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) OH⁻ → O²⁻
   b) H₂O₂ → HO₂⁻
   c) H₂PO₄⁻ → HPO₄²⁻
   d) H₃O⁺ → H₂O
   e) H₂SO₃ → HSO₃⁻

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) NaOH yes
   b) NaCl not in water
   c) CH₃OH no
   d) KCN yes
   e) KH₂PO₄ yes

9. Explain how the reaction Ag⁺⁺ + Cl⁻ → AgCl is a Lewis acid-base reaction. Is it a Brønsted acid-base reaction?
   Ag⁺⁺ is a Lewis acid that accepts a lone pair on Cl⁻ 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 << 1.
    a) Aqueous sodium sulfate is added to hydrobromic acid.
       SO₄²⁻ + H₃O⁺ → HSO₄⁻ + H₂O (weak base + strong acid)
    b) Aqueous NH₄Cl and aqueous KF are mixed.
       NH₄⁺ + F⁻ → 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.
       CN⁻ + H₂SO₃ → HSO₃⁻ + HCN
    d) Acetic acid and aqueous sodium hypochlorite are mixed.
       CH₃COOH + OCl⁻ → CH₃COO⁻ + HOCl (produced acid is weaker than reacting acid)
    e) Hydrogen sulfide is bubbled into water.
       H₂S + H₂O → no reaction (this is the Kₐ reaction of a weak acid and the extent is very small)

13. Explain why HCl is a strong acid, but 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?
    NH₃ is strong enough to remove a proton from H₃PO₄ and H₂PO₄⁻, but it is too weak to react with HPO₄²⁻. Therefore,
    the predominant phosphorus containing species is HPO₄²⁻.

17. What is an autoionization reaction? Write the chemical equation for the autoionization of ammonia. H₃O⁺ is 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 NH₃ + NH₃ → NH₂⁻ + NH₄⁺. The strongest acid that can exist in liquid ammonia is NH₄⁺ and the
    strongest base is NH₂⁻.

19. Determine the hydronium and hydroxide ion concentrations in the following solutions:
    a) detergent; pH = 10.3
       [H₃O⁺] = 5 x 10⁻¹¹ M  [OH⁻] = 2 x 10⁻⁴ M
    b) stomach acid; pH = 2.4
       [H₃O⁺] = 4 x 10⁻³ M  [OH⁻] = 3 x 10⁻¹² M
    c) beer; pH = 4.2
       [H₃O⁺] =6 x 10⁻⁵ M  [OH⁻] = 2 x 10⁻¹⁰ M
    d) milk of magnesia; pH = 10.5
       [H₃O⁺] =3 x 10⁻¹¹ M  [OH⁻] = 3 x 10⁻⁴ M

21. Determine the pKₐ of each of the following acids:
    a) tartaric acid
       H₂C₄H₄O₆  Kₐ = 1.0 x 10⁻³  pKₐ = -log (1.0 x 10⁻³) = 3.00
    b) boric acid
       H₃BO₃  Kₐ = 5.8 x 10⁻¹⁰  pKₐ = -log (5.8 x 10⁻¹⁰) = 9.24
23. Determine the $K_a$ of each of the following acids:
   a) hypobromous acid $\text{HOBr}$  $pK_a = 8.64$  $K_a = 10^{-8.64} = 2.3 \times 10^{-9}$
   b) saccharin $\text{H}_3\text{NC}_4\text{H}_4\text{SO}_3$  $pK_a = 11.68$  $K_a = 10^{-11.68} = 2.1 \times 10^{-12}$

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 $pOH = -\log[\text{OH}^-]$ 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 $\text{Al(OH)}_3$ in water is $2.9 \times 10^{-9}$ M at 25 °C. What is the pH of a saturated solution of $\text{Al(OH)}_3$?
   $[\text{OH}^-] = 3 \times (2.9 \times 10^{-9}) = 8.7 \times 10^{-9}$ M, but a hydroxide must produce an basic solution, so $[\text{OH}^-] > 1.0 \times 10^{-7}$ M. Consequently, the hydroxide produced by the water must exceed that produced by such an insoluble hydroxide.
   $$2\text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{H}_3\text{O}^+$$
   initial  $8.7 \times 10^{-9}$  $0$
   $\Delta$  $+x$  $+x$
   Eq  $8.7 \times 10^{-9} + x$  $x$

   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, $x^2 + 8.7 \times 10^{-9}x - 1.0 \times 10^{-14} = 0$

   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 $\text{OH}^-$ ion is produced by the water than is added.

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

33. What is pH of a neutral solution at 37 °C?
   $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ in a neutral solution, and $K_w = [\text{H}_3\text{O}^+] [\text{OH}^-] = 2.42 \times 10^{-14}$

   so, $[\text{H}_3\text{O}^+] = \sqrt{K_w} = \sqrt{2.42 \times 10^{-14}} = 1.56 \times 10^{-7}$ 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 $\text{HCl}$  $pH = -\log 0.066 = 1.18$
   b) $0.21$ M $\text{KOH}$  $pOH = -\log 0.21 = 0.68$  $pH = 14.00 - 0.68 = 13.32$
   c) $0.11$ M $\text{Ba(OH)}_2$  $[\text{OH}^-] = 2(0.11) = 0.22$ 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,

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

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?

$$[H_3O^{+}] = 10^{-3.84} = 1.5 \times 10^{-4} \text{ M}$$

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)$_2$ would have to be dissolved in water to prepare 500.0 mL of a pH = 9.80 solution?

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

$0.5000 \text{ L solution} \times \frac{6.5 \times 10^{-5} \text{ mol OH}^-}{1 \text{ L solution}} \times \frac{1 \text{ mol Ba(OH)}_2}{2 \text{ mol OH}^-} \times \frac{171.3 \text{ g Ba(OH)}_2}{1 \text{ mol Ba(OH)}_2} = 2.8 \times 10^{-3} \text{ g} = 2.8 \text{ mg}$

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 p$K_a$ of the acid into its $K_a$.

a) 0.80 M cyanic acid (HCNO, $pK_a = 3.46$)

$$K_a = 10^{-3.46} = 3.5 \times 10^{-4}$$

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

b) $4.4 \times 10^{-3}$ M hydrazoic acid (HN$_3$, $pK_a = 4.6$)

$$K_a = 10^{-4.6} = 3 \times 10^{-5}$$

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

c) 3.0 M arsenic acid (H$_3$AsO$_4$, $pK_a = 2.26$)

$$K_a = 10^{-2.26} = 5.5 \times 10^{-3}$$

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

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:

$$\text{CH}_3\text{COOH} + \text H_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text H_3\text{O}^+$$

Initial: $0.75 - 0 0$

$\Delta: -x +x +x$

Final: $0.75 -x x x$

$$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}$; from Appendix H;

$K_a$ is small, so we can simplify the above relationship

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

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = 2.44 \text{ when extra figures are kept from previous calculations}$$
51. What is the hypochlorite ion concentration in a 0.14 M-solution of HOCl? What is the pH of the solution?

\[
\text{HOCl} + \text{H}_2\text{O} \rightleftharpoons \text{OCl}^- + \text{H}_3\text{O}^+
\]

\[
\begin{align*}
\text{Eq} & \quad 0.14 - x = x & x = 0.004 \\
\text{Assume x is negligible, then } x & = [\text{OCl}^-] = \sqrt{K_{ac}c} = \sqrt{(3.5 \times 10^{-8})(0.14)} = 7.0 \times 10^{-5} \text{ M} = [\text{H}_3\text{O}^+] \\
\text{pH} & = -\log[\text{H}_3\text{O}^+] = -\log 7.0 \times 10^{-5} = 4.15
\end{align*}
\]

53. What is the pH of 0.044 M HF?

\[
\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-; K = 7.2 \times 10^{-4} \quad \text{If } x = [\text{H}_3\text{O}^+] = [\text{F}^-], \text{ then } [\text{HF}] = 0.044 - x. \text{ Solve } Ka \text{ expression for } x.
\]

\[
\begin{align*}
\frac{x^2}{0.044 - x} = 7.2 \times 10^{-4}; & \quad x^2 + 7.2 \times 10^{-4} x - 3.2 \times 10^{-5} = 0 \\
x = & \frac{-7.2 \times 10^{-4} + \sqrt{(7.2 \times 10^{-4})^2 - 4(1)(-3.2 \times 10^{-5})}}{2(1)} = 5.3 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+]; \quad \text{pH} = -\log(5.3 \times 10^{-3}) = 2.28
\end{align*}
\]

55. What is the Ka of chloroacetic acid, ClCH₂COOH, if a 0.085-M solution has a pH of 2.00?

\[
\text{ClCH}_2\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{ClCH}_2\text{COO}^- + \text{H}_3\text{O}^+
\]

\[
\begin{align*}
\text{Eq} & \quad 0.085 - 0.010 \quad 0.010 \quad 0.010 \\
Ka & = \frac{[\text{ClCH}_2\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{ClCH}_2\text{COOH}]} = \frac{(0.010)(0.010)}{0.075} = 1.3 \times 10^{-3}
\end{align*}
\]

57. What is the pKa of iodic acid if the iodate ion concentration in a 0.066-M solution of HIO₃ is 0.050 M?

\[
\text{HIO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{IO}_3^- + \text{H}_3\text{O}^+
\]

\[
\begin{align*}
\text{Eq} & \quad 0.066 - 0.050 \quad 0.050 \quad 0.050 \\
K_a & = \frac{[\text{IO}_3^-][\text{H}_3\text{O}^+]}{[\text{HIO}_3]} = \frac{(0.050)(0.050)}{0.016} = 0.16; \quad \text{pKa} = -\log 0.16 = 0.81
\end{align*}
\]

59. What is the percent dissociation of 0.26 M uric acid (pKₐ = 3.89)?

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

\[
\text{K}_a = 10^{-pK_a} = 10^{-3.89} = 1.3 \times 10^{-4}
\]

\[
\frac{\text{K}_a \times 100\% = \sqrt{\frac{1.3 \times 10^{-4}}{0.26} \times 100\% = 2.2\%}}{c_o}
\]

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

\[
\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+; \quad K_a = 5.6 \times 10^{-10}
\]

\[
\begin{align*}
[\text{NH}_3] = [\text{H}_3\text{O}^+] & = 10^{-pH} = 10^{-4.62} = 2.4 \times 10^{-5} \text{ M} \quad [\text{NH}_4^+] = c - 2.4 \times 10^{-5} \sim c \\
K & = 5.6 \times 10^{-10} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]^2} \quad \text{or } c = \frac{(2.4 \times 10^{-5})(2.4 \times 10^{-5})}{5.6 \times 10^{-10}} = 1.03 \text{ M}
\end{align*}
\]

500 mL of solution must contain (0.50 L)(1.03 mol/L) = 0.51 mol, which has a mass of 0.51 mol x 53.5 g/mol = 27.5 g (between 27 g and 28 g).

63. Vitamin C is ascorbic acid, H₂C₆H₈O₆. Calculate the pH, [H₂C₆H₈O₆], [HC₆H₈O₆⁻] and [C₆H₈O₆²⁻] in a 0.075-M solution of ascorbic acid.

From Appendix C, K₁ = 8.0x10⁻⁵ and K₂ = 1.6x10⁻¹². K₂ is much smaller, so concentrations achieved in the first deprotonation are not affected by the second. The reaction table for the first ionization is

\[
\begin{align*}
\text{C}_6\text{H}_8\text{O}_6^- & \quad + \quad \text{H}_2\text{O} \quad \rightleftharpoons \quad \text{HC}_6\text{H}_8\text{O}_6^- \quad + \quad \text{H}_3\text{O}^+ \\
\text{Initial:} & \quad 0.075 \text{ M} & - & 0 & 0 \\
\Delta & \quad -x & - & +x & +x \\
\text{Final:} & \quad 0.075 - x & 0 & x & x
\end{align*}
\]

Assuming the extent of reaction is negligible,

\[
[\text{H}_3\text{O}^+] = [\text{HC}_6\text{H}_8\text{O}_6^-] = \sqrt{K_{a1}c_o} = \sqrt{(8.0 \times 10^{-5})(0.075)} = 0.0024 \text{ M}, \text{ which is less than 5\% of } 0.075 \text{ M, so the approximation is acceptable.} \quad [\text{C}_6\text{H}_8\text{O}_6^-] = 0.075 - 0.002 = 0.073 \text{ M.} \quad \text{pH} = -\log (0.0022) = 2.66
\]

The concentration of the ion produced in the second ionization equals K₂, so [C₆H₈O₆²⁻] = 1.6 x 10⁻¹² M.
65. What are concentrations of all species in a 0.16-M solution of malonic acid (H_{2}C_{3}H_{2}O_{4})? \( K_1 = 1.5 \times 10^{-3} \) and \( K_2 = 2.0 \times 10^{-6} \)

\[
H_{2}C_{3}H_{2}O_{4} + H_2O \rightleftharpoons H_2C_{3}H_{2}O_{4}^{-} + H_{3}O^{+}
\]

Let \( x = [H_2C_{3}H_{2}O_{4}^{-}] = [H_{3}O^{+}] \) then \( [H_{2}C_{3}H_{2}O_{4}] = 0.16 - x \)

\[
x^2 = 0.0015; \quad (0.0015)^{2} - 4(1)(-2.4 \times 10^{-4}) = 0.015 M = [H_{3}O^{+}] = [H_2C_{3}H_{2}O_{4}^{-}]
\]

\[
[H_{2}C_{3}H_{2}O_{4}] = 0.16 - 0.015 = 0.15 M; \quad [C_{3}H_{2}O_{4}^{2-}] = K_2 = 2.0 \times 10^{-6} M
\]

The concentration of \( C_{3}H_{2}O_{4}^{2-} \) produced in the second step is negligible compared to 0.015 M, so the assumption that \( [C_{3}H_{2}O_{4}^{2-}] = K_2 \) 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

\[
\frac{15 \text{ g NaOCl}}{100 \text{ g sol'n}} \times \frac{1 \text{ mol NaOCl}}{74.44 \text{ g NaOCl}} \times \frac{1 \text{ L sol'n}}{1000 \text{ mL sol'n}} = 0.20 \text{ M NaOCl}
\]

The hypochlorite ion is a weak base

\[
\text{OCl}^{-} + H_2O \rightleftharpoons HOCl + OH^{-}
\]

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

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

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

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.

\[
\frac{5.25 \text{ g NaOCl}}{100 \text{ L}} \times \frac{1 \text{ mol NaOCl}}{74.44 \text{ g}} = 0.0705 \text{ M NaOCl}
\]

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

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

69. What are the NH_4^+ concentration and pH of a 0.16-M solution of NH_3?

\[
\text{NH}_3 + H_2O \rightleftharpoons \text{NH}_4^+ + \text{OH}^{-}
\]

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

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

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

71. The pH of 0.083 M aniline (C_6H_5NH_2) is 8.76. What are its \( K_b \) and \( pK_b \)?

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

\[
\text{C}_6\text{H}_5\text{NH}_2 + H_2O \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^{+} + \text{OH}^{-}
\]

\[
K_b = \frac{[\text{C}_6\text{H}_5\text{NH}_3^{+}] [\text{OH}^{-}]}{[\text{C}_6\text{H}_5\text{NH}_2]} = \frac{(5.8 \times 10^{-6})^2}{0.083} = 4.1 \times 10^{-10} \rightarrow pK_b = -\log 4.1 \times 10^{-5} = 9.39
\]
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. \[ \% = \frac{K_b}{c_o} \times 100\% = \frac{1.8 \times 10^{-5}}{0.12} \times 100\% = 1.2\% \text{ (less than 5% so assumption valid)} \]

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

1) \( \text{PO}_4^{3-} + \text{H}_2\text{O} \rightleftharpoons \text{HPO}_4^{2-} + \text{OH}^- \quad K = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.8 \times 10^{-13}} = 0.021 \); Let \( x = [\text{HPO}_4^{2-}] = [\text{OH}^-] \) & \( 0.084 - x = [\text{PO}_4^{3-}] \)

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

\[ x = \frac{-0.021 + \sqrt{(0.021)^2 - 4(1)(-1.75 \times 10^{-3})}}{2(1)} = 0.033 \text{ M} = [\text{OH}^-] = [\text{HPO}_4^{1-}] \]

\[ [\text{PO}_4^{3-}] = 0.084 - 0.033 = 0.051 \text{ M} \]

2) \( \text{HPO}_4^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{PO}_4^{1-} + \text{OH}^- \quad K_2 = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7} \)

Small \( K \), so assume \( x \) is negligible, which means that \( [\text{H}_2\text{PO}_4^{1-}] = K_2 = 1.6 \times 10^{-7} \text{ M} \)

3) \( \text{H}_2\text{PO}_4^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{PO}_4 + \text{OH}^- \quad K_3 = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{7.5 \times 10^{-3}} = 1.3 \times 10^{-12} \); \[ [\text{OH}^-] = 0.033 \text{ M} \] & \[ [\text{H}_2\text{PO}_4^{1-}] = 1.6 \times 10^{-7} \text{ M} \]

\[ 1.3 \times 10^{-12} = \frac{[\text{H}_3\text{PO}_4](0.033)}{1.6 \times 10^{-7}}; \quad [\text{H}_3\text{PO}_4] = \frac{(1.3 \times 10^{-12})(1.6 \times 10^{-7})}{0.033} = 6.5 \times 10^{-18} \text{ M} \]

\[ \text{pOH} = -\log(0.033) = 1.48; \quad \text{pH} = 14.00 - 1.48 = 12.52 \]

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

\( \text{NH}_4^+ \) is acidic, while \( \text{Cl}^- \) is neutral, so \( \text{NH}_4\text{Cl} \) is an acidic salt: \[ [\text{H}_3\text{O}^+] = \sqrt{(5.6 \times 10^{-10})(0.12)} = 8.2 \times 10^{-6} \text{ M}; \text{pH} = 5.09 \]

 b) 0.096 M KCN

\( K^+ \) is neutral, but \( \text{CN}^- \) is basic, so \( \text{KCN} \) is a basic salt: \[ K_v = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-10}} = 2.5 \times 10^{-5} \]

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

 c) 0.10 M KHSO_3

\( K^+ \) is neutral, and \( \text{HSO}_3^- \) is amphiprotic, so use Equation 6.12. \[ \text{pH} = \frac{1}{2}(pK_1 + pK_2) = \frac{1}{2}(1.82 + 7.00) = 4.41 \]