

## Chapter 6

### Acids and Bases

1. Define a conjugate acid-base pair.

A conjugate acid-base pair is a weak acid and a weak base that differ by only one proton.

3. Indicate the conjugate base for each of the following:

- a)  $\text{OH}^{1-}$     $\text{O}^{2-}$                       b)  $\text{H}_2\text{O}_2$     $\text{HO}_2^{1-}$                       c)  $\text{H}_2\text{PO}_4^{1-}$     $\text{HPO}_4^{2-}$   
d)  $\text{H}_3\text{O}^{1+}$     $\text{H}_2\text{O}$                       e)  $\text{H}_2\text{SO}_3$     $\text{HSO}_3^{1-}$

5. Are all Brønsted bases also Lewis bases? Explain.

Yes! Both types of base must have a lone pair to form a covalent bond to an acid.

7. Which of the following are Brønsted bases?

- a)  $\text{NaOH}$  yes                              b)  $\text{NaCl}$  not in water                      c)  $\text{CH}_3\text{OH}$  no  
d)  $\text{KCN}$  yes                              e)  $\text{KH}_2\text{PO}_4$  yes

9. Explain how the reaction  $\text{Ag}^{1+} + \text{Cl}^{1-} \rightarrow \text{AgCl}$  is a Lewis acid-base reaction. Is it a Brønsted acid-base reaction?

$\text{Ag}^{1+}$  is a Lewis acid that accepts a lone pair on  $\text{Cl}^{1-}$  to form a bond that has substantial covalent character. It is not a Brønsted acid-base reaction because no proton is transferred.

11. Write Brønsted acid-base reactions or indicate no reaction if  $K \ll 1$ .

- a) Aqueous sodium sulfate is added to hydrobromic acid.  
 $\text{SO}_4^{2-} + \text{H}_3\text{O}^{1+} \rightarrow \text{HSO}_4^{1-} + \text{H}_2\text{O}$  (weak base + strong acid)
- b) Aqueous  $\text{NH}_4\text{Cl}$  and aqueous  $\text{KF}$  are mixed.  
 $\text{NH}_4^{1+} + \text{F}^{1-} \rightarrow$  no reaction (HF, the produced acid, is much stronger than the reacting acid)
- c) Aqueous sodium cyanide is added a large excess of sulfurous acid.  
 $\text{CN}^{1-} + \text{H}_2\text{SO}_3 \rightarrow \text{HSO}_3^{1-} + \text{HCN}$
- d) Acetic acid and aqueous sodium hypochlorite are mixed.  
 $\text{CH}_3\text{COOH} + \text{OCl}^{1-} \rightarrow \text{CH}_3\text{COO}^{1-} + \text{HOCl}$  (produced acid is weaker than reacting acid)
- e) Hydrogen sulfide is bubbled into water.  
 $\text{H}_2\text{S} + \text{H}_2\text{O} \rightarrow$  no reaction (this is the  $K_a$  reaction of a weak acid and the extent is very small)

13. Explain why  $\text{HCl}$  is a strong acid, but  $\text{HF}$  is a weak acid.

The HF bond is much stronger than the HCl bond.

15. What is the predominate phosphorus containing species in a solution prepared by adding some phosphoric acid to a solution containing a large excess of ammonia?

$\text{NH}_3$  is strong enough to remove a proton from  $\text{H}_3\text{PO}_4$  and  $\text{H}_2\text{PO}_4^{1-}$ , but it is too weak to react with  $\text{HPO}_4^{2-}$ . Therefore, the predominant phosphorus containing species is  $\text{HPO}_4^{2-}$ .

17. What is an autoionization reaction? Write the chemical equation for the autoionization of ammonia.  $\text{H}_3\text{O}^{1+}$  and  $\text{OH}^{1-}$  are the strongest acid and base that can exist in aqueous solutions because water reacts with any acids or bases that are stronger than these acids. This is known as the leveling effect. What are the strongest acid and base that can exist in liquid ammonia?

Autoionization reactions are reactions between two identical molecules that produce ion. The autoionization of ammonia is  $\text{NH}_3 + \text{NH}_3 \rightarrow \text{NH}_2^{1-} + \text{NH}_4^{1+}$ . The strongest acid that can exist in liquid ammonia is  $\text{NH}_4^{1+}$  and the strongest base is  $\text{NH}_2^{1-}$ .

19. Determine the hydronium and hydroxide ion concentrations in the following solutions:

- a) detergent; pH = 10.3                       $[\text{H}_3\text{O}^{1+}] = 5 \times 10^{-11} \text{ M}$                        $[\text{OH}^-] = 2 \times 10^{-4} \text{ M}$   
b) stomach acid; pH = 2.4                       $[\text{H}_3\text{O}^{1+}] = 4 \times 10^{-3} \text{ M}$                        $[\text{OH}^-] = 3 \times 10^{-12} \text{ M}$   
c) beer; pH = 4.2                               $[\text{H}_3\text{O}^{1+}] = 6 \times 10^{-5} \text{ M}$                        $[\text{OH}^-] = 2 \times 10^{-10} \text{ M}$   
d) milk of magnesia; pH = 10.5                       $[\text{H}_3\text{O}^{1+}] = 3 \times 10^{-11} \text{ M}$                        $[\text{OH}^-] = 3 \times 10^{-4} \text{ M}$

21. Determine the  $\text{pK}_a$  of each of the following acids:

- a) tartaric acid                       $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$                        $K_a = 1.0 \times 10^{-3}$                        $\text{pK}_a = -\log(1.0 \times 10^{-3}) = 3.00$   
b) boric acid                               $\text{H}_3\text{BO}_3$                                $K_a = 5.8 \times 10^{-10}$                        $\text{pK}_a = -\log(5.8 \times 10^{-10}) = 9.24$

# Acids and Bases

23. Determine the  $K_a$  of each of the following acids:

- a) hypobromous acid      HOBr       $pK_a = 8.64$        $K_a = 10^{-8.64} = 2.3 \times 10^{-9}$   
b) saccharin       $H_3NC_7H_4SO_3$        $pK_a = 11.68$        $K_a = 10^{-11.68} = 2.1 \times 10^{-12}$

25. What is the  $pK_b$  of the conjugate base of each of the acids in Exercise 21:

- a) tartaric acid       $pK_a = 3.00$        $pK_b = 14.00 - 3.00 = 11.00$   
b) boric acid       $pK_a = 9.24$        $pK_b = 14.00 - 9.24 = 4.76$

27. What is the  $K_a$  of the acids whose conjugate bases have the following  $pK_b$ 's?

Use the expression  $pK_a = pK_w - pK_b$  and assume a temperature of 25 °C where  $pK_w = 14.00$ .

- a) 8.37       $pK_a = 14.00 - 8.37 = 5.63$        $K_a = 10^{-5.63} = 2.3 \times 10^{-6}$   
b) 12.66       $pK_a = 14.00 - 12.66 = 1.34$        $K_a = 10^{-1.34} = 4.6 \times 10^{-2}$   
c) 0.22       $pK_a = 14.00 - 0.22 = 13.78$        $K_a = 10^{-13.78} = 1.7 \times 10^{-14}$

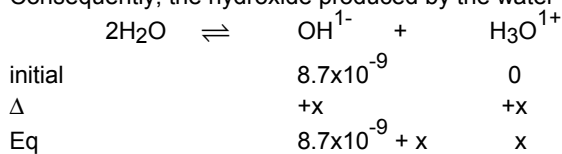
29. Determine the pH and pOH of solutions with the following hydroxide ion concentrations:

Use the expression  $pOH = -\log[OH^{1-}]$  and  $pH = pK_w - pOH$  and assume a temperature of 25 °C where  $pK_w = 14.00$ .

- a)  $7.5 \times 10^{-8}$  M       $pOH = -\log(7.5 \times 10^{-8}) = 7.12$        $pH = 14.00 - 7.12 = 6.88$   
b)  $3.9 \times 10^{-4}$  M       $pOH = -\log(3.9 \times 10^{-4}) = 3.41$        $pH = 14.00 - 3.41 = 10.59$   
c)  $1.0 \times 10^{-12}$  M       $pOH = -\log(1.0 \times 10^{-12}) = 12.00$        $pH = 14.00 - 12.00 = 2.00$

31.\* The solubility of  $Al(OH)_3$  in water is  $2.9 \times 10^{-9}$  M at 25 °C. What is the pH of a saturated solution of  $Al(OH)_3$ ?

$[OH^{1-}] = 3(2.9 \times 10^{-9}) = 8.7 \times 10^{-9}$  M, but a hydroxide must produce an basic solution, so  $[OH^{1-}] > 1.0 \times 10^{-7}$  M. Consequently, the hydroxide produced by the water must exceed that produced by such an insoluble hydroxide.



The  $K_w$  expression is  $K_w = (8.7 \times 10^{-9} + x)(x)$  or  $1.0 \times 10^{-14} = 8.7 \times 10^{-9}x + x^2$ .

Putting the  $K_w$  expression in the form of the quadratic equation,  $1x^2 + 8.7 \times 10^{-9}x - 1.0 \times 10^{-14} = 0$

Solving for x with the quadratic equation,  $x = \frac{-8.7 \times 10^{-9} \pm \sqrt{(8.7 \times 10^{-9})^2 + 4(1)(1.0 \times 10^{-14})}}{2} = 9.6 \times 10^{-8}$  M

Nine times more  $OH^{1-}$  ion is produced by the water than is added.

$[OH^{1-}] = 8.7 \times 10^{-9} + 9.6 \times 10^{-8} = 10.5 \times 10^{-8}$  M, so the pOH is 6.98, and pH = 7.02.

33. What is pH of a neutral solution at 37 °C?

$[H_3O^{1+}] = [OH^{1-}]$  in a neutral solution, and  $K_w = [H_3O^{1+}][OH^{1-}] = 2.42 \times 10^{-14}$

so,  $[H_3O^{1+}] = \sqrt{K_w} = \sqrt{2.42 \times 10^{-14}} = 1.56 \times 10^{-7}$  M

$pH = -\log(1.56 \times 10^{-7}) = 6.81$

35. The  $pK_a$  of a weak acid is 5.21 at 37 °C. What is the  $pK_b$  of its conjugate base?

$pK_w = -\log K_w = -\log(2.42 \times 10^{-14}) = 13.62$

$pK_b = pK_w - pK_a = 13.62 - 5.21 = 8.41$

37. What is the pH of each of the following aqueous solutions?

- a) 0.066 M HCl       $pH = -\log 0.066 = 1.18$   
b) 0.21 M KOH       $pOH = -\log 0.21 = 0.68$        $pH = 14.00 - 0.68 = 13.32$   
c) 0.11 M  $Ba(OH)_2$        $[OH^{1-}] = 2(0.11) = 0.22$  M  
    $pOH = -\log 0.22 = 0.66$        $pH = 14.00 - 0.66 = 13.34$

**39. To what volume must 5.0 mL of 6.0 M HCl be diluted to prepare a solution with pH = 1.22?**

The final concentration of hydronium ion is  $10^{-1.22} = 0.060$  M. Using the dilution expression from Chapter 2, we solve for the final volume,

$$V_f = \frac{M_i V_i}{M_f} = \frac{(6.0 \text{ mmol/mL})(5.0 \text{ mL})}{0.060 \text{ mmol/mL}} = 5.0 \times 10 \text{ mL} = 0.50 \text{ L}$$

**41. What volume of HCl gas measured at 300. K and 1 atm is required to prepare 5.0 L of hydrochloric acid with a pH of 3.84?**

$$[\text{H}_3\text{O}^{1+}] = 10^{-3.84} = 1.5 \times 10^{-4} \text{ M}$$

The number of moles of HCl required is  $(5.0 \text{ L})(1.5 \times 10^{-4} \text{ M}) = 7.2 \times 10^{-4} \text{ mol}$

Use the ideal gas law to determine the volume of gas

$$V = \frac{nRT}{P} = \frac{(7.2 \times 10^{-4} \text{ mol})(0.0821 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(300 \text{ K})}{1.0 \text{ atm}} = 0.018 \text{ L}$$

**43. How many grams of Ba(OH)<sub>2</sub> would have to be dissolved in water to prepare 500.0 mL of a pH= 9.80 solution?**

$$\text{pOH} = 14.00 - 9.80 = 4.20; [\text{OH}^{1-}] = 10^{-4.20} = 6.3 \times 10^{-5} \text{ M}$$

$$0.5000 \text{ L solution} \times \frac{6.3 \times 10^{-5} \text{ mol OH}^{1-}}{1 \text{ L solution}} \times \frac{1 \text{ mol Ba(OH)}_2}{2 \text{ mol OH}^{1-}} \times \frac{171.3 \text{ g Ba(OH)}_2}{1 \text{ mol Ba(OH)}_2} = 2.8 \times 10^{-3} \text{ g} = 2.8 \text{ mg}$$

**45. What is the 5% rule?**

The amount of acid that reacts can be assumed negligible in the subtraction from the original amount of acid if no more than 5% of the acid reacts.

**47. Use the 5% rule to determine whether the equilibrium concentration of the acid can be approximated by its makeup concentration.**

Proceed as in Exercise 41 after converting the pK<sub>a</sub> of the acid into its K<sub>a</sub>.

**a) 0.80 M cyanic acid (HCNO, pK<sub>a</sub> = 3.46)**

$$K_a = 10^{-3.46} = 3.5 \times 10^{-4} \quad \sqrt{\frac{3.5 \times 10^{-4}}{0.80}} \times 100\% = 2.1\% \quad \text{approximation is valid}$$

**b) 4.4 x 10<sup>-3</sup> M hydrazoic acid (HN<sub>3</sub>, pK<sub>a</sub> = 4.6)**

$$K_a = 10^{-4.6} = 3 \times 10^{-5} \quad \sqrt{\frac{3 \times 10^{-5}}{4.4 \times 10^{-3}}} \times 100\% = 8\% \quad \text{approximation is NOT valid}$$

**c) 3.0 M arsenic acid (H<sub>3</sub>AsO<sub>4</sub>, pK<sub>a</sub> = 2.26)**

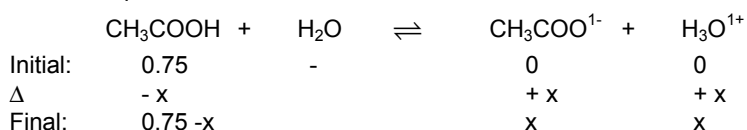
$$K_a = 10^{-2.26} = 5.5 \times 10^{-3} \quad \sqrt{\frac{5.5 \times 10^{-3}}{3.0}} \times 100\% = 4.3\% \quad \text{approximation is valid}$$

**49. What is the pH of vinegar, a 4.5% solution of acetic acid?**

First, determine the concentration of acetic acid in a 4.5% solution.

$$\frac{4.5 \text{ g CH}_3\text{COOH}}{100 \text{ g sol'n}} \times \frac{1 \text{ mol CH}_3\text{COOH}}{60.054 \text{ g CH}_3\text{COOH}} \times \frac{1.0 \text{ g sol'n}}{1.0 \text{ mL sol'n}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.75 \text{ M CH}_3\text{COOH}$$

Next, set up the dissociation reaction:



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = \frac{x^2}{0.75-x} = 1.8 \times 10^{-5}; \text{ from Appendix H;}$$

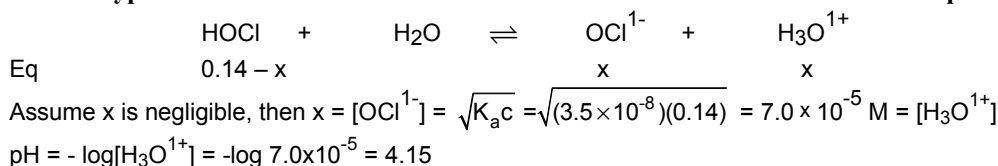
K<sub>a</sub> is small, so we can simplify the above relationship

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{0.75}, \text{ so } x = [\text{H}_3\text{O}^{1+}] = \sqrt{K_a c} = \sqrt{(1.8 \times 10^{-5})(0.75)} = 3.7 \times 10^{-3} \text{ M}$$

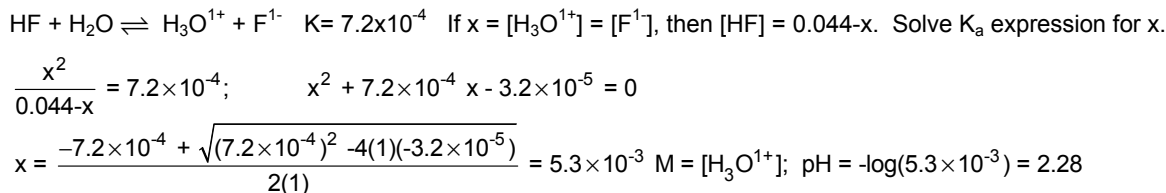
pH = -log[H<sub>3</sub>O<sup>+</sup>] = 2.44 when extra figures are kept from previous calculations

# Acids and Bases

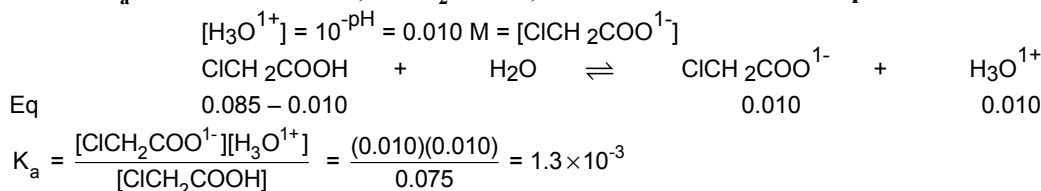
51. What is the hypochlorite ion concentration in a 0.14 M-solution of HOCl? What is the pH of the solution?



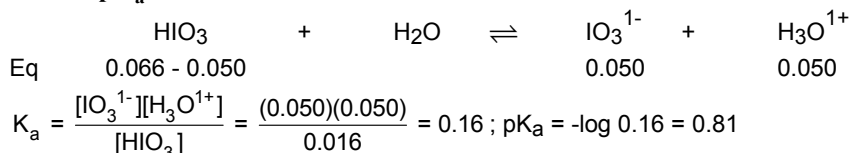
53. What is the pH of 0.044 M HF?



55. What is the  $K_a$  of chloroacetic acid,  $\text{ClCH}_2\text{COOH}$ , if a 0.085-M solution has a pH of 2.00?



57. What is the  $\text{p}K_a$  of iodic acid if the iodate ion concentration in a 0.066-M solution of  $\text{HIO}_3$  is 0.050 M?

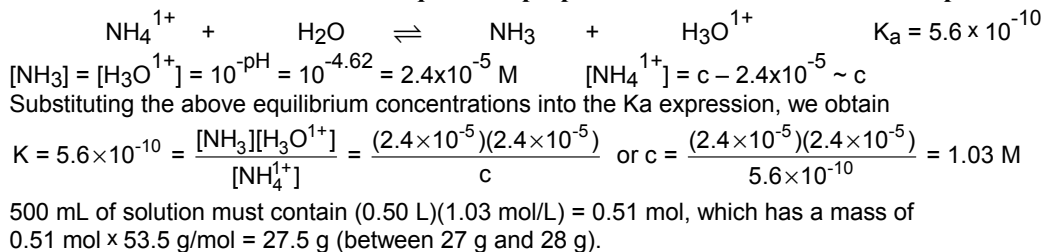


59. What is the percent dissociation of 0.26 M uric acid ( $\text{p}K_a = 3.89$ )?

Use Equation 6.7, but realize that the equation is only good if less than 5% dissociates.

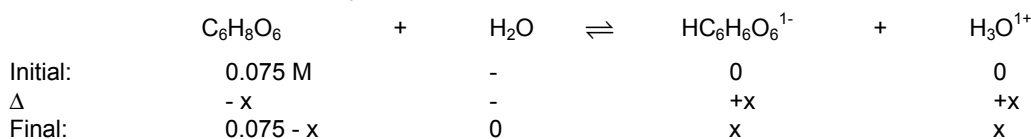
$$K_a = 10^{-\text{p}K_a} = 10^{-3.89} = 1.3 \times 10^{-4} \quad \sqrt{\frac{K_a}{c_0}} \times 100\% = \sqrt{\frac{1.3 \times 10^{-4}}{0.26}} \times 100\% = 2.2\%$$

61. What mass of ammonium chloride is required to prepare 500. mL of a solution with a pH of 4.62?



63. Vitamin C is ascorbic acid,  $\text{H}_2\text{C}_6\text{H}_8\text{O}_6$ . Calculate the pH,  $[\text{H}_2\text{C}_6\text{H}_8\text{O}_6]$ ,  $[\text{HC}_6\text{H}_6\text{O}_6^{1-}]$  and  $[\text{C}_6\text{H}_6\text{O}_6^{2-}]$  in a 0.075-M solution of ascorbic acid.

From Appendix C,  $K_1 = 8.0 \times 10^{-5}$  and  $K_2 = 1.6 \times 10^{-12}$ .  $K_2$  is much smaller, so concentrations achieved in the first deprotonation are not affected by the second. The reaction table for the first ionization is

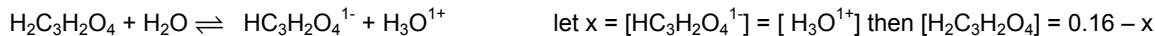


Assuming the extent of reaction is negligible,

$$[\text{H}_3\text{O}^{1+}] = [\text{HC}_6\text{H}_6\text{O}_6^{1-}] = \sqrt{K_a c_0} = \sqrt{(8.0 \times 10^{-5})(0.075)} = 0.0024 \text{ M}, \text{ which is less than 5\% of } 0.075 \text{ M, so the approximation is acceptable. } [\text{C}_6\text{H}_8\text{O}_6] = 0.075 - 0.002 = 0.073 \text{ M. } \text{pH} = -\log(0.0022) = 2.66$$

The concentration of the ion produced in the second ionization equals  $K_2$ , so  $[\text{C}_6\text{H}_6\text{O}_6^{2-}] = 1.6 \times 10^{-12} \text{ M}$ .

65. What are concentrations of all species in a 0.16-M solution of malonic acid ( $\text{H}_2\text{C}_3\text{H}_2\text{O}_4$ )?  $K_1 = 1.5 \times 10^{-3}$  and  $K_2 = 2.0 \times 10^{-6}$



$$\frac{x^2}{0.16-x} = 0.0015; \quad x^2 + 0.0015x - 2.4 \times 10^{-4} = 0$$

$$x = \frac{-0.0015 + \sqrt{(0.0015)^2 - 4(1)(-2.4 \times 10^{-4})}}{2(1)} = 0.015 \text{ M} = [\text{H}_3\text{O}^{1+}] = [\text{HC}_3\text{H}_2\text{O}_4^{1-}]$$

$$[\text{H}_2\text{C}_3\text{H}_2\text{O}_4] = 0.16 - 0.015 = 0.15 \text{ M}; \quad [\text{C}_3\text{H}_2\text{O}_4^{2-}] = K_2 = 2.0 \times 10^{-6} \text{ M}$$

The concentration of  $\text{C}_3\text{H}_2\text{O}_4^{2-}$  produced in the second step is negligible compared to 0.015 M, so the assumption that  $[\text{C}_3\text{H}_2\text{O}_4^{2-}] = K_2$  is valid.

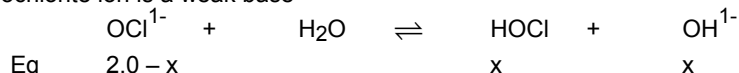
67. The hypochlorite ion is the active ingredient in bleach.

- a) Industrial bleach is 15% NaOCl by mass. What is the hypochlorous acid concentration in industrial bleach?

First, convert the mass percent to a molarity

$$\frac{15 \text{ g NaOCl}}{100 \text{ g sol'n}} = \frac{15 \text{ g NaOCl} \times \frac{1 \text{ mol NaOCl}}{74.5 \text{ g NaOCl}}}{100 \text{ g sol'n} \times \frac{1 \text{ mL sol'n}}{1 \text{ g sol'n}} \times \frac{1 \text{ L sol'n}}{1000 \text{ mL sol'n}}} = \frac{0.20 \text{ mol NaOCl}}{0.10 \text{ L sol'n}} = 2.0 \text{ M}$$

The hypochlorite ion is a weak base



Assume that  $[\text{OCl}^{1-}] = 2.0 \text{ M}$  at equilibrium because  $c_0$  is very large and  $\text{OCl}^{1-}$  is a weak base.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-8}} = 2.9 \times 10^{-7} = \frac{[\text{HOCl}][\text{OH}^{1-}]}{[\text{OCl}]} = \frac{x^2}{2.0}$$

$$x = \sqrt{(2.9 \times 10^{-7})(2.0)} = 7.6 \times 10^{-4} \text{ M} = [\text{HOCl}], \text{ which is negligible compared to } 2.0 \text{ M}.$$

- b) Household bleach is a 5.25% solution of NaOCl. What is its pH?

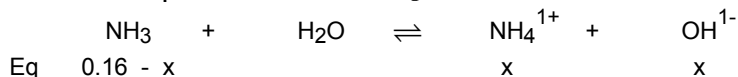
Determine the molar concentration of the NaOCl. Assuming a 100. g solution (also 100. mL solution since the density is 1.0 g/mL), 5.25% translates to 5.25 g.

$$\frac{5.25 \text{ g NaOCl}}{0.100 \text{ L}} \times \frac{1 \text{ mol}}{74.44 \text{ g}} = 0.705 \text{ M NaOCl}$$

$K_b = 2.9 \times 10^{-7}$  from part a, so the hydroxide ion concentration is

$$[\text{OH}^{1-}] = \sqrt{K_b \times c_0} = \sqrt{(2.9 \times 10^{-7})(0.70)} = 4.5 \times 10^{-4} \text{ M}, \text{ so } \text{pOH} = -\log [\text{OH}^{1-}] = 3.35 \text{ and } \text{pH} = 14.00 - 3.35 = 10.65$$

69. What are the  $\text{NH}_4^{1+}$  concentration and pH of a 0.16-M solution of  $\text{NH}_3$ ?



$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}$$

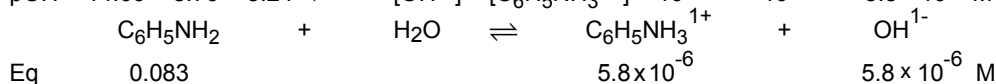
Assume that  $x$  is negligible in the subtraction, so Equation 6.7 can be used.

$$[\text{NH}_4^{1+}] = [\text{OH}^{1-}] = \sqrt{K_b c_0} = \sqrt{(1.8 \times 10^{-5})(0.16)} = 1.7 \times 10^{-3} \text{ M} \text{ (which is less than 5\% reaction)}$$

$$\text{pOH} = -\log 1.7 \times 10^{-3} = 2.77, \text{ so } \text{pH} = 14.00 - 2.77 = 11.23$$

71. The pH of 0.083 M aniline ( $\text{C}_6\text{H}_5\text{NH}_2$ ) is 8.76. What are its  $K_b$  and  $\text{p}K_b$ ?

$$\text{pOH} = 14.00 - 8.76 = 5.24 \rightarrow [\text{OH}^{1-}] = [\text{C}_6\text{H}_5\text{NH}_3^{1+}] = 10^{-\text{pOH}} = 10^{-5.24} = 5.8 \times 10^{-6} \text{ M}$$



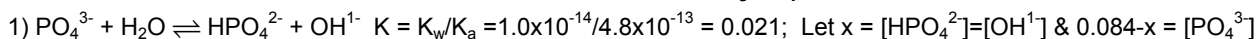
$$K_b = \frac{[\text{C}_6\text{H}_5\text{NH}_3^{1+}][\text{OH}^{1-}]}{[\text{C}_6\text{H}_5\text{NH}_2]} = \frac{(5.8 \times 10^{-6})^2}{0.083} = 4.1 \times 10^{-10} \rightarrow \text{p}K_b = -\log 4.1 \times 10^{-10} = 9.39$$

# Acids and Bases

73. What percent of the ammonia molecules in a 0.12-M solution have reacted to produce ammonium ions?

Modify Equation 6.6 for bases.  $\% = \sqrt{\frac{K_b}{c_o}} \times 100\% = \sqrt{\frac{1.8 \times 10^{-5}}{0.12}} \times 100\% = 1.2\%$  (less than 5% so assumption valid)

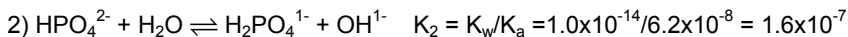
75. What are the concentrations of all P containing species in 0.084 M  $K_3PO_4$ ? What is the pH of the solution?



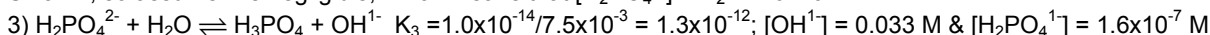
$$\frac{x^2}{0.084 - x} = 0.021; \quad x^2 + 0.021x - 1.75 \times 10^{-3} = 0$$

$$x = \frac{-0.021 + \sqrt{(0.021)^2 - 4(1)(-1.75 \times 10^{-3})}}{2(1)} = 0.033 \text{ M} = [OH^{1-}] = [HPO_4^{2-}]$$

$$[PO_4^{3-}] = 0.084 - 0.033 = 0.051 \text{ M}$$



Small K, so assume x is negligible, which means that  $[H_2PO_4^{1-}] = K_2 = 1.6 \times 10^{-7} \text{ M}$



$$1.3 \times 10^{-12} = \frac{[H_3PO_4](0.033)}{1.6 \times 10^{-7}}; \quad [H_3PO_4] = \frac{(1.3 \times 10^{-12})(1.6 \times 10^{-7})}{0.033} = 6.5 \times 10^{-18} \text{ M}$$

$$pOH = -\log(0.033) = 1.48; \quad pH = 14.00 - 1.48 = 12.52$$

77. Define the term salt. Given an example of a neutral, a basic, and an acidic salt.

A salt is an ionic compound formed in an acid-base reaction. Neutral:  $KClO_4$ ; Basic: KCN; Acidic:  $NH_4NO_3$ .

79. Indicate whether each of the following is an acidic, a basic, or a neutral salt.

- a)  $K_2SO_4$  neutral                      b)  $K_2SO_3$  basic                      c)  $KHSO_4$  acidic  
d)  $K_2CO_3$  basic                      e)  $KOCl$  basic                      f)  $NH_4OCl$  basic ( $K_b > K_a$ )

81. What is the pH of each of the following salt solutions?

a) 0.12 M  $NH_4Cl$

$NH_4^{1+}$  is acidic, while  $Cl^{1-}$  is neutral, so  $NH_4Cl$  is an acidic salt:  $[H_3O^{1+}] = \sqrt{(5.6 \times 10^{-10})(0.12)} = 8.2 \times 10^{-6} \text{ M}$ ;  $pH = 5.09$

b) 0.096 M KCN

$K^{1+}$  is neutral, but  $CN^{1-}$  is basic, so KCN is a basic salt:  $K_b = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-10}} = 2.5 \times 10^{-5}$

$$[OH^{1-}] = \sqrt{(2.5 \times 10^{-5})(0.096)} = 1.55 \times 10^{-3} \text{ M}; \quad pOH = 2.81; \quad pH = 11.19$$

c) 0.10 M  $KHSO_3$

$K^{1+}$  is neutral, and  $HSO_3^{1-}$  is amphiprotic, so use Equation 6.12.  $pH = \frac{1}{2}(pK_1 + pK_2) = \frac{1}{2}(1.82 + 7.00) = 4.41$