15.98 The pH = 3.86 for a 0.064 M solution of a monoprotic acid. Is this a strong acid?

Determine the [H⁺] for this solution: [H⁺] = 10⁻ᵖᴴ = 10⁻³.86 = 1.4×10⁻⁴ M. Since this is << 0.064, the acid must be a weak acid. To further confirm, the % ionization would be:

%ionization = \frac{[H^+]}{[HA]₀} × 100% = \frac{1.4×10⁻⁴ M}{0.064 M} × 100% = 0.22% ≪ 100%

15.100 HA and HB are both weak acids, but HB is a stronger acid than HA. Will it take a larger volume of 0.1M NaOH to neutralize 50.0 mL of 0.10M HA or 50.0 mL or 0.10M HB?

The volume of NaOH required will be the same. Since a strong base (NaOH) is being added to the weak acids, both reactions will be driven to completion (100% to right).

HA + NaOH → H₂O + NaA
HB + NaOH → H₂O + NaB

At the equivalence point, the moles of acid must equal the moles of base. Thus, since the moles of acid is the same for both titrations, the moles of NaOH added must be the same for both titrations. Therefore the volume of base required to react with the same concentration of acid solutions (either both weak, both strong, or one strong and one weak) will be the same.

15.104 Use the \( K_a \) values of the respective acids to calculate the equilibrium constant for the following reaction:

\[
\text{CH}_3\text{COOH}(aq) + \text{NO}_2^-(aq) \rightleftharpoons \text{CH}_3\text{COO}^-(aq) + \text{HNO}_2(aq)
\]

We can write two equilibria that add up to the equilibrium in the problem.

\[
\begin{align*}
\text{CH}_3\text{COOH}(aq) & \rightleftharpoons H^+(aq) + \text{CH}_3\text{COO}^-(aq) \quad K_a = \frac{[H^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5} \\
H^+(aq) + \text{NO}_2^-(aq) & \rightleftharpoons \text{HNO}_2(aq) \quad K' = \frac{1}{K_a(\text{HNO}_2)} = \frac{1}{4.4 \times 10^{-4}} = 2.2 \times 10^3 \\
K' & = \frac{[\text{HNO}_2]}{[H^+][\text{NO}_2^-]}
\end{align*}
\]

\[
\text{CH}_3\text{COOH}(aq) + \text{NO}_2^-(aq) \rightleftharpoons \text{CH}_3\text{COO}^-(aq) + \text{HNO}_2(aq) \quad K = \frac{[\text{CH}_3\text{COO}^-][\text{HNO}_2]}{[\text{CH}_3\text{COOH}][\text{NO}_2^-]} = K_a \times K' = 4.0 \times 10^{-2}
\]

15.106 Determine the pH of a 0.20 M ammonium acetate (CH₃COONH₄) solution. Hint: Ammonium acetate consists of ammonium (weak acid, NH₄⁺) and acetate (a weak base, CH₃COO⁻). Look up or determine the \( K_a \) of ammonium and the \( K_b \) of acetate.

An ammonium acetate solution consists of the ions, NH₄⁺ and CH₃COO⁻. This is a salt solution.

Ammonium is a weak acid. \( K_a(\text{NH}_4^+) = 5.6 \times 10^{-10} \quad [K_b(\text{NH}_3) = 1.8 \times 10^{-5}] \)

Acetate is a weak base. \( K_a(\text{CH}_3\text{COO}^-) = 5.6 \times 10^{-10} \quad [K_b(\text{CH}_3\text{COO}^-) = 1.8 \times 10^{-5}] \)

In this specific case the \( K_a \) of ammonium ion is the same as the \( K_b \) of acetate ion. The two are of exactly (to two significant figures) equal strength. The solution will have pH 7.00.

However, be aware— not all salt solutions have a pH of 7!!!
15.107 Novacaine is a weak base ($K_b = 8.91 \times 10^{-6}$). What is the ratio of the concentration of the base to that of its acid in the blood plasma (pH = 7.40) of a patient? 
*Hint: Write the base as Nov and the acid as HNov$^\dagger$. You are looking for $[\text{Nov}] / [\text{HNov}^\dagger]$. Write out the reaction of Nov and H$_2$O. Write out the $K_b$ expression. Then, plug in the $K_b$ and the $[\text{OH}^-]$ and solve for the $[\text{Nov}] / [\text{HNov}^\dagger]$ ratio. (No ICE table is needed. You have all equilibrium concentrations.)* 

Then the equilibrium reaction is:

$\text{Nov}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HNov}^\dagger(\text{aq}) + \text{OH}^- (\text{aq})$

$pH = 7.40$ so $pOH = 14 - 7.40 = 6.60$

$[\text{OH}^-] = 10^{-6.60} = 2.51 \times 10^{-7}$

$K_b = \frac{[\text{HNov}^\dagger][\text{OH}^-]}{[\text{Nov}]}$, so $\frac{[\text{Nov}]}{[\text{HNov}^\dagger]} = \frac{[\text{OH}^-]}{K_b} = \frac{2.51 \times 10^{-7}}{8.91 \times 10^{-6}} = 0.028$

15.108 Which of the following is the stronger base: NF$_3$ or NH$_3$? (Hint: Think about which one is a stronger Lewis Base and use the concept that F is more electronegative than H.)

The fact that fluorine attracts electrons in a molecule more strongly than hydrogen causes NF$_3$ to be a poor electron pair donor and a poor base (poor Lewis base). NH$_3$ is the stronger base.

15.112 What is the pH of 250.0 mL of an aqueous solution containing 0.616 g of the strong acid trifluoromethane sulfonic acid?

First we must calculate the molarity of the trifluoromethane sulfonic acid. (MM = 150.1 g/mol)

$$Molarity = \frac{0.616 \text{ g} \times \frac{1 \text{ mol}}{150.1 \text{ g}}}{0.250 \text{ L}} = 0.0164 \text{ M}$$

Since trifluoromethane sulfonic acid is a strong acid and is 100% ionized, the $[\text{H}^+]$ is 0.0165 M.

$pH = -\log(0.0164) = 1.79$

15.117 Given the equation: $\text{HbH}^\dagger + \text{O}_2 \rightleftharpoons \text{HbO}_2 + \text{H}^\dagger$

(a) What form of hemoglobin is favored in the lungs where oxygen concentration is highest?

From the equilibrium equation, high oxygen concentration puts stress on the left side of the equilibrium and thus shifts the concentrations to the right to compensate. HbO$_2$ is favored.

(b) In body tissues, where the cells release CO$_2$, the blood is more acidic. What form of hemoglobin is favored under this condition?

High acid, $\text{H}^\dagger$ concentration, places stress on the right side of the equation forcing concentrations on the left side to increase, thus releasing oxygen and increasing the concentration of HbH$^\dagger$.

(c) When a person hyperventilates, the concentration of CO$_2$ in his or her blood decreases. How does this action affect the above equilibrium? Why should a person who is hyperventilating, breath into a paper bag? Removal of CO$_2$ decreases $\text{H}^\dagger$ (in the form of carbonic acid), thus shifting the reaction to the right. More HbO$_2$ will form. Breathing into a paper bag increases the concentration of CO$_2$ (re-breathing the exhaled CO$_2$), thus causing more O$_2$ to be released as explained above.
15.123 How many grams of NaCN would you need to dissolve in enough water to make exactly 250 mL of solution with a pH = 10.00? NaCN consists of Na\(^+\) and CN\(^-\). CN\(^-\) is a base. Write the reaction of CN\(^-\) in water and look up its K\(_b\). You can figure out the [OH\(^-\)] from pH. Use an ICE table and fill in what you know and solve for the initial [CN\(^-\)]. Use that concentration to find grams of NaCN.

When the pH=10.00, the pOH= 4.00. Thus, [OH\(^-\)] = 1.0 \times 10^{-4} M. The concentration of HCN must be the same. (Why?) If the concentration of NaCN is \(x\), the table looks like:

<table>
<thead>
<tr>
<th>Initial (M)</th>
<th>Change (M)</th>
<th>Equilibrium (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CN(^-)(aq) + H(_2)O(l) ⇌ HCN(aq) + OH(^-)(aq)</td>
<td>([-1.0 \times 10^{-4}]</td>
<td>[1.0 \times 10^{-4}]</td>
</tr>
<tr>
<td>(x)</td>
<td>(0)</td>
<td>((1.0 \times 10^{-4})^2)</td>
</tr>
<tr>
<td>(0)</td>
<td>(+1.0 \times 10^{-4})</td>
<td>((x - 1.0 \times 10^{-4}))</td>
</tr>
</tbody>
</table>

\[
K_b = \frac{[HCN][OH^-]}{[CN^-]} = \frac{(1.0 \times 10^{-4})^2}{(x - 1.0 \times 10^{-4})} = 2.0 \times 10^{-5}
\]

(Taken from Table 15.3)

\[
x = 6.0 \times 10^{-4} M = [CN^-]_0
\]

Amount of NaCN = 250 mL \times \frac{6.0 \times 10^{-4} \text{ mol NaCN}}{1000 \text{ mL}} \times \frac{49.01 \text{ g NaCN}}{1 \text{ mol NaCN}} = 7.4 \times 10^{-3} \text{ g NaCN}

15.125 Calculate the pH of a 1 L solution containing 0.150 mol of CH\(_3\)COOH and 0.100 mole of HCl. A weak acid (CH\(_3\)COOH) is being added to a strong acid (HCl). Thus, the pH is totally determined by the HCl because the contribution of H\(^+\) by the weak acid is negligible.

\([\text{HCl}] = 0.100 \text{ M}, \text{ so } [\text{H}^+] = 0.100 \text{ M} \quad \text{pH} = 1.000\]

15.129 Describe the hydration of SO\(_2\) as a Lewis acid–base reactions. Like carbon dioxide, sulfur dioxide behaves as a Lewis acid by accepting a pair of electrons from the Lewis base water. The Lewis acid–base adduct rearranges to form sulfurous acid in a manner exactly analogous to the rearrangement of the carbon dioxide–water adduct to form carbonic acid that is presented on page 665 of the textbook.

15.133 Both NH\(_2^+\) and N\(^3-\) ions do not exist in aqueous solutions because they are stronger bases than hydroxide.

Write equations showing the reactions of these ions in water.

(a) \(\text{NH}_2^+ + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{OH}^-\)

(b) \(\text{N}^3^- + 3\text{H}_2\text{O} \rightarrow \text{NH}_3 + 3\text{OH}^-\)

(b) Which ion is the stronger base?

\(\text{N}^3^-\) is the stronger base since each ion produces 3 OH\(^-\) ions.

15.138 Which of the following does not represent a Lewis acid–base reaction?

- a) \(\text{H}_2\text{O} + \text{H}^+ \rightleftharpoons \text{H}_3\text{O}^+\)
- b) \(\text{NH}_3 + \text{BF}_3 \rightleftharpoons \text{H}_3\text{NBF}_3\)
- c) \(\text{PF}_3 + \text{F}_2 \rightleftharpoons \text{PF}_5\)
- d) \(\text{Al(OH)}_3 + \text{OH}^- \rightleftharpoons \text{Al(OH)}_4^-\)

(c), PF\(_3\) + F\(_2\) → PF\(_5\), does not represent a Lewis acid–base reaction. In this reaction, the F–F single bond is broken and single bonds are formed between P and each F atom. For a Lewis acid–base reaction, the Lewis acid is an electron–pair acceptor and the Lewis base is an electron–pair donor.
15.139 True or False? If false explain why the statement is wrong

(a) All Lewis acids are Brønsted acids: False.
A Lewis acid such as CO\(_2\) is not a Brønsted acid. It does not have a hydrogen ion to donate.

(b) The conjugate base of an acid always carries a negative charge. False.
Consider the weak acid, NH\(_4^+\). The conjugate base of this acid is NH\(_3\), which is neutral.

(c) The percent ionization of a base increases with its concentration in solution. False.
The percent ionization of a base decreases with increasing concentration of base in solution.

(d) A solution of barium fluoride is acidic. False.
A solution of barium fluoride is basic. The fluoride ion, F\(^-\), is the conjugate base of the weak acid, HF. It will hydrolyze to produce OH\(^-\) ions.

A. Write balanced molecular and net ionic equations for the following acid-base reactions:

1. Sulfuric acid solution + sodium hydroxide solution
   Molecular: H\(_2\)SO\(_4\) (aq) + 2 NaOH (aq) → 2 H\(_2\)O (l) + Na\(_2\)SO\(_4\) (aq)
   Net Ionic: H\(^+\) (aq) + OH\(^-\) (aq) → H\(_2\)O (l)
   or:
   Molecular: H\(_2\)SO\(_4\) (aq) + NaOH (aq) → H\(_2\)O (l) + NaHSO\(_4\) (aq)
   Net Ionic: H\(^+\) (aq) + OH\(^-\) (aq) → H\(_2\)O (l)

2. Sodium hydroxide solution + nitrous acid solution
   Molecular: NaOH (aq) + HNO\(_2\) (aq) → NaNO\(_2\) (aq) + H\(_2\)O (l)
   Net Ionic: OH\(^-\) (aq) + HNO\(_2\) (aq) → NO\(_2^-\) (aq) H\(_2\)O (l)

3. Hydrochloric acid solution + magnesium hydroxide solid
   Molecular: 2 HCl (aq) + Mg(OH)\(_2\) (s) → MgCl\(_2\) (aq) + 2 H\(_2\)O (l)
   Net Ionic: 2 H\(^+\) (aq) + Mg(OH)\(_2\) (s) → Mg\(^2+\) (aq) + H\(_2\)O (l)