w$ ). All conjugate bases of strong acids have no base strength.

$$\text{H}_2\text{O}: K_b = K_w = 1.0 \times 10^{-14}; \quad \text{NH}_3: K_b = 1.8 \times 10^{-5}; \quad \text{C}_5\text{H}_5\text{N}: K_b = 1.7 \times 10^{-9}$$

Base strength =  $\text{NH}_3 > \text{C}_5\text{H}_5\text{N} > \text{H}_2\text{O} > \text{NO}_3^-$  (As  $K_b$  increases, base strength increases.)

87. a.  $\text{C}_6\text{H}_5\text{NH}_2$       b.  $\text{C}_6\text{H}_5\text{NH}_2$       c.  $\text{OH}^-$       d.  $\text{CH}_3\text{NH}_2$

The base with the largest  $K_b$  value is the strongest base ( $K_{b, \text{C}_6\text{H}_5\text{NH}_2} = 3.8 \times 10^{-10}$ ,  $K_{b, \text{CH}_3\text{NH}_2} = 4.4 \times 10^{-4}$ ).  $\text{OH}^-$  is the strongest base possible in water.

89.  $\text{NaOH}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$ ;  $\text{NaOH}$  is a strong base that completely dissociates into  $\text{Na}^+$  and  $\text{OH}^-$ . The initial concentration of  $\text{NaOH}$  will equal the concentration of  $\text{OH}^-$  donated by  $\text{NaOH}$ .

$$\text{a. } [\text{OH}^-] = 0.10 \text{ M}; \quad \text{pOH} = -\log[\text{OH}^-] = -\log(0.10) = 1.00$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 1.00 = 13.00$$

Note that  $\text{H}_2\text{O}$  is also present, but the amount of  $\text{OH}^-$  produced by  $\text{H}_2\text{O}$  will be insignificant compared to the  $0.10 \text{ M OH}^-$  produced from the  $\text{NaOH}$ .

- b. The  $[\text{OH}^-]$  concentration donated by the  $\text{NaOH}$  is  $1.0 \times 10^{-10} \text{ M}$ . Water by itself donates  $1.0 \times 10^{-7} \text{ M}$ . In this exercise, water is the major  $\text{OH}^-$  contributor, and  $[\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$ .

$$\text{pOH} = -\log(1.0 \times 10^{-7}) = 7.00; \quad \text{pH} = 14.00 - 7.00 = 7.00$$

$$\text{c. } [\text{OH}^-] = 2.0 \text{ M}; \quad \text{pOH} = -\log(2.0) = -0.30; \quad \text{pH} = 14.00 - (-0.30) = 14.30$$

91. a. Major species:  $\text{K}^+$ ,  $\text{OH}^-$ ,  $\text{H}_2\text{O}$  ( $\text{KOH}$  is a strong base.)

$$[\text{OH}^-] = 0.015 \text{ M}, \quad \text{pOH} = -\log(0.015) = 1.82; \quad \text{pH} = 14.00 - \text{pOH} = 12.18$$

91. b. Major species:  $\text{Ba}^{2+}$ ,  $\text{OH}^-$ ,  $\text{H}_2\text{O}$ ;  $\text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq})$ ; because each mole of the strong base  $\text{Ba}(\text{OH})_2$  dissolves in water to produce two mol  $\text{OH}^-$ ,  $[\text{OH}^-] = 2(0.015 \text{ M}) = 0.030 \text{ M}$ .

$$\text{pOH} = -\log(0.030) = 1.52; \text{pH} = 14.00 - 1.52 = 12.48$$

93.  $\text{pOH} = 14.00 - 11.56 = 2.44$ ;  $[\text{OH}^-] = [\text{KOH}] = 10^{-2.44} = 3.6 \times 10^{-3} \text{ M}$

$$0.8000 \text{ L} \times \frac{3.6 \times 10^{-3} \text{ mol KOH}}{\text{L}} \times \frac{56.11 \text{ g KOH}}{\text{mol KOH}} = 0.16 \text{ g KOH}$$

95.  $\text{NH}_3$  is a weak base with  $K_b = 1.8 \times 10^{-5}$ . The major species present will be  $\text{NH}_3$  and  $\text{H}_2\text{O}$  ( $K_b = K_w = 1.0 \times 10^{-14}$ ). Because  $\text{NH}_3$  has a much larger  $K_b$  value than  $\text{H}_2\text{O}$ ,  $\text{NH}_3$  is the stronger base present and will be the major producer of  $\text{OH}^-$ . To determine the amount of  $\text{OH}^-$  produced from  $\text{NH}_3$ , we must perform an equilibrium calculation using the  $K_b$  reaction for  $\text{NH}_3$ .



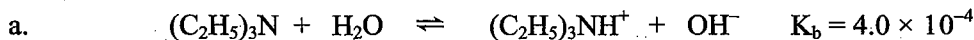
Initial	0.150 M	0	~0
	x mol/L $\text{NH}_3$ reacts with $\text{H}_2\text{O}$ to reach equilibrium		
Change	-x	+x	+x
Equil.	0.150 - x	x	x

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{x^2}{0.150 - x} \approx \frac{x^2}{0.150} \quad (\text{assuming } x \ll 0.150)$$

$x = [\text{OH}^-] = 1.6 \times 10^{-3} \text{ M}$ ; check assumptions:  $x$  is 1.1% of 0.150, so the assumption  $0.150 - x \approx 0.150$  is valid by the 5% rule. Also, the contribution of  $\text{OH}^-$  from water will be insignificant (which will usually be the case). Finishing the problem:

$$\text{pOH} = -\log[\text{OH}^-] = -\log(1.6 \times 10^{-3} \text{ M}) = 2.80; \text{pH} = 14.00 - \text{pOH} = 14.00 - 2.80 = 11.20.$$

97. These are solutions of weak bases in water. In each case we must solve an equilibrium weak base problem.



Initial	0.20 M	0	~0
	x mol/L of $(\text{C}_2\text{H}_5)_3\text{N}$ reacts with $\text{H}_2\text{O}$ to reach equilibrium		
Change	-x	+x	+x
Equil.	0.20 - x	x	x

$$K_b = 4.0 \times 10^{-4} = \frac{[(\text{C}_2\text{H}_5)_3\text{NH}^+][\text{OH}^-]}{[(\text{C}_2\text{H}_5)_3\text{N}]} = \frac{x^2}{0.20 - x} \approx \frac{x^2}{0.20}, \quad x = [\text{OH}^-] = 8.9 \times 10^{-3} \text{ M}$$

Assumptions good ( $x$  is 4.5% of 0.20).  $[\text{OH}^-] = 8.9 \times 10^{-3} \text{ M}$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{8.9 \times 10^{-3}} = 1.1 \times 10^{-12} \text{ M}; \text{pH} = 11.96$$

97



Initial	0.20 M	0	~0
Equil.	0.20 - x	x	x

$$K_b = 1.1 \times 10^{-8} = \frac{x^2}{0.20 - x} \approx \frac{x^2}{0.20}, \quad x = [\text{OH}^-] = 4.7 \times 10^{-5} \text{ M}; \quad \text{assumptions good.}$$

$$[\text{H}^+] = 2.1 \times 10^{-10} \text{ M}; \quad \text{pH} = 9.68$$

99. This is a solution of a weak base in water. We must solve the weak base equilibrium problem.



Initial	0.20 M	0	~0
<i>x</i> mol/L C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub> reacts with H <sub>2</sub> O to reach equilibrium			
Change	-x	+x	+x
Equil.	0.20 - x	x	x

$$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]} = \frac{x^2}{0.20 - x} \approx \frac{x^2}{0.20} \quad (\text{assuming } x \ll 0.20)$$

$$x = 1.1 \times 10^{-2}; \quad \text{checking assumption: } \frac{1.1 \times 10^{-2}}{0.20} \times 100 = 5.5\%$$

The assumption fails the 5% rule. We must solve exactly using either the quadratic equation or the method of successive approximations (see Appendix 1 of the text). Using successive approximations and carrying extra significant figures:

$$\frac{x^2}{0.20 - 0.011} = \frac{x^2}{0.189} = 5.6 \times 10^{-4}, \quad x = 1.0 \times 10^{-2} \text{ M} \quad (\text{consistent answer})$$

$$x = [\text{OH}^-] = 1.0 \times 10^{-2} \text{ M}; \quad [\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12} \text{ M}; \quad \text{pH} = 12.00$$

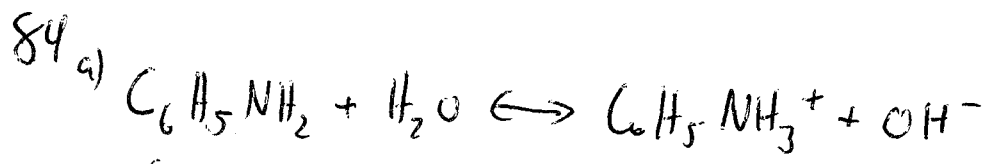
101. To solve for percent ionization, we first solve the weak base equilibrium problem.



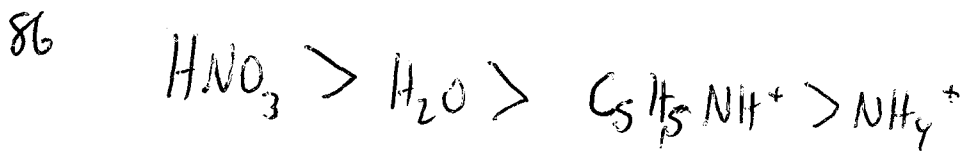
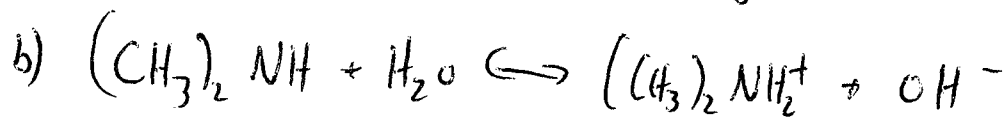
Initial	0.10 M	0	~0
Equil.	0.10 - x	x	x

$$K_b = 1.8 \times 10^{-5} = \frac{x^2}{0.10 - x} \approx \frac{x^2}{0.10}, \quad x = [\text{OH}^-] = 1.3 \times 10^{-3} \text{ M}; \quad \text{assumptions good.}$$

$$\text{Percent ionization} = \frac{x}{[\text{NH}_3]_0} \times 100 = \frac{1.3 \times 10^{-3} \text{ M}}{0.10 \text{ M}} \times 100 = 1.3\%$$



14.6a



Largest  $K_a$

Smallest  $K_b$