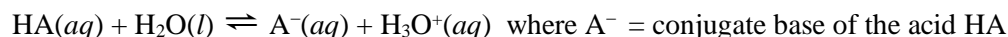# Chapter 13 Review Questions and Text Homework Solutions

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 (H_3O^+). For a general weak acid HA, the K_a reaction is:

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

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 sub-stance. For the general K_a reaction, the K_a expression is:

$$K_a = \frac{[A^-][H_3O^+]}{[HA]} \text{ or } K_a = \frac{[H^+][A^-]}{[HA]} \text{ (for the abbreviated } K_a \text{ reaction)}$$

    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:

$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq) \text{ where } BH^+ = \text{conjugate acid of the base B}$$
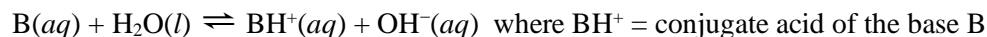
    d. The K_b equilibrium constant for the general K_b reaction is: $K_b = \dfrac{[BH^+][OH^-]}{[B]}$

    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, HNO_2, $HC_2H_3O_2$, and H_2SO_4 are Cl^-, NO_2^-, $C_2H_3O_2^-$, and HSO_4^-, respectively. The conjugate acids of the bases NH_3, C_5H_5N, and $HONH_2$ are NH_4^+, $C_5H_5NH^+$, and $HONH_3^+$, 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:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \text{ or } H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$
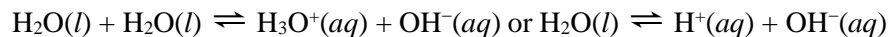
    c. The K_w equilibrium constant is also called the ion-product constant or the dissociation constant of water. It is the equilibrium constant for the autoionization reaction of water:

$$K_w = [H_3O^+][OH^-] \text{ or } K_w = [H^+][OH^-]$$

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 $pH + pOH = 14.00$

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

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 $K_a \gg 1$ 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 base, 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 H_2SO_4 can be different from this because H_2SO_4 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 H_2SO_4.

K_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_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):

$$K_a = \frac{x^2}{0.10 - x} \approx \frac{x^2}{0.10}$$

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:
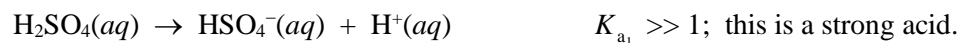
$$K_b = \frac{x^2}{0.25 - x} \approx \frac{x^2}{0.25} \quad \text{(assuming } [B]_0 = 0.25 \ M)$$

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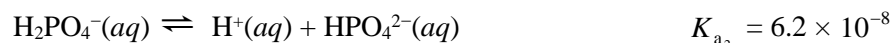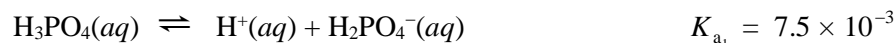
7. Monoprotic acid: An acid with one acidic proton; the general formula for monoprotic acids is HA.

Diprotic acid: An acid with two acidic protons (H_2A).

Triprotic acid: An acid with three acidic protons (H_3A).

$$H_2SO_4(aq) \rightarrow HSO_4^-(aq) + H^+(aq) \qquad\qquad K_{a_1} \gg 1; \text{ this is a strong acid.}$$

$$HSO_4^-(aq) \rightleftharpoons SO_4^{2-}(aq) + H^+(aq) \qquad\qquad K_{a_2} = 0.012; \text{ this is a weak acid.}$$

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_2} reaction for HSO_4^-.

$$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq) \qquad\qquad K_{a_1} = 7.5 \times 10^{-3}$$

$$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq) \qquad\qquad K_{a_2} = 6.2 \times 10^{-8}$$

$$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq) \qquad\qquad K_{a_3} = 4.8 \times 10^{-13}$$

When H_3PO_4 is added to water, the three acids that are present are H_3PO_4, $H_2PO_4^-$, and HPO_4^{2-}. H_3PO_4, with the largest K_a value, is the strongest of these weak acids. The conjugate bases of the three acids are $H_2PO_4^-$, HPO_4^{2-}, and PO_4^{3-}. Because HPO_4^{2-} is the weakest acid (smallest K_a value), its conjugate base (PO_4^{3-}) 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 ($K_a < 1$), $CH_3CO_2^-$ is a stronger base than H_2O.

    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:

    $$CH_3CO_2^- + H_2O \rightleftharpoons CH_3CO_2H + OH^-$$
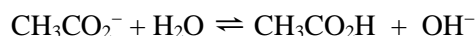
Now the competition is between $CH_3CO_2^-$ and OH^- for the proton. Hydroxide ion is the strongest base possible in water. The equilibrium above lies far to the left, resulting in a K_b value less than one. Those species we specifically call weak bases ($10^{-14} < K_b < 1$) lie between H_2O and OH^- in base strength. Weak bases are stronger bases than water but are weaker bases than OH^-.

The NH_4^+ ion is a weak acid because it lies between H_2O and H_3O^+ (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 H_3O^+ (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^-, NO_3^-, ClO_4^-, and HSO_4^-.

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^+, Ca^{2+}, Sr^{2+}, and Ba^{2+}.

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, KNO_2, $Ca(CN)_2$, and $RbC_2H_3O_2$. 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 NH_4Cl, $C_5H_5NHNO_3$, CH_3NH_3I, and $C_2H_5NH_3ClO_4$. 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 $NaCl$, KNO_3, $BaBr_2$, and $Sr(ClO_4)_2$.

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).

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. H_2CO_3 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 $K_{a_1} \gg K_{a_2}$. Because $K_{a_1} \ll 1$, the $[H^+]$ < 0.10 M; only a small percentage of the 0.10 M H_2CO_3 will dissociate into HCO_3^- and H^+. So statement a best describes the 0.10 M H_2CO_3 solution. H_2SO_4 is a strong acid as well as a very good weak acid ($K_{a_1} \gg 1$, $K_{a_2} = 1.2 \times 10^{-2}$). All the 0.10 M H_2SO_4 solution will dissociate into 0.10 M H^+ and 0.10 M HSO_4^-. However, because HSO_4^- is a good weak acid due to the relatively large K_a value, some of the 0.10 M HSO_4^- will dissociate into some more H^+ and SO_4^{2-}. Therefore, the $[H^+]$ will be greater than 0.10 M but will not reach 0.20 M because only some of 0.10 M HSO_4^- will dissociate. Statement c is best for a 0.10 M H_2SO_4 solution.

34.   a.   Sulfur reacts with oxygen to produce SO_2 and SO_3.   These sulfur oxides both react with water to produce H_2SO_3 and H_2SO_4, 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.

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. HNO_2: weak acid beaker

       b. HNO_3: strong acid beaker

       c.   HCl: strong acid beaker

       d.   HF: weak acid beaker

       e. $HC_2H_3O_2$: weak acid beaker

46.   a. $[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{1.0 \times 10^{-14}}{1.5} = 6.7 \times 10^{-15}\,M;$ basic

       b. $[H^+] = \dfrac{1.0 \times 10^{-14}}{3.6 \times 10^{-15}} = 2.8\,M;$ acidic

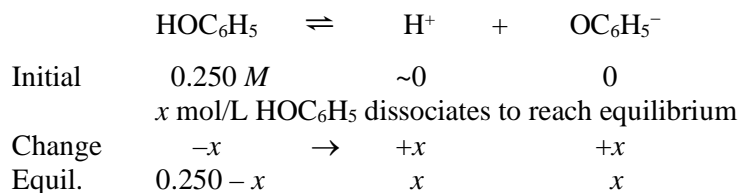   c. $[H^+] = \dfrac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7}\,M;$ neutral

       d. $[H^+] = \dfrac{1.0 \times 10^{-14}}{7.3 \times 10^{-4}} = 1.4 \times 10^{-11}\,M;$ basic

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

       $[OH^-] = \dfrac{K_w}{[H^+]} = \dfrac{1.0 \times 10^{-14}}{5.5 \times 10^{-9}} = 1.8 \times 10^{-6}\,M$ 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.

64.   a. HOC_6H_5 ($K_a = 1.6 \times 10^{-10}$) and H_2O ($K_a = K_w = 1.0 \times 10^{-14}$) are the major species. The major equilibrium is the dissociation of HOC_6H_5. Solving the weak acid problem:

$$HOC_6H_5 \ \rightleftharpoons \ H^+ \ + \ OC_6H_5^-$$

Initial	0.250 M	~0	0
	x mol/L HOC_6H_5 dissociates to reach equilibrium		
Change	$-x$ $\rightarrow$	$+x$	$+x$
Equil.	$0.250 - x$	x	x

$K_a = 1.6 \times 10^{-10} = \dfrac{[H^+][OC_6H_5^-]}{[HOC_6H_5]} = \dfrac{x^2}{0.250 - x} \approx \dfrac{x^2}{0.250}$ (assuming $x \ll 0.250$)

$x = [H^+] = 6.3 \times 10^{-6} M$; 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 H_2O are the major species. HCN is the major source of H^+.

$$HCN \rightleftharpoons H^+ + CN^-$$

Initial	0.250 M	~0	0

x mol/L HCN dissociates to reach equilibrium

Change	$-x$	$\rightarrow$	$+x$	$+x$
Equil.	$0.250 - x$		x	x

$$K_a = 6.2 \times 10^{-10} = \frac{[H^+][CN^-]}{[HCN]} = \frac{x^2}{0.250 - x} \approx \frac{x^2}{0.250} \quad (\text{assuming } x \ll 0.250)$$

$x = [H^+] = 1.2 \times 10^{-5} M$; checking assumption: x is 4.8×10^{-3}% of 0.250.

Assumptions good. $pH = -\log(1.2 \times 10^{-5}) = 4.92$

76.
$$HX \rightleftharpoons H^+ + X^-$$

Initial	I	~0	0	where I = $[HX]_0$

x mol/L HX dissociates to reach equilibrium

Change	$-x$	$\rightarrow$	$+x$	$+x$
Equil.	$I - x$		x	x

From the problem, $x = 0.25(I)$ and $I - x = 0.30 M$.

$I - 0.25(I) = 0.30 M$, $I = 0.40 M$ and $x = 0.25(0.40 M) = 0.10 M$

$$K_a = \frac{[H^+][X^-]}{[HX]} = \frac{x^2}{I - x} = \frac{(0.10)^2}{0.30} = 0.033$$

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

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

$$[Sr(OH)_2] = \frac{3.2 \times 10^{-4} \text{ mol OH} -}{L} \times \frac{1 \text{ mol Sr(OH)}_2}{2 \text{ mol OH}^-} = 1.6 \times 10^{-4} M \text{ Sr(OH)}_2$$

A $1.6 \times 10^{-4} M$ $Sr(OH)_2$ solution will produce a $pH = 10.50$ solution.

104.
$$HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^- \qquad K_b = 1.1 \times 10^{-8}$$

Initial	I		0	~0	I = $[HONH_2]_0$
Equil.	$I - x$		x	x

$$K_b = 1.1 \times 10^{-8} = \frac{x^2}{I - x}$$

From problem, $pH = 10.00$, so $pOH = 4.00$ and $x = [OH^-] = 1.0 \times 10^{-4} M$.

$$1.1 \times 10^{-8} = \frac{(1.0 \times 10^{-4})^2}{I - (1.0 \times 10^{-4})}, \quad I = 0.91 \ M$$

$$\text{Mass HONH}_2 = 0.2500 \ L \times \frac{0.91 \ \text{mol HONH}_2}{L} \times \frac{33.03 \ \text{g HONH}_2}{\text{mol HONH}_2} = 7.5 \ \text{g HONH}_2$$

106. $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq}) \rightleftharpoons \text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(\text{aq}) + \text{H}^+(\text{aq})$ 　　　$K_{a_1} = \dfrac{[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-][\text{H}^+]}{[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]}$

   $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(\text{aq}) \rightleftharpoons \text{HC}_6\text{H}_5\text{O}_7^{2-}(\text{aq}) + \text{H}^+(\text{aq})$ 　　　$K_{a_2} = \dfrac{[\text{HC}_6\text{H}_5\text{O}_7^{2-}][\text{H}^+]}{[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]}$

   $\text{HC}_6\text{H}_5\text{O}_7^{2-}(\text{aq}) \rightleftharpoons \text{C}_6\text{H}_5\text{O}_7^{3-}(\text{aq}) + \text{H}^+(\text{aq})$ 　　　$K_{a_3} = \dfrac{[\text{C}_6\text{H}_5\text{O}_7^{3-}][\text{H}^+]}{[\text{HC}_6\text{H}_5\text{O}_7^{2-}]}$
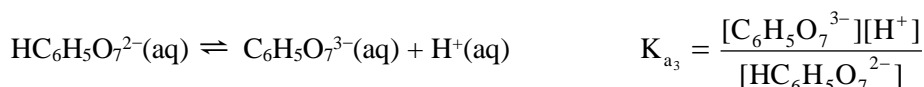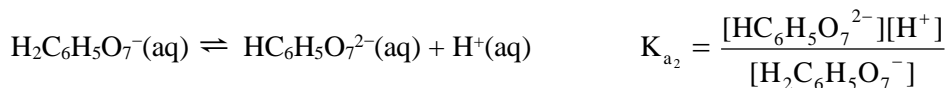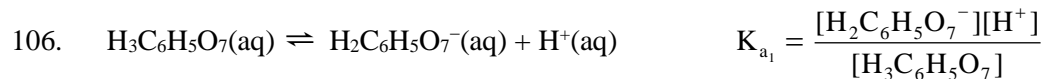
116.   Because NH_3 is a weaker base (smaller K_b value) than CH_3NH_2, the conjugate acid of NH_3 will be a stronger acid than the conjugate acid of CH_3NH_2. Thus NH_4^+ is a stronger acid than CH_3NH_3^+.

124. $\text{BHCl} \rightarrow \text{BH}^+ + \text{Cl}^-$; 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^+	$\rightleftharpoons$	B	+	H^+
Initial	0.10 M		0		~0

   x mol/L BH^+ dissociates to reach equilibrium

Change	$-x$	$\rightarrow$	$+x$		$+x$
Equil.	$0.10 - x$		x		x

   $K_a = \dfrac{[\text{B}][\text{H}^+]}{[\text{BH}^+]} = \dfrac{x^2}{0.10 - x}$; 　from the problem, pH = 5.82:

   $$[\text{H}^+] = x = 10^{-5.82} = 1.5 \times 10^{-6} \ M; \quad K_a = \frac{(1.5 \times 10^{-6})^2}{0.10 - (1.5 \times 10^{-6})} = 2.3 \times 10^{-11}$$

   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 $\text{CH}_3\text{NH}_3\text{Cl}$.

136.   a.　Basic; $\text{Li}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2 \ \text{LiOH}(\text{aq})$; LiOH is a strong base.

   b.　Acidic; $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{CO}_3(\text{aq})$; H_2CO_3 is a weak diprotic acid.

   c.　Basic; $\text{SrO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Sr(OH)}_2(\text{aq})$; Sr(OH)_2 is a strong base.

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-}.