Chapter 13 Review Questions and Text Homework Solutions

Review Questions

2. a. The K_a reaction always refers to an acid reacting with water to produce the conjugate base of the acid and the hydronium ion (H₃O⁺). For a general weak acid HA, the K_a reaction is:

$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$$
 where $A^- = \text{conjugate base of the acid HA}$

This reaction is often abbreviated as: $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

b. The K_a equilibrium constant is the equilibrium constant for the K_a reaction of some substance. For the general K_a reaction, the K_a expression is:

$$K_{a} = \frac{[A^{-}][H_{3}O^{+}]}{[HA]}$$
 or $K_{a} = \frac{[H^{+}][A^{-}]}{[HA]}$ (for the abbreviated K_{a} reaction)

c. The K_b reaction always refers to a base reacting with water to produce the conjugate acid of the base and the hydroxide ion (OH⁻). For a general base, B, the K_b reaction is:

 $B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$ where $BH^+ = \text{conjugate acid of the base } B$

d. The K_b equilibrium constant for the general K_b reaction is: $K_b = \frac{[BH^+][OH^-]}{[B]}$

- e. A conjugate acid-base pair consists of two substances related to each other by the donating and accepting of a single proton (H⁺). The conjugate bases of the acids HCl, HNO₂, HC₂H₃O₂, and H₂SO₄ are Cl⁻, NO₂⁻, C₂H₃O₂⁻, and HSO₄⁻, respectively. The conjugate acids of the bases NH₃, C₅H₅N, and HONH₂ are NH₄⁺, C₅H₅NH⁺, and HONH₃⁺, respectively. Conjugate acid-base pairs only differ by H⁺ in their respective formulas.
- 3. a. Amphoteric: A substance that can behave either as an acid or as a base.
 - b. The K_w reaction is also called the autoionization of water reaction. The reaction always occurs when water is present as the solvent. The reaction is:

 $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \text{ or } H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

c. The K_w equilibrium constant is also called the ion-product constant or the dissociation constant of water. It is the equilibrium constant for the autoionization reaction of water:

 $K_{\rm w} = [{\rm H}_3{\rm O}^+][{\rm O}{\rm H}^-]$ or $K_{\rm w} = [{\rm H}^+][{\rm O}{\rm H}^-]$

At typical solution temperatures of 25°C, $K_w = 1.0 \times 10^{-14}$.

d. pH is a mathematical term which is equal to the $-\log$ of the H⁺ concentration of a solution $(pH = -\log[H^+])$.

- e. pOH is a mathematical term which is equal to the $-\log$ of the OH⁻ concentration of a solution (pOH = $-\log[OH^{-}]$).
- f. The p of any quantity is the -log of that quantity. So: $pK_w = -\log K_w$. At 25°C, $pK_w = -\log(1.0 \times 10^{-14}) = 14.00$.

Neutral solution at 25°C: $K_w = 1.0 \times 10^{-14} = [H^+][OH^-]$ and pH + pOH = 14.00

$$[H^+] = [OH^-] = 1.0 \times 10^{-7} M; \text{ pH} = pOH = -\log(1.0 \times 10^{-7}) = 7.00$$

Acidic solution at 25°C:

$$[H^+] > [OH^-]; [H^+] > 1.0 \times 10^{-7} M; [OH^-] < 1.0 \times 10^{-7} M; pH < 7.00; pOH > 7.00$$

Basic solution at 25°C:

 $[OH^{-}] > [H^{+}]; [OH^{-}] > 1.0 \times 10^{-7} M; [H^{+}] < 1.0 \times 10^{-7} M; pOH < 7.00; pH > 7.00$

As a solution becomes more acidic, $[H^+]$ increases, so $[OH^-]$ decreases, pH decreases, and pOH increases. As a solution becomes more basic, $[OH^-]$ increases, so $[H^+]$ decreases, pH increases, and pOH decreases.

4. The K_a value refers to the reaction of an acid reacting with water to produce the conjugate base and H_3O^+ . The stronger the acid, the more conjugate base and H_3O^+ produced, and the larger the K_a value.

Strong acids are basically 100% dissociated in water. Therefore, the strong acids have a $K_a >> 1$ because the equilibrium position lies far to the right. The conjugate bases of strong acids are terrible bases; much worse than water, so we can ignore their basic properties in water.

Weak acids are only partially dissociated in water. We say that the equilibrium lies far to the left, thus giving values for $K_a < 1$ (weak acids have mostly reactants at equilibrium and few products present). The conjugate bases of weak acids are better bases than water. When we have a solution composed of just the conjugate base of a weak acid in water, the resulting pH is indeed basic (pH > 7.0). In general, as the acid strength increases, the conjugate base strength decreases, or as acid strength decreases, the conjugate base strength increases. They are inversely related.

Base strength is directly related to the K_b value. The larger the K_b value, the more OH⁻ produced from the K_b reaction, and the more basic the solution (the higher the pH). Weak bases have a $K_b < 1$ and their conjugate acids behave as weak acids in solution. As the strength of the base increases, the strength of the conjugate acid gets weaker; the stronger the base, the weaker the conjugate acid, or the weaker the stronger the conjugate acid.

5. Strong acids are assumed 100% dissociated in water, and we assume that the amount of H⁺ donated by water is negligible. Hence, the equilibrium [H⁺] of a strong acid is generally equal to the initial acid concentration ([HA]₀). Note that solutions of H₂SO₄ can be different from this because H₂SO₄ is a diprotic acid. Also, when you have very dilute solutions of a strong acid, the H⁺ contribution from water by itself must be considered. The strong acids to memorize are HCl, HBr, HI, HNO₃, HClO₄, and H₂SO₄.

 $K_{\rm a}$ values for weak acids are listed in Table 13.2 and in Appendix 5 of the text. Because weak acids only partially dissociate in water, we must solve an equilibrium problem to determine how much H⁺ is added to water by the weak acid. We write down the $K_{\rm a}$ reaction, set-up the ICE table, then solve the equilibrium problem. The two assumptions generally made are that acids are less than 5% dissociated in water and that the H⁺ contribution from water is negligible.

The 5% rule comes from the assumptions that weak acids are less than 5% dissociated. When this is true, the mathematics of the problem are made much easier. The equilibrium expression we get for weak acids in water generally has the form (assuming an initial acid concentration of 0.10 M):

$$K_{\rm a} = \frac{x^2}{0.10 - x} \approx \frac{x^2}{0.10}$$

The 5% rule refers to assuming $0.10 - x \approx 0.10$. The assumption is valid if x is less than 5% of the number the assumption was made against ([HA]₀). When the 5% rule is valid, solving for x is very straight forward. When the 5% rule fails, we must solve the mathematical expression exactly using the quadratic equation (or your graphing calculator). Even if you do have a graphing calculator, reference Appendix 1 to review the quadratic equation. Appendix 1 also discusses the method of successive approximations which can also be used to solve quadratic (and cubic) equations.

6. Strong bases are soluble ionic compounds containing the OH⁻ anion. Strong bases increase the OH⁻ concentration in water by just dissolving. Thus, for strong bases like LiOH, NaOH, KOH, RbOH, and CsOH, the initial concentration of the strong base equals the equilibrium [OH⁻] of water.

The other strong bases to memorize have +2 charged metal cations. The soluble ones to know are Ca(OH)₂, Sr(OH)₂, and Ba(OH)₂. These are slightly more difficult to solve because they donate 2 moles OH⁻ for every mole of salt dissolved. Here, the [OH⁻] is equal to two times the initial concentration of the soluble alkaline earth hydroxide salt dissolved.

Neutrally charged organic compounds containing at least one nitrogen atom generally behave as weak bases. The nitrogen atom has an unshared pair of electrons around it. This lone pair of electrons is used to form a bond to H^+ .

Weak bases only partially react with water to produce OH^- . To determine the amount of OH^- produced by the weak base (and, in turn, the pH of the solution), we set-up the ICE table using the K_b reaction of the weak base. The typical weak base equilibrium expression is:

$$K_{\rm b} = \frac{x^2}{0.25 - x} \approx \frac{x^2}{0.25}$$
 (assuming [B]₀ = 0.25 *M*)

Solving for *x* gives us the $[OH^-]$ in solution. We generally assume that weak bases are less than 5% reacted with water and that the OH^- contribution from water is negligible. The 5% assumption makes the math easier. By assuming an expression like $0.25 M - x \approx 0.25 M$, the calculation is very straight forward. The 5% rule applied here is that if $(x/0.25) \times 100$ is less than 5%, the assumption is valid. When the assumption is not valid, then we solve the equilibrium expression exactly using the quadratic equation (or by the method of successive approximations).

7. Monoprotic acid: An acid with one acidic proton; the general formula for monoprotic acids is HA.

Diprotic acid: An acid with two acidic protons (H₂A).

Triprotic acid: An acid with three acidic protons (H₃A).

$\mathrm{H}_2\mathrm{SO}_4(aq) \rightarrow \mathrm{HSO}_4^-(aq) + \mathrm{H}^+(aq)$	$K_{a_1} >> 1$; this is a strong acid.
$\mathrm{HSO}_{4^{-}}(aq) \rightleftharpoons \mathrm{SO}_{4^{2^{-}}}(aq) + \mathrm{H}^{+}(aq)$	$K_{a_2} = 0.012$; this is a weak acid.

When H_2SO_4 is dissolved in water, the first proton is assumed 100% dissociated because H_2SO_4 is a strong acid. After H_2SO_4 dissociates, we have H^+ and HSO_4^- present. HSO_4^- is a weak acid and can donate some more protons to water. To determine the amount of H^+ donated by HSO_4^- , one must solve an equilibrium problem using the K_{a_1} reaction for HSO_4^- .

$$H_{3}PO_{4}(aq) \rightleftharpoons H^{+}(aq) + H_{2}PO_{4}^{-}(aq) \qquad K_{a_{1}} = 7.5 \times 10^{-3}$$

$$H_{2}PO_{4}^{-}(aq) \rightleftharpoons H^{+}(aq) + HPO_{4}^{2-}(aq) \qquad K_{a_{2}} = 6.2 \times 10^{-8}$$

$$HPO_{4}^{2-}(aq) \rightleftharpoons H^{+}(aq) + PO_{4}^{3-}(aq) \qquad K_{a_{2}} = 4.8 \times 10^{-13}$$

When H₃PO₄ is added to water, the three acids that are present are H₃PO₄, H₂PO₄⁻, and HPO₄²⁻. H₃PO₄, with the largest K_a value, is the strongest of these weak acids. The conjugate bases of the three acids are H₂PO₄⁻, HPO₄²⁻, and PO₄³⁻. Because HPO₄²⁻ is the weakest acid (smallest K_a value), its conjugate base (PO₄³⁻) will have the largest K_b value and is the strongest base.

See Examples 13.15-13.17 of the text on the strategies used to solve for the pH of polyprotic acids. The strategy to solve most polyprotic acid solutions is covered in Example 13.15. For typical polyprotic acids, $K_{a_1} \gg K_{a_2}$ (and K_{a_3} if a triprotic acid). Because of this, the dominant producer of H⁺ in solution is just the K_{a_1} reaction. We set-up the equilibrium problem using the K_{a_1} reaction and solve for H⁺. We then assume that the H⁺ donated by the K_{a_2} (and K_{a_3} if triprotic) reaction is negligible, that is, the H⁺ donated by the K_{a_1} reaction is assumed to be the H⁺ donated by the entire acid system. This assumption is great when $K_{a_1} \gg K_{a_2}$ (roughly a 1000 fold difference in magnitude).

Examples 13.16 and 13.17 cover strategies for the other type of polyprotic acid problems. This other type is solutions of H_2SO_4 . As discussed previously, H_2SO_4 problems are both a strong acid and a weak acid problem in one. To solve for the [H⁺], we sometimes must worry about the H⁺ contribution from HSO_4^- . Example 13.16 is an example of an H_2SO_4 solution where the HSO_4^- contribution of H⁺ can be ignored. Example 13.17 illustrates an H_2SO_4 problem where we can't ignore the H⁺ contribution from HSO_4^- .

8. a. H_2O and $CH_3CO_2^-$

- b. An acid-base reaction can be thought of as a competition between two opposing bases. Since this equilibrium lies far to the left ($K_a < 1$), CH₃CO₂⁻ is a stronger base than H₂O.
- c. The acetate ion is a better base than water and produces basic solutions in water. When we put acetate ion into solution as the only major basic species, the reaction is:

$$CH_3CO_2^- + H_2O \rightleftharpoons CH_3CO_2H + OH^-$$

Now the competition is between $CH_3CO_2^-$ and OH^- for the proton. Hydroxide ion is the strongest base possible in water. The equilibrium above lies far to the left, resulting in a K_b value less than one. Those species we specifically call weak bases $(10^{-14} < K_b < 1)$ lie between H_2O and OH^- in base strength. Weak bases are stronger bases than water but are weaker bases than OH^- .

The NH₄⁺ ion is a weak acid because it lies between H₂O and H₃O⁺ (H⁺) in terms of acid strength. Weak acids are better acids than water, thus their aqueous solutions are acidic. They are weak acids because they are not as strong as H₃O⁺ (H⁺). Weak acids only partially dissociate in water and have K_a values between 10⁻¹⁴ and 1.

For a strong acid HX having $K_a = 1 \times 10^6$, the conjugate base, X⁻, has $K_b = K_w/K_a = 1.0 \times 10^{-14}/1 \times 10^6 = 1 \times 10^{-20}$ (a very small value).

The conjugate bases of strong acids have extremely small values for K_b ; so small that they are worse bases than water ($K_b \ll K_w$). Therefore, conjugate bases of strong acids have no basic properties in water. They are present, but they only balance charge in solution and nothing else. The conjugate bases of the six strong acids are Cl⁻, Br⁻, I⁻, NO₃⁻, ClO₄⁻, and HSO₄⁻.

Summarizing the acid-base properties of conjugates:

- a. The conjugate base of a weak acid is a weak base $(10^{-14} < K_b < 1)$
- b. The conjugate acid of a weak base is a weak acid $(10^{-14} < K_a < 1)$
- c. The conjugate base of a strong acid is a worthless base ($K_b \ll 10^{-14}$)
- d. The conjugate acid of a strong base is a worthless acid ($K_a \ll 10^{-14}$)

Identifying/recognizing the acid-base properties of conjugates is crucial in order to understand the acid-base properties of salts. The salts we will give you will be salts containing the conjugates discussed above. Your job is to recognize the type of conjugate present, and then use that information to solve an equilibrium problem (if necessary).

9. A salt is an ionic compound composed of a cation and an anion.

Weak base anions: These are the conjugate bases of the weak acids having the HA general formula. Table 13.2 of the text lists several HA type acids. Some weak base anions derived from the acids in Table 13.2 are ClO_2^- , F^- , NO_2^- , $C_2H_3O_2^-$, OCl^- , and CN^- .

Garbage anions (those anions with no basic or acidic properties): These are the conjugate bases of the strong acids having the HA general formula. These are Cl^- , NO_3^- , Br^- , I^- , and ClO_4^- .

Weak acid cations: These are the conjugate acids of the weak bases which contain nitrogen. Table 13.3 of the text lists several nitrogen-containing bases. Some weak acid cations derived from the weak bases in Table 13.3 are NH_4^+ , $CH_3NH_3^+$, $C_2H_5NH_3^+$, $C_6H_5NH_3^+$, and $C_5H_5NH^+$.

Garbage cations (those cations with no acidic properties or basic properties): The most common ones used are the cations in the strong bases. These are Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Ca²⁺, Sr²⁺, and Ba²⁺.

We mix and match the cations and anions to get what type of salt we want. For a weak base salt, we combine a weak base anion with a garbage cation. Some weak base salts are NaF, KNO₂, Ca(CN)₂, and RbC₂H₃O₂. To determine the pH of a weak base salt, we write out the K_b reaction for the weak base anion and determine K_b (= K_w/K_a). We set up the ICE table under the K_b reaction, and then solve the equilibrium problem to calculate [OH⁻] and, in turn, pH.

For a weak acid salt, we combine a weak acid cation with a garbage anion. Some weak acid salts are NH₄Cl, C₅H₅NHNO₃, CH₃NH₃I, and C₂H₅NH₃ClO₄. To determine the pH, we write out the K_a reaction for the weak acid cation and determine K_a (= K_w/K_b). We set up the ICE table under the K_a reaction, and then solve the equilibrium problem to calculate [H⁺] and, in turn, pH.

For a neutral (pH = 7.0) salt, we combine a garbage cation with a garbage anion. Some examples are NaCl, KNO₃, BaBr₂, and Sr(ClO₄)₂.

For salts that contain a weak acid cation and a weak base anion, we compare the K_a value of the weak acid cation to the K_b value for the weak base anion. When $K_a > K_b$, the salt produces an acidic solution (pH < 7.0). When $K_b > K_a$, the salt produces a basic solution. And when $K_a = K_b$, the salt produces a neutral solution (pH = 7.0).

Text Homework

- 20. Acidic solutions (at 25°C) have an $[H^+] > 1.0 \times 10^{-7} M$, which gives a pH < 7.0. Because $[H^+][OH^-] = 1.0 \times 10^{-14}$ and pH + pOH = 14.00 for an aqueous solution at 25°C, an acidic solution must also have $[OH^-] < 1.0 \times 10^{-7} M$ and pOH > 7.00. From these relationships, the solutions in parts a, b, and d are acidic. The solution in part c will have a pH > 7.0 (pH = 14.00 4.51 = 9.49) and is therefore not acidic (solution is basic).
- 21. Basic solutions (at 25°C) have an $[OH^-] > 1.0 \times 10^{-7} M$, which gives a pOH < 7.0. Because $[H^+][OH^-] = 1.0 \times 10^{-14}$ and pH + pOH = 14.00 for any aqueous solution at 25°C, a basic solution must also have $[H^+] < 1.0 \times 10^{-7} M$ and pH > 7.00. From these relationships, the solutions in parts b, c, and d are basic solutions. The solution in part a will have a pH < 7.0 (pH = 14.00 11.21 = 2.79) and is therefore not basic (solution is acidic).
- 23. 10.78 (4 S.F.); 6.78 (3 S.F.); 0.78 (2 S.F.); a pH value is a logarithm. The numbers to the left of the decimal point identify the power of 10 to which [H⁺] is expressed in scientific notation, for example, 10⁻¹¹, 10⁻⁷, 10⁻¹. The number of decimal places in a pH value identifies the number of significant figures in [H⁺]. In all three pH values, the [H⁺] should be expressed only to two significant figures because these pH values have only two decimal places.
- a. These are solutions of strong acids like HCl, HBr, HI, HNO₃, H₂SO₄, and HClO₄. So
 0.10 *M* solutions of any of the acids would be examples of a strong electrolyte solution that is very acidic.
 - b. These are solutions containing salts of the conjugate acids of the bases in Table 13-3. These conjugate acids are all weak acids, and they are cations with a 1+ charge. NH₄Cl, CH₃NH₃NO₃,

and $C_2H_5NH_3Br$ are three examples of this type of slightly acidic salts. Note that the anions used to form these salts are conjugate bases of strong acids; this is so because they have no acidic or basic properties in water (except for HSO_4^- , which has weak acid properties).

- c. These are solutions of strong bases like LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)₂, Sr(OH)₂, and Ba(OH)₂. All these strong bases are strong electrolytes.
- d. These are solutions containing salts of the conjugate bases of the neutrally charged weak acids in Table 13-2. These conjugate bases are all weak bases, and they are anions with a 1- charge. Three examples of this type of slightly basic salts are NaClO₂, KC₂H₃O₂, and CaF₂. The cations used to form these salts are Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Ca²⁺, Sr²⁺, and Ba²⁺ because these cations have no acidic or basic properties in water. Notice that these are the cations of the strong bases you should memorize.
- e. There are two ways to make a neutral salt solution. The easiest way is to combine a conjugate base of a strong acid (except for HSO_4^-) with one of the cations from a strong base. These ions have no acidic/basic properties in water, so salts of these ions are neutral. Three examples are NaCl, KNO₃, and SrI₂. Another type of strong electrolyte that can produce neutral solutions are salts that contain an ion with weak acid properties combined with an ion of opposite charge having weak base properties. If the K_a for the weak acid ion is equal to the K_b for the weak base ion, then the salt will produce a neutral solution. The most common example of this type of salt is ammonium acetate (NH₄C₂H₃O₂). For this salt, K_a for NH₄⁺ = K_b for C₂H₃O₂⁻ = 5.6 $\times 10^{-10}$. This salt at any concentration produces a neutral solution.

29. a.
$$H_2O(1) + H_2O(1) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$
 or

 $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \quad K = K_w = [H^+][OH^-]$

b. $HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$ or

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq) \qquad K = K_a = \frac{[H^+][F^-]}{[HF]}$$

- c. $C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq)$ $K = K_b = \frac{[C_5H_5NH^+][OH^-]}{[C_5H_5N]}$
- 30. Only statement a is true (assuming the species is not amphoteric). You cannot add a base to water and get an acidic pH (pH < 7.0). For statement b, you can have negative pH values; this just indicates an $[H^+] > 1.0 M$. For statement c, a dilute solution of a strong acid can have a higher pH than a more concentrated weak acid solution. For statement d, the Ba(OH)₂ solution will have an $[OH^-]$ twice of the same concentration of KOH, but this does not correspond to a pOH value twice that of the same concentration of KOH (prove it to yourselves).
- 32. H₂CO₃ is a weak acid with $K_{a_1} = 4.3 \times 10^{-7}$ and $K_{a_2} = 5.6 \times 10^{-11}$. The [H⁺] concentration in solution will be determined from the K_{a_1} reaction because $K_{a_1} >> K_{a_2}$. Because $K_{a_1} << 1$, the [H⁺] < 0.10 *M*; only a small percentage of the 0.10 *M* H₂CO₃ will dissociate into HCO₃⁻ and H⁺. So statement a best describes the 0.10 *M* H₂CO₃ solution. H₂SO₄ is a strong acid as well as a very good weak acid ($K_{a_1} >> 1$, $K_{a_2} = 1.2 \times 10^{-2}$). All the 0.10 *M* H₂SO₄ solution will dissociate into 0.10 *M* H⁺ and 0.10 *M* HSO₄⁻. However, because HSO₄⁻ is a good weak acid due to the relatively large

 K_a value, some of the 0.10 *M* HSO₄⁻ will dissociate into some more H⁺ and SO₄²⁻. Therefore, the [H⁺] will be greater than 0.10 *M* but will not reach 0.20 *M* because only some of 0.10 *M* HSO₄⁻ will dissociate. Statement c is best for a 0.10 *M* H₂SO₄ solution.

- 34. a. Sulfur reacts with oxygen to produce SO₂ and SO₃. These sulfur oxides both react with water to produce H₂SO₃ and H₂SO₄, respectively. Acid rain can result when sulfur emissions are not controlled. Note that, in general, nonmetal oxides react with water to produce acidic solutions.
 - b. CaO reacts with water to produce Ca(OH)₂, a strong base. A gardener mixes lime (CaO) into soil in order to raise the pH of the soil. The effect of adding lime is to add Ca(OH)₂. Note that, in general, metal oxides react with water to produce basic solutions.
- a. HClO₄(aq) + H₂O(l) → H₃O⁺(aq) + ClO₄⁻(aq). Only the forward reaction is indicated because HClO₄ is a strong acid and is basically 100% dissociated in water. For acids, the dissociation reaction is commonly written without water as a reactant. The common abbreviation for this reaction is HClO₄(aq) → H⁺(aq) + ClO₄⁻(aq). This reaction is also called the K_a reaction because the equilibrium constant for this reaction is designated as K_a.
 - b. Propanoic acid is a weak acid, so it is only partially dissociated in water. The dissocia-tion reaction is:

 $CH_3CH_2CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CH_2CO_2^-(aq)$ or

 $CH_3CH_2CO_2H(aq) \rightleftharpoons H^+(aq) + CH_3CH_2CO_2^-(aq).$

c. NH_4^+ is a weak acid. Similar to propanoic acid, the dissociation reaction is:

 $NH_4^+(aq) + H_2O(1) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$ or $NH_4^+(aq) \rightleftharpoons H^+(aq) + NH_3(aq)$

37. An acid is a proton (H⁺) donor, and a base is a proton acceptor. A conjugate acid-base pair differs by only a proton (H⁺).

	Acid	Base	Conjugate Base of Acid	Conjugate Acid of Base
a.	H ₂ CO ₃	H ₂ O	HCO ₃ -	$\mathrm{H_3O^+}$
b.	$C_5H_5NH^+$	H_2O	C_5H_5N	$\mathrm{H}_{3}\mathrm{O}^{+}$
c.	$C_5H_5NH^+$	HCO ₃ ⁻	C ₅ H ₅ N	H ₂ CO ₃

- 39. Strong acids have a K_a >> 1, and weak acids have K_a < 1. Table 13-2 in the text lists some K_a values for weak acids. K_a values for strong acids are hard to determine, so they are not listed in the text. However, there are only a few common strong acids so, if you memorize the strong acids, then all other acids will be weak acids. The strong acids to memorize are HCl, HBr, HI, HNO₃, HClO₄, and H₂SO₄.
 - a. HClO₄ is a strong acid.
 - b. HOCl is a weak acid ($K_a = 3.5 \times 10^{-8}$).

- c. H_2SO_4 is a strong acid.
- d. H_2SO_3 is a weak diprotic acid because the K_{a1} and K_{a2} values are much less than 1.
- 40. The beaker on the left represents a strong acid in solution; the acid HA is 100% dissociated into the H⁺ and A⁻ ions. The beaker on the right represents a weak acid in solution; only a little bit of the acid HB dissociates into ions, so the acid exists mostly as undissociated HB molecules in water.
 - a. HNO₂: weak acid beaker
 - b. HNO₃: strong acid beaker
 - c. HCl: strong acid beaker
 - d. HF: weak acid beaker
 - e. $HC_2H_3O_2$: weak acid beaker
- 41. The K_a value is directly related to acid strength. As K_a increases, acid strength increases. For water, use K_w when comparing the acid strength of water to other species. The K_a values are:

HClO₄: strong acid (K_a >> 1); HClO₂: $K_a = 1.2 \times 10^{-2}$

 NH_4^+ : $K_a = 5.6 \times 10^{-10}$; H_2O : $K_a = K_w = 1.0 \times 10^{-14}$

From the K_a values, the ordering is $HClO_4 > HClO_2 > NH_4^+ > H_2O$.

- 43. a. HCl is a strong acid, and water is a very weak acid with $K_a = K_w = 1.0 \times 10^{-14}$. HCl is a much stronger acid than H₂O.
 - b. H_2O , $K_a = K_w = 1.0 \times 10^{-14}$; HNO₂, $K_a = 4.0 \times 10^{-4}$; HNO₂ is a stronger acid than H₂O because K_a for HNO₂ > K_w for H₂O.
 - c. HOC_6H_5 , $K_a = 1.6 \times 10^{-10}$; HCN, $K_a = 6.2 \times 10^{-10}$; HCN is a slightly stronger acid than HOC_6H_5 because K_a for HCN > K_a for HOC₆H₅.
- 45. At 25°C, the relationship $[H^+][OH^-] = K_w = 1.0 \times 10^{-14}$ always holds for aqueous solutions. When $[H^+]$ is greater than $1.0 \times 10^{-7} M$, the solution is acidic; when $[H^+]$ is less than $1.0 \times 10^{-7} M$, the solution is basic; when $[H^+] = 1.0 \times 10^{-7} M$, the solution is neutral. In terms of $[OH^-]$, an acidic solution has $[OH^-] < 1.0 \times 10^{-7} M$, a basic solution has $[OH^-] > 1.0 \times 10^{-7} M$, and a neutral solution has $[OH^-] = 1.0 \times 10^{-7} M$.
 - a. $[OH^{-}] = \frac{K_w}{[H^{+}]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7} M$; the solution is neutral.
 - b. $[OH^{-}] = \frac{1.0 \times 10^{-14}}{8.3 \times 10^{-16}} = 12 M$; the solution is basic.
 - c. $[OH^{-}] = \frac{1.0 \times 10^{-14}}{12} = 8.3 \times 10^{-16} M$; the solution is acidic.

d.
$$[OH^{-}] = \frac{1.0 \times 10^{-14}}{5.4 \times 10^{-5}} = 1.9 \times 10^{-10} M$$
; the solution is acidic.

46. a.
$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.5} = 6.7 \times 10^{-15} M$$
; basic
b. $[H^+] = \frac{1.0 \times 10^{-14}}{3.6 \times 10^{-15}} = 2.8 M$; acidic
c. $[H^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7} M$; neutral
d. $[H^+] = \frac{1.0 \times 10^{-14}}{7.3 \times 10^{-4}} = 1.4 \times 10^{-11} M$; basic

53.
$$pOH = 14.0 - pH = 14.0 - 2.1 = 11.9; [H^+] = 10^{-pH} = 10^{-2.1} = 8 \times 10^{-3} M$$
 (1 sig. fig.)
 $[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{8 \times 10^{-3}} = 1 \times 10^{-12} M$ or $[OH^-] = 10^{-pOH} = 10^{-11.9} = 1 \times 10^{-12} M$

The sample of gastric juice is acidic because the pH is less than 7.00 at 25°C.

54.
$$pH = 14.00 - pOH = 14.00 - 5.74 = 8.26; [H^+] = 10^{-pH} = 10^{-8.26} = 5.5 \times 10^{-9} M$$

$$[OH^{-}] = \frac{K_{w}}{[H^{+}]} = \frac{1.0 \times 10^{-14}}{5.5 \times 10^{-9}} = 1.8 \times 10^{-6} M \text{ or } [OH^{-}] = 10^{-pOH} = 10^{-5.74} = 1.8 \times 10^{-6} M$$

The solution of baking soda is basic because the pH is greater than 7.00 at 25°C.

- 55. All the acids in this problem are strong acids that are always assumed to completely dissociate in water. The general dissociation reaction for a strong acid is $HA(aq) \rightarrow H^+(aq) + A^-(aq)$, where A^- is the conjugate base of the strong acid HA. For 0.250 *M* solutions of these strong acids, 0.250 *M* H⁺ and 0.250 *M* A⁻ are present when the acids completely dissociate. The amount of H⁺ donated from water will be insignificant in this problem since H₂O is a very weak acid.
 - a. Major species present after dissociation = H^+ , ClO_4^- , and H_2O ;

$$pH = -log[H^+] = -log(0.250) = 0.602$$

b. Major species = H^+ , NO_3^- , and H_2O ; pH = 0.602

$$[\mathrm{H}^+] = \frac{(2.5 \times 10^{-3} + 1.5 \times 10^{-2}) \,\mathrm{mol}}{0.2000 \,\mathrm{L}} = 0.088 \,M; \ \mathrm{pH} = -\log(0.088) = 1.06$$

- 57. Strong acids are assumed to completely dissociate in water; for example; $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ or $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$.
 - a. A 0.10 *M* HCl solution gives $0.10 M H^+$ and $0.10 M Cl^-$ because HCl completely dissociates. The amount of H⁺ from H₂O will be insignificant.

 $pH = -log[H^+] = -log(0.10) = 1.00$

- b. $5.0 M H^+$ is produced when $5.0 M HClO_4$ completely dissociates. The amount of H⁺ from H₂O will be insignificant. pH = $-\log(5.0) = -0.70$ (Negative pH values just indicate very concentrated acid solutions.)
- c. $1.0 \times 10^{-11} M \,\mathrm{H^+}$ is produced when $1.0 \times 10^{-11} M \,\mathrm{HI}$ completely dissociates. If you take the negative log of 1.0×10^{-11} , this gives pH = 11.00. This is impossible! We dissolved an acid in water and got a basic pH. What we must consider in this problem is that water by itself donates $1.0 \times 10^{-7} M \,\mathrm{H^+}$. We can normally ignore the small amount of H⁺ from H₂O except when we have a very dilute solution of an acid (as in the case here). Therefore, the pH is that of neutral water (pH = 7.00) because the amount of HI present is insignificant.
- 59. $[H^+] = 10^{-pH} = 10^{-2.50} = 3.2 \times 10^{-3} M$. Because HI is a strong acid, a $3.2 \times 10^{-3} M$ HI solution will produce $3.2 \times 10^{-3} M$ H⁺, giving a pH = 2.50.
- 64. a. HOC₆H₅ (K_a = 1.6×10^{-10}) and H₂O (K_a = K_w = 1.0×10^{-14}) are the major species. The major equilibrium is the dissociation of HOC₆H₅. Solving the weak acid problem:

$$HOC_{6}H_{5} \rightleftharpoons H^{+} + OC_{6}H_{5}^{-}$$
Initial 0.250 *M* ~0 0
x mol/L HOC_{6}H_{5} dissociates to reach equilibrium
Change $-x \rightarrow +x +x +x$
Equil. 0.250 $-x x x x$

$$K_{a} = 1.6 \times 10^{-10} = \frac{[H^{+}][OC_{6}H_{5}^{-}]}{[HOC_{6}H_{5}]} = \frac{x^{2}}{0.250 - x} \approx \frac{x^{2}}{0.250} \text{ (assuming } x << 0.250)$$

 $x = [H^+] = 6.3 \times 10^{-6} M$; checking assumption: x is 2.5×10^{-3} % of 0.250, so assumption is valid by the 5% rule.

 $pH = -\log(6.3 \times 10^{-6}) = 5.20$

b. HCN ($K_a = 6.2 \times 10^{-10}$) and H₂O are the major species. HCN is the major source of H⁺.

HCN ≓ H^+ +CN⁻ ~0 Initial 0.250 M0 x mol/L HCN dissociates to reach equilibrium Change -x \rightarrow +x+xEquil. 0.250 - xx х $K_{a} = 6.2 \times 10^{-10} = \frac{[H^{+}][CN^{-}]}{[HCN]} = \frac{x^{2}}{0.250 - x} \approx \frac{x^{2}}{0.250} \quad (assuming \ x << 0.250)$ $x = [H^+] = 1.2 \times 10^{-5} M$; checking assumption: x is 4.8×10^{-3} % of 0.250. Assumptions good. $pH = -log(1.2 \times 10^{-5}) = 4.92$

65. This is a weak acid in water. Solving the weak acid problem:

 $K_a = 7.2 \times 10^{-4}$ HF H^+ F^{-} ⇒ +Initial 0.020 M~0 0 mol/L HF dissociates to reach equilibrium Change +x-x \rightarrow +x0.020 - xEquil. х х $K_a = 7.2 \times 10^{-4} = \frac{[H^+][F^-]}{[HF]} = \frac{x^2}{0.020 - x} \gg \frac{x^2}{0.020}$ (assuming x << 0.020) $x = [H^+] = 3.8 \times 10^{-3} M$; check assumptions:

$$\frac{x}{0.020} \times 100 = \frac{3.8 \times 10^{-3}}{0.020} \times 100 = 19\%$$

The assumption $x \ll 0.020$ is not good (x is more than 5% of 0.020). We must solve $x^2/(0.020 - x) = 7.2 \times 10^{-4}$ exactly by using either the quadratic formula or the method of successive approximations (see Appendix 1 of the text). Using successive approximations, we let 0.016 *M* be a new approximation for [HF]. That is, in the denominator try x = 0.0038 (the value of x we calculated making the normal assumption) so that 0.020 - 0.0038 = 0.016; then solve for a new value of x in the numerator.

$$\frac{x^2}{0.020 - x} \approx \frac{x^2}{0.016} = 7.2 \times 10^{-4}, \ x = 3.4 \times 10^{-3}$$

We use this new value of x to further refine our estimate of [HF], that is, 0.020 - x = 0.020 - 0.0034 = 0.0166 (carrying an extra sig. fig.).

$$\frac{x^2}{0.020 - x} \approx \frac{x^2}{0.0166} = 7.2 \times 10^{-4}, \ x = 3.5 \times 10^{-3}$$

We repeat until we get a self-consistent answer. This would be the same answer we would get solving exactly using the quadratic equation. In this case it is, $x = 3.5 \times 10^{-3}$. Thus:

$$[H^+] = [F^-] = x = 3.5 \times 10^{-3} M; [OH^-] = K_w/[H^+] = 2.9 \times 10^{-12} M$$

$$[HF] = 0.020 - x = 0.020 - 0.0035 = 0.017 M; pH = 2.46$$

Note: When the 5% assumption fails, use whichever method you are most comfortable with to solve exactly. The method of successive approximations is probably fastest when the percent error is less than $\sim 25\%$ (unless you have a graphing calculator).

67. $HC_3H_5O_2$ ($K_a = 1.3 \times 10^{-5}$) and H_2O ($K_a = K_w = 1.0 \times 10^{-14}$) are the major species present. $HC_3H_5O_2$ will be the dominant producer of H^+ because $HC_3H_5O_2$ is a stronger acid than H_2O . Solving the weak acid problem:

	$\mathrm{HC_{3}H_{5}O_{2}}$	≓	H^{+}	+	$C_3H_5O_2^-$	
Initial	0.100 M		~0		0	
	$x \mod/L H$	C_3H_5C	D ₂ disso	ociates	to reach	equilibrium
Change	-x	\rightarrow	+x		+x	
Equil.	0.100 - x		x		x	

$$\begin{aligned} \mathbf{K}_{a} &= 1.3 \times 10^{-5} = \frac{[\mathrm{H}^{+}][\mathrm{C}_{3}\mathrm{H}_{5}\mathrm{O}_{2}^{-1}]}{[\mathrm{HC}_{3}\mathrm{H}_{5}\mathrm{O}_{2}]} = \frac{x^{2}}{0.100 - x} \approx \frac{x^{2}}{0.100} \\ x &= [\mathrm{H}^{+}] = 1.1 \times 10^{-3} M; \ \mathrm{pH} = -\mathrm{log}(1.1 \times 10^{-3}) = 2.96 \\ \text{Assumption follows the 5% rule (x is 1.1% of 0.100).} \\ [\mathrm{H}^{+}] &= [\mathrm{C}_{3}\mathrm{H}_{5}\mathrm{O}_{2}^{-}] = 1.1 \times 10^{-3} M; \ [\mathrm{OH}^{-}] = \mathrm{K}_{w}/[\mathrm{H}^{+}] = 9.1 \times 10^{-12} M \\ [\mathrm{HC}_{3}\mathrm{H}_{5}\mathrm{O}_{2}] = 0.100 - 1.1 \times 10^{-3} = 0.099 M \\ \text{Percent dissociation} &= \frac{[\mathrm{H}^{+}]}{[\mathrm{HC}_{3}\mathrm{H}_{5}\mathrm{O}_{2}]_{0}} \times 100 = \frac{1.1 \times 10^{-3}}{0.100} \times 100 = 1.1\% \\ \mathrm{HX} \qquad \rightleftharpoons \qquad \mathrm{H}^{+} \qquad + X^{-} \\ \mathrm{Initial} \qquad \mathrm{I} \qquad \sim 0 \qquad 0 \qquad \text{where I} = [\mathrm{HX}]_{0} \\ x \mod/L \operatorname{HX} \operatorname{dissociates to reach equilibrium} \\ \mathrm{Change} \qquad -x \qquad \rightarrow \qquad +x \qquad +x \\ \mathrm{Equil.} \qquad \mathrm{I} - x \qquad x \qquad x \\ \mathrm{From the problem}, x = 0.25(\mathrm{I}) \ \text{and } \mathbf{I} - x = 0.30 M. \\ \mathrm{I} - 0.25(\mathrm{I}) = 0.30 M, \ \mathrm{I} = 0.40 M \ \text{and } x = 0.25(0.40 M) = 0.10 M \\ \mathrm{K}_{a} = \frac{[\mathrm{H}^{+}][\mathrm{X}^{-}]}{[\mathrm{HX}]} = \frac{x^{2}}{\mathrm{I} - x} = \frac{(0.10)^{2}}{0.30} = 0.033 \\ \end{array}$$

76.

87.

77. HClO₄ is a strong acid with $[H^+] = 0.040 M$. This equals the $[H^+]$ in the trichloroacetic acid. Set up the problem using the K_a equilibrium reaction for CCl₃CO₂H.

	CCl ₃ CO ₂ H	⇒	H^{+}	+	$CCl_3CO_2^-$
Initial	0.050 M		~0		0
Equil.	0.050 - x		x		x

$$K_{a} = \frac{[H^{+}][CCl_{3}CO_{2}^{-}]}{[CCl_{3}CO_{2}H]} = \frac{x^{2}}{0.050 - x}; \text{ from the problem, } x = [H^{+}] = 4.0 \times 10^{-2} M$$
$$K_{a} = \frac{(4.0 \times 10^{-2})^{2}}{0.050 - (4.0 \times 10^{-2})} = 0.16$$

83. All K_b reactions refer to the base reacting with water to produce the conjugate acid of the base and OH⁻.

a.
$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

b. $C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq)$
a. $C_6H_5NH_2$
b. $C_6H_5NH_2$
c. OH^-
d. CH_3NH_2

The base with the largest K_b value is the strongest base ($K_{b, C_6H_5NH_2} = 3.8 \times 10^{-10}$, $K_{b, CH_3NH_2} = 4.4 \times 10^{-4}$). OH⁻ is the strongest base possible in water.

- 89. NaOH(aq) → Na⁺(aq) + OH⁻(aq); NaOH is a strong base that completely dissociates into Na⁺ and OH⁻. The initial concentration of NaOH will equal the concentration of OH⁻ do-nated by NaOH.
 - a. $[OH^{-}] = 0.10 M$; $pOH = -log[OH^{-}] = -log(0.10) = 1.00$

pH = 14.00 - pOH = 14.00 - 1.00 = 13.00

Note that H_2O is also present, but the amount of OH^- produced by H_2O will be insignificant compared to the 0.10 *M* OH⁻ produced from the NaOH.

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

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

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

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

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

94.
$$pH = 10.50; pOH = 14.00 - 10.50 = 3.50; [OH-] = 10-3.50 = 3.2 \times 10-4 M$$

 $Sr(OH)_2(aq) \rightarrow Sr^{2+}(aq) + 2 OH^{-}(aq); Sr(OH)_2 \text{ donates } 2 \text{ mol } OH^{-} \text{ per mol } Sr(OH)_2.$

A $1.6 \times 10^{-4} M$ Sr(OH)₂ solution will produce a pH = 10.50 solution.

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

a.
$$(C_2H_5)_3N + H_2O \rightleftharpoons (C_2H_5)_3NH^+ + OH^- K_b = 4.0 \times 10^{-4}$$

Initial 0.20 *M* 0 ~0
x mol/L of $(C_2H_5)_3N$ reacts with H₂O to reach equilibrium
Change $-x \rightarrow +x +x +x$
Equil. $0.20 - x \qquad x \qquad x$
 $K_b = 4.0 \times 10^{-4} = \frac{[(C_2H_5)_3NH^+][OH^-]}{[(C_2H_5)_3N]} = \frac{x^2}{0.20 - x} \approx \frac{x^2}{0.20}$, $x = [OH^-] = 8.9 \times 10^{-3} M$

Assumptions good (x is 4.5% of 0.20). $[OH^{-}] = 8.9 \times 10^{-3} M$ $[H^{+}] = \frac{K_w}{[OH^{-}]} = \frac{1.0 \times 10^{-14}}{8.9 \times 10^{-3}} = 1.1 \times 10^{-12} M$; pH = 11.96 b. HONH₂ + H₂O \rightleftharpoons HONH₃⁺ + OH⁻ K_b = 1.1 × 10⁻⁸ Initial 0.20 M 0 ~0 Equil. 0.20 - x x x x K_b = 1.1 × 10⁻⁸ = $\frac{x^2}{0.20 - x} \approx \frac{x^2}{0.20}$, $x = [OH^{-}] = 4.7 \times 10^{-5} M$; assumptions good. $[H^{+}] = 2.1 \times 10^{-10} M$; pH = 9.68

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

$$C_{2}H_{5}NH_{2} + H_{2}O \rightleftharpoons C_{2}H_{5}NH_{3}^{+} + OH^{-} K_{b} = 5.6 \times 10^{-4}$$
Initial 0.20 *M* 0 ~0
x mol/L C_{2}H_{5}NH_{2} reacts with H_{2}O to reach equilibrium
Change -*x* $\rightarrow +x +x$
Equil. 0.20 - *x x x*
 $K_{b} = \frac{[C_{2}H_{5}NH_{3}^{+}][OH^{-}]}{[C_{2}H_{5}NH_{2}]} = \frac{x^{2}}{0.20 - x} \approx \frac{x^{2}}{0.20}$ (assuming *x* << 0.20)
 $x = 1.1 \times 10^{-2}$; 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 equa-tion or the method of successive approximations (see Appendix 1 of the text). Using successive approximations and carrying extra significant figures:

$$\frac{x^2}{0.20 - 0.011} = \frac{x^2}{0.189} = 5.6 \times 10^{-4}, \ x = 1.0 \times 10^{-2} M \text{ (consistent answer)}$$
$$x = [\text{OH}^-] = 1.0 \times 10^{-2} M; \ [\text{H}^+] = \frac{\text{K}_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12} M; \ \text{pH} = 12.00$$

103. Using the K_b reaction to solve where PT = p-toluidine ($CH_3C_6H_4NH_2$):

Because pH = 8.60: pOH = 14.00 - 8.60 = 5.40 and $[OH^{-}] = x = 10^{-5.40} = 4.0 \times 10^{-6} M$

$$K_{b} = \frac{(4.0 \times 10^{-6})^{2}}{0.016 - (4.0 \times 10^{-6})} = 1.0 \times 10^{-9}$$
104. HONH₂ + H₂O \rightleftharpoons HONH₃⁺ + OH⁻ K_b = 1.1 × 10⁻⁸
Initial I 0 ~0 ~0 I = [HONH₂]₀
Equil. I-x x x x
K_b = 1.1 × 10⁻⁸ = $\frac{x^{2}}{1-x}$
From problem, pH = 10.00, so pOH = 4.00 and x = [OH⁻] = 1.0 × 10⁻⁴ M.
1.1 × 10⁻⁸ = $\frac{(1.0 \times 10^{-4})^{2}}{1 - (1.0 \times 10^{-4})}$, I = 0.91 M
Mass HONH₂ = 0.2500 L × $\frac{0.91 \text{ mol HONH}_{2}}{L}$ × $\frac{33.03 \text{ g HONH}_{2}}{\text{mol HONH}_{2}}$ = 7.5 g HONH₂
106. H₃C₆H₅O₇(aq) \rightleftharpoons H₂C₆H₅O₇⁻(aq) + H⁺(aq) K_{a1} = $\frac{[H_{2}C_{6}H_{5}O_{7}^{-}][H^{+}]}{[H_{3}C_{6}H_{5}O_{7}^{-}]}$
H₂C₆H₅O₇⁻(aq) \rightleftharpoons H_CC₆H₅O₇²⁻(aq) + H⁺(aq) K_{a2} = $\frac{[HC_{6}H_{5}O_{7}^{-2}][H^{+}]}{[H_{2}C_{6}H_{5}O_{7}^{-}]}$

- 116. Because NH_3 is a weaker base (smaller K_b value) than CH_3NH_2 , the conjugate acid of NH_3 will be a stronger acid than the conjugate acid of CH_3NH_2 . Thus NH_4^+ is a stronger acid than $CH_3NH_3^+$.
- 124. BHCl \rightarrow BH⁺ + Cl⁻; Cl⁻ is the conjugate base of the strong acid HCl, so Cl⁻ has no acidic/ basic properties. BH⁺ is a weak acid because it is the conjugate acid of a weak base B. Determining the K_a value for BH⁺:

 BH^+ \Rightarrow В + H^+ Initial 0.10 M 0 ~0 $x \text{ mol/L BH}^+$ dissociates to reach equilibrium Change -x+x+x \rightarrow Equil. 0.10 - xx x $K_a = \frac{[B][H^+]}{[BH^+]} = \frac{x^2}{0.10 - x}$; from the problem, pH = 5.82: $[\mathrm{H}^+] = x = 10^{-5.82} = 1.5 \times 10^{-6} M; \quad \mathrm{K}_\mathrm{a} = \frac{(1.5 \times 10^{-6})^2}{0.10 - (1.5 \times 10^{-6})} = 2.3 \times 10^{-11}$ K_b for the base $B = K_w/K_a = (1.0 \times 10^{-14})/(2.3 \times 10^{-11}) = 4.3 \times 10^{-4}$.

From Table 13.3 of the text, this K_b value is closest to CH_3NH_2 , so the unknown salt is CH_3NH_3Cl .

- 136. a. Basic; $\text{Li}_2O(s) + H_2O(l) \rightarrow 2 \text{ LiOH}(aq)$; LiOH is a strong base.
 - b. Acidic; $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$; H_2CO_3 is a weak diprotic acid.
 - c. Basic; $SrO(s) + H_2O(l) \rightarrow Sr(OH)_2(aq)$; $Sr(OH)_2$ is a strong base.
- 143. At pH = 2.000, $[H^+] = 10^{-2.000} = 1.00 \times 10^{-2} M$ At pH = 4.000, $[H^+] = 10^{-4.000} = 1.00 \times 10^{-4} M$ Moles H⁺ present = 0.0100 L × $\frac{0.0100 \text{ mol } \text{H}^+}{\text{L}} = 1.00 \times 10^{-4} \text{ mol } \text{H}^+$ Let V = total volume of solution at pH = 4.000:

 $1.00 \times 10^{-4} \text{ mol/L} = \frac{1.00 \times 10^{-4} \text{ mol H}^+}{\text{V}}, \text{ V} = 1.00 \text{ L}$

Volume of water added = 1.00 L - 0.0100 L = 0.99 L = 990 mL

- 144. Conjugate acid-base pairs differ by an H^+ in the formula. Pairs in parts a, c, and d are conjugate acid-base pairs. For part b, HSO_4^- is the conjugate base of H_2SO_4 . In addition, HSO_4^- is the conjugate acid of SO_4^{2-} .
- 161. In deciding whether a substance is an acid or a base, strong or weak, you should keep in mind a couple of ideas:
 - (1) There are only a few common strong acids and strong bases, all of which should be memorized. Common strong acids = HCl, HBr, HI, HNO₃, HClO₄, and H₂SO₄. Common strong bases = LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)₂, Sr(OH)₂, and Ba(OH)₂.
 - (2) All other acids and bases are weak and will have K_a and K_b values of less than 1 but greater than K_w (1.0×10^{-14}). Reference Table 13.2 for K_a values for some weak acids and Table 13.3 for K_b values for some weak bases. There are too many weak acids and weak bases to memorize them all. Therefore, use the tables of K_a and K_b values to help you identify weak acids and weak bases. Appendix 5 contains more complete tables of K_a and K_b values.
 - a. weak acid ($K_a = 4.0 \times 10^{-4}$)
 - c. weak base (K_b = 4.38×10^{-4})
 - e. weak base ($K_b = 1.8 \times 10^{-5}$)
 - g. weak acid ($K_a = 1.8 \times 10^{-4}$)
 - i. strong acid

- b. strong acid
- d. strong base
- f. weak acid ($K_a = 7.2 \times 10^{-4}$)
- h. strong base