CHEMISTRY 104 – Help Sheet #12
Acids and Bases (Part II)

Do the topics appropriate for your course
https://clc.chem.wisc.edu (Resources page)
Prepared by Dr. Tony Jacob

Nuggets: Autoionization; pH equations; pH Calculations Strong Acids/Strong Bases; pH Calculations Weak Acids/Weak Bases; %ionization; Kn, Kb, pKn, pKb relationships; Which Side of Reaction Favored (revisited); Polyprotic Acids; Lewis acids/bases; Salts

WATER – AUTOIONIZATION (self-ionizing)

\[ \text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \] and \( K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \)

\[ K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C} \]

\[ x = 1.0 \times 10^{-7} = [\text{H}_3\text{O}^+] = [\text{OH}^-] \text{ in pure water} \]

pH EQUATIONS

\[ \text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{and} \quad \text{pOH} = -\log[\text{OH}^-]; \]

\[ [\text{H}_3\text{O}^+] = 1 \times 10^{-\text{pH}} \quad \text{and} \quad [\text{OH}^-] = 1 \times 10^{-\text{pOH}} \]

\[ K_w = [\text{H}_2\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ and } \text{pH} + \text{pOH} = 14 \]

\[ \text{pH < 7 } \rightarrow \text{acidic } ([\text{H}_3\text{O}^+] > [\text{OH}^-]) \]

\[ \text{pH > 7 } \rightarrow \text{basic } ([\text{OH}^-] > [\text{H}_3\text{O}^+]) \]

\[ \text{pH = 7 } \rightarrow \text{neutral } ([\text{OH}^-] = [\text{H}_3\text{O}^+]) \]

Note: \( K_w \) varies with \( T \); at 25°C \( K_w = 1.0 \times 10^{-14} \)

STRONG ACID pH Calculation

Example 1: a. What is the pH of a 0.0010M HBr solution? b. What is the pOH and [OH⁻]?

Answer 1 Part a: Step 1. Write reaction.

\[ \text{HBr} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Br}^- \]

Note: Strong acid (memorized) \( \rightarrow \) reacts/dissociates to completion \( \rightarrow 100\% \)

Step 2. Write an ICE Table.

\[
\begin{array}{ccc}
\text{ICE} & \text{HBr} & \text{H}_2\text{O} & \text{H}_3\text{O}^+ & \text{Br}^- \\
\text{I} & 0.0010 & 0 & 0 & 0 \\
\text{C} & -0.0010 & -0.0010 & +0.0010 & +0.0010 \\
\text{E} & 0 & 0 & 0.0010 & 0.0010 \\
\end{array}
\]

(Note: An ICE Table implies “equilibrium” but with strong acids they go to completion. Using an ICE Table is still helpful as it shows how the reaction proceeds.)

Step 3. \( \text{pH} = -\log[\text{H}_3\text{O}^+] \)

\( \text{pH} = -\log(0.0010) = 3.00 \)

NOTE: Since HBr is a strong acid, it is not necessary to write an ICE Table since all of the acid dissociates.

(e.g., 0.0010M HBr \( \rightarrow 0.0010M \text{H}_3\text{O}^+ \))

Answer 1 Part b:

\( \text{pOH} = 14.00 - \text{pH} \)

\( \text{pOH} = 14.00 - 3.00 = 11.00 \)

\[ [\text{OH}^-] = 10^{-\text{pOH}}; \quad [\text{OH}^-] = 10^{-11} = 1.00 \times 10^{-11} \]
IONIZATION CONSTANT EXPRESSION (revisited with calculations)

**WEAK ACIDS** – $K_a$ (acid ionization constant)

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq) \quad K_a = \frac{[H_3O^+][F^-]}{[HF]} \quad (K_a \text{ always written with acid as reactant})$$

**WEAK BASES** – $K_b$ (base ionization constant) for weak bases

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \quad K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \quad (K_b \text{ always written with base as reactant})$$

**Approximations:** Approximations can be made to simplify calculations; there are different methods to determine if an approximation is valid; follow the method provided by the instructor. The method used in the examples on this HelpSheet is:

If $100(K_a) < \text{[initial concentration]} \rightarrow$ then approximation is valid. When making approximations, only drop the “x” term when it is added to or subtracted from a number and not when it is by itself. For example, $0.10 - x$ becomes $0.10$ ($-x$ is dropped); $x^2$ stays as $x^2$ since it is not added/subtracted from a number. Likewise, $0.0015 + 3x$ becomes $0.0015$ ($+3x$ is dropped).

**WEAK ACID/BASE pH Calculations**

### Example 2: What is the pH of a 0.15M HF solution? $K_a$ for HF = $6.8 \times 10^{-4}$

**Answer 2:**

1. **Write reaction.** $HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$
2. **Write an ICE Table.**

<table>
<thead>
<tr>
<th></th>
<th>$HF$</th>
<th>$H_2O$</th>
<th>$H_3O^+$</th>
<th>$F^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.15</td>
<td>0.15</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>$x$</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>E</td>
<td>0.15</td>
<td>0</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

3. **Write $K_a$ expression.** $K_a = \frac{[H_3O^+][F^-]}{[HF]}$

4. **Substitute $K_a$ and values from ICE Table into $K_a$ expression.**

$$6.8 \times 10^{-4} = \frac{(x)(x)}{0.15-x}$$

5. **Check if approximation for $x$ is valid; if it is valid make the approximation.**

Check if $(100)K_a < [HF]_0$ is true; $(100)6.8 \times 10^{-4} = 0.15M \rightarrow$ true$ightarrow$

ok to make approximation: $6.8 \times 10^{-4} = \frac{(x)(x)}{0.15}$

6. **Solve for $x$.**

$$x^2 = 0.000102; x = 0.0101 = [H_3O^+];$$

**determine pH:** $pH = -\log[H_3O^+];$

$pH = -\log(0.0101) = 1.996;$

$pH = 2.00$

### Example 3: What is the pH of a 0.75M $NH_3$ solution? ($K_b = 1.8 \times 10^{-5}$)

**Answer 3:**

1. **Write reaction.** $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
2. **Write an ICE Table.**

<table>
<thead>
<tr>
<th></th>
<th>$NH_3$</th>
<th>$H_2O$</th>
<th>$NH_4^+$</th>
<th>$OH^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.75</td>
<td>0.75</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>$-x$</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>E</td>
<td>0.75</td>
<td>0</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

3. **Write $K_b$ expression.** $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$

4. **Substitute values into $K_b$.**

$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.75-x}$$

5. **Check if approximation for $x$ is valid; if it is valid make approximation.**

Check if $(100)K_b < [NH_3]_0$ is true; $(100)1.8 \times 10^{-5} < 0.75M \rightarrow$ true$ightarrow$

ok to make approximation: $1.8 \times 10^{-5} = \frac{(x)(x)}{0.75}$

6. **Solve for $x$.**

$$x = 0.00367 = [OH^-];$$

**determine pOH:** $pOH = -\log[OH^-];$

$pOH = -\log(0.00367) = 2.438$

$pH = 14.00 - 2.438 = 11.57$

### PERCENT IONIZATION

$$\% \text{ ionization} = \frac{\text{amount that reacts}}{\text{initial concentration}} \times 100\% \quad \% \text{ ionization} = \frac{x}{[\text{base}]_0} \times 100\% \quad ("x" \text{ from ICE table}; x = [H_3O^+] \text{ or } x = [OH^-])$$

- Comparing different acids/bases with the same concentration: **stronger acid/base $\Rightarrow \% \text{ ionization}$ ↑
- Comparing same acids/bases with different concentrations: **more dilute solution $\Rightarrow \% \text{ ionization}$ ↑

### Example 4: What is the percent ionization for 0.75M $NH_3$ (Example 3 above)?

**Answer 4:**

$$\% \text{ ionization} = \frac{[OH^-]}{[\text{base}]_0} \times 100\% = \frac{0.00367}{0.75} \times 100\% = 0.489\%$$
**Example 5:** a. Which solution has the greatest percent ionization, 0.0010M HF, 0.010M HF, 0.10M HF?

b. Which solution has the greatest percent ionization, 0.5M HF \((K_a = 6.8 \times 10^{-4})\), 0.5M HCN \((K_a = 3.3 \times 10^{-10})\), 0.5M CH₃COOH \((K_a = 1.8 \times 10^{-5})\)

**Answer 5:** a. 0.0010M HF {for the same acid as the solution becomes more dilute the %ionization increases}  
b. 0.5M HF {for different acids with the same concentrations the stronger acid will have a greater percent ionization}

**How are \(K_a\) and \(K_b\) related?** \(K_w = K_a \times K_b\) (\(K_a\) and \(K_b\) are a conjugate pair of acids/bases)

Shown below is the HF acid dissociation and the conjugate base F⁻ hydrolysis reactions. The two reactions are added and chemicals on both sides of the reaction canceled. What is left is the autoionization reaction for water! Recall when 2 reactions are added, the \(K_{\text{add}} = K_1 \times K_2\). Hence, \(K_w = K_a \times K_b\) where \(K_a\) and \(K_b\) are an acid/base conjugate pair.

\[
\begin{align*}
\text{HF(aq) + H}_2\text{O(l)} & \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq) & K_{a} \\
\text{F}^-(aq) + \text{H}_2\text{O(l)} & \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HF(aq) + OH}^-(aq) & K_{b} \\
& \equiv \text{H}_2\text{O(l) + H}_2\text{O(l) + F}^-(aq) + \text{H}_3\text{O}^+(aq) + \text{F}^-(aq) + \text{HF(aq) + OH}^-(aq) \\
& \equiv \text{H}_2\text{O(l) + H}_3\text{O}^+(aq) + \text{OH}^-(aq) & K_w
\end{align*}
\]

\(K_a\) and \(K_b\) EQUATIONS

\[
\begin{align*}
pK_a &= -\log K_a \quad \text{and} \quad pK_b = -\log K_b \\
K_a &= 10^{-pK_a} \quad \text{and} \quad K_b = 10^{-pK_b} \\
K_w &= K_a \times K_b = 1.0 \times 10^{-14} \quad \text{and} \quad pK_a + pK_b = 14
\end{align*}
\]

Relationships (revisited):

\(K_a \uparrow\) or \(pK_a \downarrow\) \(\Rightarrow\) acid strength \(\uparrow\), \(\text{pH} \downarrow\), \([\text{H}_3\text{O}^+] \uparrow\), \(\text{pOH} \uparrow\), \([\text{OH}^-] \downarrow\), conductivity \(\uparrow\), \([\text{ions}] \uparrow\)

\(K_b \uparrow\) or \(pK_b \downarrow\) \(\Rightarrow\) base strength \(\uparrow\), \(\text{pH} \uparrow\), \([\text{H}_3\text{O}^+] \downarrow\), \(\text{pOH} \downarrow\), \([\text{OH}^-] \uparrow\), conductivity \(\uparrow\), \([\text{ions}] \uparrow\)

**Example 6:** Determine the \(pK_a\), \(K_a\), and \(pK_b\) of NH₃. \((K_b = 1.8 \times 10^{-5})\)

**Answer 6:** Find \(K_a\)

\[
\begin{align*}
pK_b &= -\log K_b \\
pK_b &= -(1.8 \times 10^{-5}) = 4.74 \\
K_b &= 5.6 \times 10^{-10}
\end{align*}
\]

\[
\begin{align*}
K_a &= K_w/K_b = 1.0 \times 10^{-14}/1.8 \times 10^{-5} \\
K_a &= 5.6 \times 10^{-10}
\end{align*}
\]

\[
\begin{align*}
pK_a &= -\log K_a \\
pK_a &= -(5.6 \times 10^{-10}) = 9.26 \quad \text{(or}\ pK_a = 14 - pK_b) \\
pK_a &= -\log K_a
\end{align*}
\]

**WHICH WAY DOES THE REACTION PROCEED?** (revisited) **Side with the weaker acid/weaker base is favored**

1. **Strong acid/base present** – side **without** the strong acid/strong base is the **favored** side

   **Example 7:** Which side of this reaction is favored? \(\text{H}_2\text{O(l) + NaCH}_3\text{COO(aq) } \rightleftharpoons \text{NaOH(aq) + CH}_3\text{COOH(aq)}\)

   **Answer 7:** Since NaOH is a strong base the reaction goes to the right (away from the strong base); left side/reactant side is favored

2. **Only weak acid/base present** (no strong acid/strong base present) – calculations rather than a table of acid/base strengths

   **Example 8:** Which side of this reaction is favored? \(\text{HF(aq) + NH}_3\text{(aq) } \rightleftharpoons \text{F}^-(aq) + \text{NH}_4^+(aq)\)

   \(\text{HF, } K_a = 7.2 \times 10^{-4}\) and \(\text{NH}_3, K_b = 1.8 \times 10^{-5}\)

   **Answer 8:**

   Acids: \(K_a = 7.2 \times 10^{-4}\) and for \(\text{NH}_4^+:\ K_a = K_w/K_b = 1.0 \times 10^{-14}/1.8 \times 10^{-5} = 5.6 \times 10^{-10} \Rightarrow \text{NH}_4^+\) is the weaker acid so the right side/product side is favored.

   Bases: \(\text{NH}_3, K_b = 1.8 \times 10^{-5}\) and for \(\text{F}^-:\ K_b = K_w/K_a = 1.0 \times 10^{-14}/7.2 \times 10^{-4} = 1.4 \times 10^{-11}\) and \(\text{F}^-\) is the weaker base so the right side/product side is favored.

   Only **one comparison is needed not both:** either \(\text{NH}_4^+\) is weaker than HF – or – \(\text{F}^-\) is weaker than \(\text{NH}_3\) (both will yield the same results; both were presented just to show that the same answer is obtained.)
POLYPROTIC ACIDS

Monoprotic: one H⁺ can dissociate (e.g., HF); Diprotic: two H⁺ can dissociate (e.g., H₂C₂O₄)

Triprotic: many H⁺ can dissociate (e.g., H₃PO₄); Polyprotic means more than one H⁺

The Kₐ for each successive H⁺ dissociation becomes smaller. Hence, when doing calculations with polyprotic acids, the first H⁺ dissociation dominates and the next H⁺ dissociation is ignored.

Example of polyprotic acid, H₃PO₄, acid and base reactions:

\[
\begin{array}{c|c|c}
\text{H₃PO₄} + \text{H₂O} \rightleftharpoons \text{H₃O}^+ + \text{H₂PO₄}^- & \text{K}_{a1} = 7.5 \times 10^{-3} & \text{PO₄}^{3-} + \text{H₂O} \rightleftharpoons \text{PO₄}^{2-} + \text{OH}^- & \text{K}_{b1} = \text{K}_w/\text{K}_{a3} = 1.0 \times 10^{-14}/3.6 \times 10^{-13} = 2.8 \times 10^{-2} \\
\text{H₂PO₄}^- + \text{H₂O} \rightleftharpoons \text{H₂O}^+ + \text{HPO₄}^{2-} & \text{K}_{a2} = 6.2 \times 10^{-8} & & \\
\text{HPO₄}^{2-} + \text{H₂O} \rightleftharpoons \text{H₂O}^+ + \text{PO₄}^{3-} & \text{K}_{a3} = 3.6 \times 10^{-13} & \\
\end{array}
\]

A way to remember which Kₐ or Kₐ goes with which reaction: H₃PO₄ \( \xrightarrow[][]{K_a} \) H₂PO₄⁻ \( \xrightarrow[]{K_a} \) HPO₄²⁻ \( \xrightarrow[]{K_a} \) PO₄³⁻.

Example 9: a. What is the pH of a 1.5M H₂SeO₃(aq) (selenous acid) solution? b. What is [SeO₃²⁻]? Kₐ₁ = 2.4 \times 10^{-3}; Kₐ₂ = 4.8 \times 10^{-9}

Answer 9: a. pH = 1.22

Write first H⁺ dissociation reaction: H₂SeO₃(aq) + H₂O(l) \( \rightleftharpoons \) HSeO₃⁻(aq) + H₂O⁺(aq); write ICE Table; to assist with which K to use: H₂SeO₃ \( \xrightarrow[]{K_{a1}} \) HSeO₃⁻ \( \xrightarrow[]{K_{b1}} \) SeO₃²⁻; Kₐ₁ = \[\frac{[\text{HSeO}_3^-][\text{H}_2\text{O}^+]}{[\text{H}_2\text{SeO}_3]}\] \( \approx \) \( \frac{x(x)}{1.50-x} \); since (100)2.4 \times 10^{-3} = 0.24 \times 1.5 = [H₂SeO₃], an approximation is valid;

\[
\begin{align*}
\text{I} & \quad 1.50 & \quad \text{X} & \quad \text{O} & \quad \text{O} \\
\text{C} & \quad x & \quad -x & \quad +x & \quad +x \\
\text{E} & \quad 1.50-x & \quad x & \quad x & \quad x \\
\end{align*}
\]

make approximation: 2.4 \times 10^{-3} \approx \frac{x^2}{1.50}; solve for x: x = 0.0600 = [H₃O⁺]; pH = -\log(0.0600) = 1.22;

HSeO₃⁻ can lose the 2nd H⁺: HSeO₃⁻ + H₂O \( \rightleftharpoons \) H₂O⁺ + SeO₃²⁻; with polyprotic acids the 1st H⁺ dominates and the 2nd H⁺ is ignored

b. [SeO₃²⁻] = 4.8 \times 10^{-9}M  (for a diprotic acid, H₂A, the [A²⁻] = Kₐ₂)

Write 2nd H⁺ dissociation reaction: HSeO₃⁻(aq) + H₂O(l) \( \rightleftharpoons \) SeO₃²⁻(aq) + H₂O⁺(aq); write an ICE Table with [HSeO₃⁻] = [H₂O⁺] = 0.0600M

from above ICE Table; Kₐ₂ = \[\frac{[\text{SeO}_3^{2-}][\text{H}_2\text{O}^+]}{[\text{HSeO}_3^-]}\] \( \approx \) \( \frac{(y)(0.0600+y)}{(0.0600-y)} \);

\[
\begin{align*}
\text{I} & \quad 0.0600 & \quad 0.0600 & \quad 0 \\
\text{C} & \quad y & \quad -y & \quad +y & \quad +y \\
\text{E} & \quad 0.0600-y & \quad 0.0600+y & \quad y \\
\end{align*}
\]

since (100)4.8 \times 10^{-9} = 4.8 \times 10^{-7} < 0.0600 = [HSeO₃⁻], an approximation is valid;

make approximation: 4.8 \times 10^{-9} \approx \frac{(y)(0.0600)}{(0.0600)}; y = 4.8 \times 10^{-9} = [\text{SeO}_3^{2-}] = K_{a2}; note: there is no significant change to [H₂O⁺] after the 2nd dissociation (0.0600 + 4.8 \times 10^{-9} ≈ 0.0600) which is why the second dissociation of a polyprotic acid is ignored when determining the pH;

Hint: In a diprotic acid, H₂A, the [A²⁻] = Kₐ₂; hence, the above ICE Table and calculations were not actually needed!

DEFINITIONS – Lewis Acids/Bases (revisited)

Lewis: Acid accepts lone pair of electrons; base donates lone pair of electrons

Lewis Acid-Base interactions: When a pair of e⁻ is donated from one species (base) to another (acid) the bond is called a coordinate covalent bond.

Common Lewis acids: metal cations (often transition metal cations) that can accept electrons because they are positive species; spectator ions (e.g., Li⁺, Na⁺) are not Lewis acids; the greater the metal charge the metal the stronger the Lewis acid (e.g., Al³⁺ is stronger than Be²⁺); the smaller the size of metal the stronger the Lewis acid (e.g., Be²⁺ is stronger than Mg²⁺)

Common Lewis bases: chemicals that can donate electron pairs; many anions are Lewis bases; other examples include: ammonia (NH₃), water (H₂O), hydroxide (OH⁻) – all of these chemical species have a lone pair of electrons to donate

SPECTATOR IONS: Cl⁻, Br⁻, I⁻, NO₃⁻, ClO₄⁻, Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Ca²⁺, Sr²⁺, Ba²⁺ – these ions are the conjugates of strong acids or bases and do not react with water (see below for examples); if these ions are present they can be ignored as far as contributing to the acidity or basicity in an aqueous solution.
**SALTS: Ionic compounds that can produce an acidic, basic, or neutral solution**

1. **SALTS that contain one or two spectator ions (salt is not from a polyprotic acid).**

   **Example 10:** Determine for each chemical whether it will produce an acidic, basic, or neutral solution. a. NaCl  b. NaF  c. NH₄NO₃

   **Answer 10a:** NaCl

   - Step 1. Let soluble salt break up: NaCl(aq) → Na⁺(aq) + Cl⁻(aq)
   - Step 2. Let each ion react with H₂O and see what forms:
     
     | Na⁺(aq) + H₂O(l) → NaOH(aq) + H⁺(aq) |
     | Cl⁻(aq) + H₂O(l) → HCl(aq) + OH⁻(aq) |

   **Hint:** If spectator ions are memorized ⇒ cross off all spectators; Na⁺(aq) and Cl⁻(aq) are crossed off

   **Step 3.** The formation of a strong acid or strong base is **not possible** so no reaction (NR) occurs in this case. If a weak acid or weak base forms, the reaction **can occur** since weak acids and weak bases can form in solution.

   - NaOH(aq) is SB: Na⁺(aq) + H₂O(l) → NR
   - HCl(aq) is SA: Cl⁻(aq) + H₂O(l) → NR

   **Step 4.** Neither H⁺ nor OH⁻ are formed ⇒ **neutral solution**

   **Answer 10b:** NaF

   - Cross off all spectator ions: Na⁺(aq) is crossed off ⇒ focus on F⁻(aq)

   **Step 2.** Let each ion react with H₂O and see what forms:

   | F⁻(aq) + H₂O(l) → HF(aq) + OH⁻(aq) |

   **Step 3.** The formation of a weak acid is possible

   - HF(aq) is a weak acid so **reaction occurs**

   **Step 4.** If OH⁻ is produced then solution is **basic**; if H₃O⁺ is produced then solution is **acidic**. OH⁻ was formed so F⁻(aq) ⇒ **basic solution**

   **Answer 10c:** NH₄NO₃

   - Cross off all spectator ions: NO₃⁻(aq) is crossed off ⇒ focus on NH₄⁺(aq)

   **Step 2.** Let each ion react: NH₄⁺(aq) + H₂O(l) → NH₃(aq) + H₃O⁺(aq)

   **Step 3.** The formation of a weak base (NH₃) is possible

   - NH₃(aq) + H₂O → NH₄⁺(aq)
   - NH₃(aq) is a weak base – **reaction occurs**

   **Step 4.** H₃O⁺ was formed; NH₄⁺(aq) ⇒ **acidic solution**

2. **SALTS that contain both weak acid and weak base can produce an acidic or basic soln. Which one?**

   **Compare K_a and K_b:** larger K wins! if K_a > K_b → acidic; if K_b > K_a → basic; if K_a = K_b → neutral

   **Example 11:** For NH₄F, determine whether it will produce an acidic, basic, or neutral solution. K_b NH₃ = 1.8 x 10⁻⁵; K_a HF = 6.8 x 10⁻⁴

   **Answer 11:**

   - Cross off all spectator ions: NO₃⁻(aq) is crossed off ⇒ focus on NH₄⁺(aq)

   **Step 2.** Let each ion react: NH₄⁺(aq) → F⁻(aq) + NH₄⁺(aq)

   - Basic reaction: F⁻ + H₂O → NH₃ + H₂O⁺

   **Step 3.** The formation of a weak acid (1st rxn) and weak base (2nd rxn) is possible; both rxns possible; use K_a and K_b to determine which dominates

   **Step 4.** Since K_a > K_b → NH₄F ⇒ **acidic solution**

3. **SALTS of polyprotic acids can produce either an acidic or basic soln. Which one?**

   **Compare K_a and K_b:** larger K wins! if K_a > K_b → acidic; if K_b > K_a → basic; if K_a = K_b → neutral

   **Example 12:** For Na₂HPO₄, determine whether it will produce an acidic, basic, or neutral solution. K_a₁ = 7.5 x 10⁻³; K_a₂ = 6.2 x 10⁻⁸; K_a₃ = 4.2 x 10⁻¹³

   **Answer 12:**

   - Cross out Na⁺ since it is a spectator ion

   **Step 2.** Let each ion react: HPO₄²⁻ + H₂O → H₂PO₄⁻ + PO₄³⁻

   - Basic reaction: HPO₄²⁻ + H₂O → H₂PO₄⁻ + OH⁻

   **Step 3.** The formation of a weak acid (1st rxn) and weak base (2nd rxn) is possible; both rxns possible; use K_a₁ and K_a₂ to determine which dominates

   **Step 4.** Since K_b₂ > K_a₃ → Na₂HPO₄ ⇒ **basic solution**

   **Determine K_a₁, K_a₂ for HPO₄²⁻:** K_a₁ HPO₄²⁻ = 1.0 x 10⁻¹⁴/1.8 x 10⁻⁵ = 5.6 x 10⁻¹⁰

   **Determine K_b₁, K_b₂ for F⁻:** K_b₁ HPO₄²⁻ = 1.0 x 10⁻¹⁴/6.8 x 10⁻⁵ = 1.5 x 10⁻¹¹
1. Calculate the pH, $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, and pOH for a $1.0 \times 10^{-3}$M solution of HCl.

2. I. What is the pH of a 0.065M Sr(OH)$_2$(aq) solution?
   a. 13.11  
   b. 0.89  
   c. 1.19  
   d. 12.81  
   e. 0.065

   II. 100.0ml of this Sr(OH)$_2$(aq) solution is diluted to 500.0ml. What is the pH of this new solution?
   a. 12.11  
   b. 1.59  
   c. 12.41  
   d. 1.89  
   e. 13.81

3. Which of the following solutions has the greatest acidity? (Multiple choice question.)
   a. A solution with pH = 2.65.
   b. A solution with $[\text{H}_3\text{O}^+] = 7.5 \times 10^{-3}$.
   c. A solution with $[\text{OH}^-] = 1.5 \times 10^{-12}$.
   d. A 0.0015M HNO$_3$(aq) solution.
   e. A solution with pOH = 1.80

4. Calculate the pH and percent ionization for each solution.
   I. 1.0M CH$_3$COOH ($K_a = 1.8 \times 10^{-5}$)
   III. 0.010M NH$_3$ ($K_b = 1.8 \times 10^{-5}$)
   II. 0.10M HF ($K_a = 6.8 \times 10^{-4}$)
   IV. 0.50M CH$_3$NH$_2$ ($K_b = 5.0 \times 10^{-4}$)

5. A 0.75M solution of HCNO(aq), cyanic acid, has pH = 1.79. What is the $K_a$ of HCNO(aq)?

6. Calculate the $K_b$ for each of the solutions.
   I. A 0.25M solution of trimethylamine, (CH$_3$)$_3$N, has a pH of 11.63. What is the $K_b$ for trimethylamine?
   II. A 1.15M solution of aniline, C$_6$H$_5$NH$_2$, has a pH of 9.35. What is the $K_b$ for aniline?

7. Five solutions are prepared: a 1.5M HCOOH, 0.15M HCOOH, 1.5M HOCl, 0.15M HOCl, and 1.5M HNO$_3$.
   HCOOH ($K_a = 3.0 \times 10^{-4}$) and HOCl ($K_a = 6.8 \times 10^{-8}$)
   Place the solutions in order from smallest to largest percent ionization. (Hint: No calculation is needed.)
   a. 1.5M HNO$_3 < 1.5M$ HCOOH < 0.15M HCOOH < 1.5M HOCl < 0.15M HOCl
   b. 0.15M HOCl < 1.5M HOCl < 0.15M HCOOH < 1.5M HCOOH < 1.5M HNO$_3$
   c. 1.5M HNO$_3 < 0.15M$ HCOOH < 1.5M HCOOH < 0.15M HOCl < 1.5M HOCl
   d. 1.5M HOCl < 0.15M HOCl < 1.5M HCOOH < 0.15M HCOOH < 1.5M HNO$_3$
   e. 0.15M HOCl < 0.15M HCOOH < 1.5M HOCl < 1.5M HCOOH < 1.5M HNO$_3$

8. a. The structure for allylamine is H$_2$C=CH—CH$_2$—NH$_2$. Draw the conjugate acid of this base.
   b. If the p$K_b$ for allylamine is 9.49, what is the $K_b$ for this base?
   c. What is the $K_a$ for this conjugate acid?
   d. What is the p$K_a$ for this conjugate acid?
9. For each reaction will the product-side or reactant-side be favored? (a few harder ones that require a calculation)
I. HCN(aq) + NH₃(aq) ⇌ CN⁻(aq) + NH₄⁺(aq) \[K_a \text{ HCN} = 3.3 \times 10^{-10}; K_b = NH_3 = 1.8 \times 10^{-5}\]
II. HCN(aq) + SO₄²⁻(aq) ⇌ CN⁻(aq) + HSO₄⁻(aq) \[K_a \text{ HCN} = 3.3 \times 10^{-10}; K_{a2} = HSO_4^- = 1.1 \times 10^{-2}\]
III. CH₃NH₂(aq) + HF(aq) ⇌ CH₃NH₃⁺(aq) + F⁻(aq) \[K_a \text{ HF} = 6.8 \times 10^{-4}; K_b = CH₃NH₂ = 5.0 \times 10^{-4}\]
IV. (CH₃)₂NH₂⁺(aq) + NO₂⁻(aq) ⇌ (CH₃)₂NH(aq) + HNO₂(aq) \[K_a \text{ HNO}_2 = 7.4 \times 10^{-4}; K_b (\text{CH₃}_2\text{NH}) = 5.8 \times 10^{-4}\]

10. What is the pH of a 0.025M solution of sodium propionate, NaCH₃CH₂COO? What is the concentration of propionic acid in the solution? (Kₐ CH₃CH₂COOH, is 1.3 x 10⁻⁵)

11. Identify the solutions made from the following chemicals as being acidic, basic, or neutral?
   a. Na₂S(aq)   b. NaCl(aq)   c. NH₃(aq)   d. H₃PO₄(aq)   e. NaF(aq)   f. NH₄NO₃(aq)
   g. NH(CH₃)₂(aq)   h. LiNO₃(aq)

12. Determine if a solution of NH₄F(aq) is acidic, basic, or neutral. (HF: Kₐ = 6.8 x 10⁻⁴; NH₃: Kₐ = 1.8 x 10⁻⁵)

13. Succinic acid is a diprotic acid with the formula, H₂C₄H₄O₄. (Multiple choice question.)
I. Which reaction below represents the Kₐ reaction?
II. Which reaction below has the largest Kₐ value?
   a. HC₄H₄O₄⁻(aq) + H₂O(l) ⇌ C₄H₄O₄²⁻(aq) + H₃O⁺(aq)
   b. HC₄H₄O₄⁻(aq) + H₂O(l) ⇌ H₂C₄H₄O₄⁴⁻(aq) + OH⁻(aq)
   c. HC₄H₄O₄⁻(aq) + OH⁻(aq) ⇌ C₄H₄O₄²⁻(aq) + H₂O(l)
   d. C₄H₄O₄²⁻(aq) + H₂O(l) ⇌ HC₄H₄O₄⁴⁻(aq) + OH⁻(aq)
   e. HC₄H₄O₄⁻(aq) + H₂O(l) ⇌ H₂C₄H₄O₄⁴⁻(aq) + OH⁻(aq)

14. a. What is the pH of a 0.75M H₂C₃H₂O₄(aq) (malonic acid) solution? Kₐ1 = 1.5 x 10⁻³; Kₐ2 = 2.0 x 10⁻⁶
   b. Calculate the concentration of [C₃H₂O₄²⁻].

15. a. HCO₃⁻ is amphiprotic and can donate or accept H⁺. Write the two reactions showing HCO₃⁻ acting as an acid and acting as a base.
   b. Given Kₐ₁ = 4.2 x 10⁻⁷ and Kₐ₂ = 4.8 x 10⁻¹¹ will HCO₃⁻ form an acidic or basic solution?

16. a. H₂PO₄⁻ is amphiprotic and can donate or accept H⁺. Write the two reactions showing H₂PO₄⁻ acting as an acid and acting as a base.
   b. Given Kₐ₁ = 7.5 x 10⁻³, Kₐ₂ = 6.2 x 10⁻⁸, and Kₐ₃ = 3.6 x 10⁻¹³ will H₂PO₄⁻ form an acidic or basic solution?

17. Which chemical does not act as a Lewis acid?
   a. Fe³⁺
   b. Cu(NO₃)₂
   c. Co(NO₃)₃
   d. H⁺
   e. PH₃
18. Using the following reaction,
\[ \text{Cr(H}_2\text{O)}_6^{3+} (\text{aq}) + 4\text{NH}_3 (\text{aq}) \rightleftharpoons \text{Cr(NH}_3)_6^{3+} (\text{aq}) + 6\text{H}_2\text{O(l)}, \]
which of the following statements are correct?

I. \( \text{NH}_3 (\text{aq}) \) is acting as a Lewis base
II. \( \text{H}_2\text{O(l)} \) in \( \text{Cr(H}_2\text{O)}_6^{3+} (\text{aq}) \) is acting as a Lewis base
III. \( \text{NH}_3 (\text{aq}) \) is acting as a Bronsted base
IV. \( \text{H}_2\text{O(l)} \) is acting as a Bronsted acid
V. \( \text{Cr}^{3+} (\text{aq}) \) in \( \text{Cr(H}_2\text{O)}_6^{3+} (\text{aq}) \) is acting as a Lewis acid

(a) I
(b) I and II
(c) I, II, III
(d) I, II, IV, V
(e) I, II, V

**ANSWERS**

1. \( \text{pH} = 3.00, \left[ \text{H}_3\text{O}^+ \right] = 1.0 \times 10^{-3} \text{ M}, \left[ \text{OH}^- \right] = 1.0 \times 10^{-11} \text{ M}, \text{pOH} = 11.00 \)

\[ \{\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^- \text{(100\%-SA)}; \left[ \text{H}_3\text{O}^+ \right] = 1.0 \times 10^{-3} \text{M}; \text{pH} = -\log(1.0 \times 10^{-3}) = 3.00; \text{pOH} = 14.00 - 3 = 11.00; \]
\[ \left[ \text{OH}^- \right] = 1.0 \times 10^{-11} \text{M} \}

2. I. a (write reaction: \( \text{Sr(OH)}_2 \rightleftharpoons \text{Sr}^{2+} + 2\text{OH}^- \); write ICE Table;
\[ \left[ \text{OH}^- \right] = 2(0.065) = 0.13 \text{M}; \text{pOH} = -\log(0.13) = 0.89; \text{pH} = 13.11 \} \]
II. c (dilution: \( M_1V_1 = M_2V_2; M_1 = \left[ \text{OH}^- \right] = 0.13 \text{M}; V_1 = 100.0 \text{ml}; V_2 = 500.0 \text{ml}; \)
\[ (0.13 \text{M})(100.0 \text{ml}) = M_2(500.0 \text{ml}); M_2 = \left[ \text{OH}^- \right]_{\text{diluted}} = 0.026 \text{M}; \]
\[ \text{pOH} = -\log(0.026) = 1.59; \text{pH} = 12.41 \} \]

3. b (greatest acidity = smallest pH; “a”: pH = 2.65; “b”: pH = \(-\log(7.5 \times 10^{-3}) = 2.12; \)
“c”: pH = \(-\log(1.5 \times 10^{-12}) = 11.82; \text{pH} = 14.00 - 11.82 = 2.18; “d”: [\text{H}_3\text{O}^+] = 0.0015 \text{M so pH} = -\log(0.0015) = 2.82; \)
“e”: pH = 14.00 – 1.80 = 12.20; solution “b” has the lowest pH and therefore the greatest acidity)
4. I. pH = 2.37; %ionization = 0.42% \{CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-\}; write ICE Table;

\[
K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}; \quad 1.8 \times 10^{-5} = \frac{(x)(x)}{(1.0-x)}; \quad 1.8 \times 10^{-5} = \frac{x^2}{(1.0-x)}; \quad \text{CH}_3\text{COOH}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CH}_3\text{COO}^-(aq)
\]

\[
\begin{array}{cccc}
I & 1.0 & \cdots & 0 \\
C & -x & -x & +x \\
E & 1.0-x & \cdots & x & x \\
\end{array}
\]

(100)\(K_a\) < [HA]₀; 100(1.8 \times 10^{-5}) = 1.8 \times 10^{-3} < 1.0 \rightarrow \text{approximation is valid};

make approximation: 1.8 \times 10^{-5} = \frac{x^2}{1.0}; \quad x = 0.00424 = [H_3O^+]; \quad \text{pH} = -\log(0.00424) = 2.37;

%ionization = \left(\frac{x}{1.0}\right) \times 100%; \quad \%\text{ionization} = \left(\frac{0.00424}{1.0}\right) \times 100% = 0.424%

II. pH = 10.62; %ionization = 4.2% \{NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-\}; write ICE Table;

\[
K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}; \quad K_b = \frac{(x)(x)}{(0.010 - x)}; \quad K_b = \frac{x^2}{(0.010 - x)}; \quad (100)K_b < [\text{Base}]_0;
\]

100(1.8 \times 10^{-5}) = 1.8 \times 10^{-3} < 0.010 \rightarrow \text{approximation is valid};

make approximation: 1.8 \times 10^{-5} = \frac{x^2}{0.010}; \quad x = 4.24 \times 10^{-4} = [OH^-];

pOH = -\log(4.24 \times 10^{-4}) = 3.37; \quad \text{pH} = 14.00 - \text{pOH} = 14.00 - 3.37 = 10.63;

%ionization = \left(\frac{x}{1.0}\right) \times 100%; \quad \%\text{ionization} = \left(\frac{4.24 \times 10^{-4}}{0.010}\right) \times 100% = 4.24%

III. pH = 2.08; %ionization = 8.3% \{HF + H_2O \rightleftharpoons H_3O^+ + F^-\}; write ICE Table;

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}; \quad 6.8 \times 10^{-4} = \frac{(x)(x)}{(10.0-x)}; \quad 6.8 \times 10^{-4} = \frac{x^2}{(10.0-x)}; \quad (100)K_a < [\text{HA}]_0;
\]

100(6.8 \times 10^{-4}) = 6.8 \times 10^{-2} < 0.10 \rightarrow \text{approximation is valid};

make approximation: 6.8 \times 10^{-4} = \frac{x^2}{0.10}; \quad x = 0.00825 = [H_3O^+]; \quad \text{pH} = -\log(0.00825) = 2.08;

%ionization = \left(\frac{x}{1.0}\right) \times 100%; \quad \%\text{ionization} = \left(\frac{0.00825}{0.10}\right) \times 100% = 8.25%

IV. pH = 12.20; %ionization = 3.2% \{CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-\}; write ICE Table;

\[
K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}; \quad K_b = \frac{(x)(x)}{(0.5-x)}; \quad K_b = \frac{x^2}{(0.50-x)}; \quad (100)K_b < [\text{Base}]_0;
\]

100(5.0 \times 10^{-4}) = 5.0 \times 10^{-2} < 0.50 \rightarrow \text{approximation is valid};

make approximation: 5.0 \times 10^{-4} = \frac{x^2}{0.50}; \quad x = 1.58 \times 10^{-2} = [OH^-]; \quad \text{pOH} = -\log(1.58 \times 10^{-2}) = 1.80;

\[
\text{pH} = 14.00 - \text{pOH} = 14.00 - 1.80 = 12.20; \quad \%\text{ionization} = \left(\frac{x}{1.0}\right) \times 100%; \quad \%\text{ionization} = \left(\frac{1.58 \times 10^{-2}}{0.50}\right) \times 100% = 3.16%
\]

5. \(3.6 \times 10^{-4}\) \{HCNO + H_2O \rightleftharpoons CNO^- + H_3O^+\}; write ICE Table;

\[
K_a = \frac{[\text{H}_3\text{O}^+]\text{[CNO}^-]}{[\text{HCNO}]}; \quad K_a = \frac{(x)(x)}{(0.75-x)}; \quad K_a = \frac{x^2}{(0.75-x)};
\]

now the pH was given = 1.79; [H_3O^+] = 10^{-1.79} = 0.01620M = x; \quad K_a = \frac{(0.01620)^2}{(0.75-0.01620)} = 3.58 \times 10^{-4}
6. I. \(7.4 \times 10^{-5}\) \((\text{CH}_3)_2\text{N} + \text{H}_2\text{O} \rightleftharpoons (\text{CH}_3)_2\text{NH}^+ + \text{OH}^-\); write ICE Table;

\[
K_b = \frac{([\text{CH}_3)_2\text{NH}^+][\text{OH}^-]}{([\text{CH}_3)_2\text{N}]}; \quad K_b = \frac{x(x)}{(0.25-x)}; \quad K_b = \frac{x^2}{(0.25-x)};
\]

the pH was given = 11.63; \(p\text{OH} = 14.00-\text{pH} = 14.00-11.63 = 2.37 \to \)

\([\text{OH}^-] = 10^{-2.37} = 0.00427\text{M} = x \text{ since I have assigned x to be } [\text{OH}^-]; \quad K_b = \frac{(0.00427)^2}{(0.25-0.00427)} = 7.42 \times 10^{-5}\}

II. \(4.3 \times 10^{-10}\) \(\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^-\); write ICE Table;

\[
K_b = \frac{([\text{C}_6\text{H}_5\text{NH}_3^+][\text{OH}^-]}{([\text{C}_6\text{H}_5\text{NH}_2]}; \quad K_b = \frac{x(x)}{(1.15-x)}; \quad K_b = \frac{x^2}{(1.15-x)};
\]

the pH was given = 9.35; \(p\text{OH} = 14.00-\text{pH} = 14.00-9.35 = 4.65 \to [\text{OH}^-] = 10^{-4.65} = 2.22 \times 10^{-5}\text{M} = x \text{ since}

I have assigned x to be [OH^-]; \quad K_b = \frac{(2.22 \times 10^{-5})^2}{(1.15-2.22 \times 10^{-5})} = 4.30 \times 10^{-10}\}

7. \{\text{HNO}_3\} \text{ is a strong acid and therefore has } \%\text{ion} \text{ = 100}\% \text{ and } \text{the greatest } \%\text{ion}; \text{ the other solutions should be divided into the}

stronger weak acid, \(\text{HCOOH}\), and the weaker weak acid, \(\text{HCl}\); at this point: \(\text{HCl} < \text{HCOOH} < \text{HNO}_3\); \text{ to distinguish between}

the 2 concentrations containing the same weak acid, recall as [HCOOH] \(\downarrow \to \%\text{ion} \uparrow \text{ likewise, } [\text{HCl}] \downarrow \to \%\text{ion} \uparrow \text{; therefore: smallest } \%\text{ion} \text{ 1.5M HCl} < 0.15M \text{ HCl} < 1.5M \text{ HCOOH} < 0.15M \text{ HCOOH} < 1.5M \text{ HNO}_3 \text{ greatest } \%\text{ion}\}

8. a. \(\text{H}_2\text{C} \rightleftharpoons \text{CH}_2\text{H}_2\text{NH}_3^+\) \{add a \text{H}^+ to make the conjugate acid\};

b. \(3.2 \times 10^{-10}\) \(\{10^pK_b = K_b = 3.2 \times 10^{-10}\}\)

c. \(3.1 \times 10^{-5}\) \(K_a \times K_b = K_w; \quad K_a = 1.0 \times 10^{-14}/3.2 \times 10^{-10} = 3.13 \times 10^{-5}\}

d. 4.51 \{-\log K_a = pK_a = -\log(3.13 \times 10^{-5}) = 4.505\}

9. In all these problems find the \(K_a\) for the acid on the reactant side and the \(K_a\) for the conjugate acid on product side; reaction will go to the weaker acid side (can also compare \(K_b\)'s; will go toward the weaker base side)

I. reactant \{\(K_a\) for HCN = 3.3 \times 10^{-10}; \(K_b\) for NH\(_3\) = 1.8 \times 10^{-5}; find either \(K_b\) for CN\(^-\) or \(K_a\) for NH\(_4^+\); \(I'll\) do both but only need to do one or the other in this case\}; \(K_b\) for CN\(^-\) = 1.0 \times 10^{-14}/3.3 \times 10^{-10} = 3.03 \times 10^{-5};

\(K_a\) for NH\(_4^+\) = 1.0 \times 10^{-14}/1.8 \times 10^{-5} = 5.56 \times 10^{-10}; \text{ compare } K_a's \text{ or } K_b's; \text{ I'll compare } K_a's \to \text{ weaker acid is HCN (smaller } K_a\) \text{ and that is the favored side = left side (reactant)}\}

II. reactant \{comparing the acids: \(K_a\) for HCN = 3.3 \times 10^{-10}; \(K_a\) for HSO\(_4^+\) = 1.1 \times 10^{-2}; \text{ compare } K_a's \to \text{ weaker acid is HCN (smaller } K_a\) \text{ and that is the favored side = left side (reactant)}\}

III. product \{\(K_a\) for HF = 6.8 \times 10^{-4}; \(K_b\) for CH\(_3\)NH\(_2\) = 5.0 \times 10^{-4}; \text{ find either } K_b \text{ for F}^- \text{ or } K_a \text{ for CH\(_3\)NH\(_3^+\); \(I'll\) do both but only need to do one or the other in this case\}; \(K_b\) for F\(^-\) = 1.0 \times 10^{-14}/6.8 \times 10^{-4} = 1.5 \times 10^{-11};

\(K_a\) for CH\(_3\)NH\(_3^+\) = 1.0 \times 10^{-14}/5.0 \times 10^{-4} = 2.0 \times 10^{-11}; \text{ compare } K_a's \text{ or } K_b's; \text{ I'll compare } K_a's \to \text{ weaker acid is CH\(_3\)NH\(_3^+\) (smaller } K_a\) \text{ and that is the favored side = right side (product)}\}

IV. reactant \{HNO\(_2\) and (CH\(_3\)\(_2\)NH\(_2^+\) \text{ are both weak acids; need to determine the } K_a \text{ of (CH\(_3\)\(_2\)NH\(_2^+\) from the } K_b \text{ of}

(CH\(_3\)\(_2\)NH; \(K_a\) (CH\(_3\)\(_2\)NH\(_2^+\) = K_w/K_b = 1.0 \times 10^{-14}/5.8 \times 10^{-4} = 1.72 \times 10^{-11}; \text{ since } K_a \text{ HNO}_2 > K_a \text{ (CH\(_3\)\(_2\)NH\(_2^+\) reaction goes toward (CH\(_3\)\(_2\)NH\(_2^+\) which is the side with the weaker acid (reactant side)); could also have compared } K_b \text{ values (don't need to do both comparisons: either compare acids or bases!)); have } K_b \text{ of (CH\(_3\)\(_2\)NH and need to determine } K_b \text{ for NO}_2^- \text{ from the } K_a \text{ of HNO}_2; \text{ K}_b \text{ NO}_2^- = K_w/K_a = 1.0 \times 10^{-14}/7.4 \times 10^{-4} = 1.35 \times 10^{-11}; \text{ since } K_b \text{ (CH\(_3\)\(_2\)NH > K_b \text{ NO}_2^- \text{ reaction goes toward NO}_2^- \text{ which is the side with the weaker base (again reactant side)}\}

\[\text{(CH}_3\text{)_2N(aq) + H}_2\text{O(l) } \rightleftharpoons \text{(CH}_3\text{)_2NH}^+(aq) + \text{OH}^-(aq)\]

\[
\begin{array}{cccc}
& 1 & 0.25 & 0 & 0 \\
\text{I} & 0 & -x & x & +x \\
\text{C} & -x & -x & +x & +x \\
\text{E} & 0.25x & -x & x & x \\
\end{array}
\]

\[
\text{(CH}_3\text{)_2N(aq) + H}_2\text{O(l) } \rightleftharpoons \text{(CH}_3\text{)_2NH}^+(aq) + \text{OH}^-(aq)\]

\[
\begin{array}{cccc}
& 1 & 1.15 & 0 & 0 \\
\text{I} & 1.15 & -x & x & +x \\
\text{C} & -x & -x & +x & +x \\
\text{E} & 1.15-x & -x & x & x \\
\end{array}
\]
10. pH = 8.64, [CH₃CH₂COOH] = 4.38 x 10⁻⁶ M

\[ \text{Na}^+ : \text{NaCH}_3\text{CH}_2\text{COO}^{-}(aq) \rightarrow \text{Na}^+(aq) + \text{CH}_3\text{CH}_2\text{COO}^{-}(aq) \]; write reaction with water:

\[ \text{CH}_3\text{CH}_2\text{COO}^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{CH}_2\text{COO}(aq) + \text{OH}^-(aq) \]; write ICE table; write Kₜ expression:

\[ K_b = \frac{[\text{CH}_3\text{CH}_2\text{COO}^-][\text{OH}^-]}{[\text{CH}_3\text{CH}_2\text{COO}^-]} \]

determine Kₜ from Kₐ: \[ K_b = K_w/K_a = 1.0 \times 10^{-14}/1.3 \times 10^{-5} = 7.69 \times 10^{-10} \]

\[ 7.69 \times 10^{-10} = \frac{(x)(x)}{(0.025)}; \quad x^2 = 1.92 \times 10^{-11}; \quad x = 4.38 \times 10^{-6} \]

\[ [\text{CH}_3\text{CH}_2\text{COOH}] = x = 4.38 \times 10^{-6} \text{M} \]

11. a. basic \( \text{Na}^+ \) is a spectator ion – ignore; let S²⁻ react with H₂O: S²⁻(aq) + H₂O(l) \rightleftharpoons HS⁻(aq) + OH⁻(aq) and check if HS⁻ is a strong acid or base; it is not \( \text{OH}^- \) formed \( \rightarrow \text{basic} \)

b. neutral \{both Na⁺ and Cl⁻ are spectator ions \( \rightarrow \text{neutral} \} \)

c. basic \{memorize NH₃ as basic; or let NH₃ react with H₂O: NH₃(aq) + H₂O(l) \rightleftharpoons NH₄⁺(aq) + OH⁻(aq) and check if NH₄⁺ is a strong acid or base; it is not \( \text{OH}^- \) formed \( \rightarrow \text{basic} \}

d. acidic \{starts with a H \( \rightarrow \text{acidic} \} \)

e. basic \{Na⁺ is a spectator ion \( \rightarrow \) ignore; let F⁻ react with H₂O: F⁻(aq) + H₂O(l) \rightleftharpoons HF(aq) + OH⁻(aq) and check if HF is a strong acid or base; it is not \( \text{OH}^- \) formed \( \rightarrow \text{basic} \}

f. acidic \{NO₃⁻ is a spectator ion \( \rightarrow \) ignore; memorize NH₄⁺ as acidic since NH₃ is basic; or let NH₄⁺ dissociate:

\[ \text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq) \] and check if NH₃ is a strong acid or base; it is not \( \text{H}^+ \) formed \( \rightarrow \text{acidic} \} \)

g. basic \{this is analogous to NH₃ with 2 H’s replaced by \( -\text{CH}_2 \) groups; memorize NH₃ (an analogous compounds) are basic; or let (CH₃)₂NH react with H₂O:

\[ (\text{CH}_3)_2\text{NH}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons (\text{CH}_3)_2\text{NH}_2^+(aq) + \text{OH}^-(aq) \] and check if \( (\text{CH}_3)_2\text{NH}_2^+ \) is a strong acid or base; it is not \( \text{OH}^- \) formed \( \rightarrow \text{basic} \}

h. neutral \{both Na⁺ and Cl⁻ are spectator ions \( \rightarrow \text{neutral} \} \)

12. a. acidic \{NH₄F(aq) \rightarrow NH₄⁺(aq) + F⁻(aq); NH₄⁺ - acid; F⁻ - base; find K values and compare – whichever is bigger wins!

\[ K_b \text{ NH}_3 = 1.8 \times 10^{-5}; \quad K_a \text{ NH}_4^+ = (1.0 \times 10^{-14})/(1.8 \times 10^{-5}) = 5.56 \times 10^{-10}; \quad K_a \text{ HF} = 6.8 \times 10^{-4}; \]

\[ K_b \text{ F}^-(1.0 \times 10^{-14})/(6.8 \times 10^{-4}) = 1.47 \times 10^{-11}; \quad \text{since } K_a \text{ NH}_4^+ > K_b \text{ F}^- \rightarrow \text{acidic solution} \}

13. I. b \{to assist with the K’s:

\[ \text{H}_2\text{C}_4\text{H}_4\text{O}_4 \rightleftharpoons \frac{K_{a1}}{K_{b2}} \quad \text{H}_3\text{C}_4\text{H}_4\text{O}_4^- \rightleftharpoons \frac{K_{a2}}{K_{b1}} \quad \text{C}_4\text{H}_4\text{O}_4^{2-} \]

\{since the acid is a diprotic acid (2 H⁺), Kₑ₂ will correspond to the first base-reaction starting with the completely deprotonated acid, \text{C}_4\text{H}_4\text{O}_4^{-2} \text{ and Kₑ₂ will correspond to the next base-reaction starting from } \text{H}_3\text{C}_4\text{H}_4\text{O}_4^- \text{ reacting with } \text{H}_2\text{O} \text{ to yield } \text{OH}^- \text{ and } \text{H}_2\text{C}_4\text{H}_4\text{O}_4 \}

II. d \{the largest Kₐ for a polyprotic acid (not asked in this question) is Kₑ₁ and starts with the fully protonated acid, \text{H}_2\text{C}_4\text{H}_4\text{O}_4, \text{ reacting with } \text{H}_2\text{O} \text{ and yielding a } \text{H}_3\text{O}^+; \text{ likewise, the largest Kₐ for a polyprotic acid is Kₑ₂ and starts with the fully deprotonated acid, } \text{C}_4\text{H}_4\text{O}_4^{-2}, \text{ reacting with } \text{H}_2\text{O} \text{ and yielding a } \text{OH}^- \}
14. a. pH = 1.47 \{H_2C_3H_2O_4 + H_2O \rightleftharpoons HC_3H_2O_4^- + H_3O^+\}; write ICE Table:

<table>
<thead>
<tr>
<th></th>
<th>HC_3H_2O_4^- + H_2O \rightleftharpoons HC_3H_2O_4^- + H_3O^+</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.75</td>
</tr>
<tr>
<td>C</td>
<td>x</td>
</tr>
<tr>
<td>E</td>
<td>0.75-x</td>
</tr>
</tbody>
</table>

K_a1 = \frac{[HC_3H_2O_4^-][H_2O]}{[HC_2H_2O_4]} \times 1.5 \times 10^{-3} = \frac{x^2}{0.75 - x};

since [H_2C_3H_2O_4] > (100)1.5 \times 10^{-3}; approximation is valid; make approximation: 1.5 \times 10^{-3} = \frac{x^2}{0.75}; x = 0.03354 = [H_3O^+];

pH = -\log(0.03354) = 1.47; HC_3H_2O_4^- can lose a 2nd H^+: HC_3H_2O_4^- + H_2O \rightleftharpoons H_3O^+ + C_3H_2O_4^-; with polyprotic acids the 1st H^+ dominates (note that K_a1 >> K_a2) and the 2nd H^+ can be ignored because its amount is quite small!

b. \[C_3H_2O_4^-\] = 2.0 \times 10^{-6}M

\{to answer the next question write the next acid dissociation reaction: HC_3H_2O_4^- + H_2O \rightleftharpoons C_3H_2O_4^- + H_3O^+\}

K_a2 = \frac{[C_3H_2O_4^-][H_3O^+]}{[C_3H_2O_4^-]} \times 2.0 \times 10^{-6} = \frac{(x)(0.03354 + x)}{(0.03354 - x)}; since [C_3H_2O_4^-] > (100)2.0 \times 10^{-6} ok to make approximation: 2.0 \times 10^{-6} = \frac{(x)(0.03354)}{(0.03354)}; x = [C_3H_2O_4^-] = 2.0 \times 10^{-6};

Note: there is no significant change to [H_3O^+] after the second dissociation which is why the second dissociation of a polyprotic acid is ignored when determining the pH

15. HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq) \text{ acting as an acid};

HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq) \text{ acting as a base}

b. HCO_3^- will act as a base

\{to assist which K to use: H_2CO_3 \rightleftharpoons K_{a1}K_{b1}\}

acid reaction: K_{a2} = 4.8 \times 10^{-11}; base reaction: K_{b2} = K_w/K_{a1} = 1.0 \times 10^{-14}/4.2 \times 10^{-7} = 2.38 \times 10^{-8}; since K_{b2} > K_{a2} then HCO_3^- will act as a base

16. a. H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq) \text{ acting as an acid};

H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq) \text{ acting as a base}

b. H_2PO_4^- will act as an acid

\{to assist which K to use: H_2PO_4^- \rightleftharpoons K_{a1}K_{b1}\}

acid reaction: K_{a2} = 6.2 \times 10^{-8}; base reaction: K_{b3} = K_w/K_{a1} = 1.0 \times 10^{-14}/7.5 \times 10^{-3} = 1.33 \times 10^{-12}; since K_{a2} > K_{b3} then H_2PO_4^- will act as an acid

17. e \{metal cations are Lewis acids and can accept electron pairs - “a”, “b”, “c”; “d” (H^+) is an acid; PH_3 is similar to NH_3 (\(\text{H}^-\text{N}^-\text{H}\)) has 5 valence electrons) and therefore has a lone pair of electrons it can donate which makes it a Lewis base; \(\text{NH}_3\) \(\text{H}^-\text{N}^-\text{H}\) replaces a H_2O (a Lewis base) so it too must be donating a pair of electrons to Cr^{3+} and therefore is also acting as a Lewis base; with Bronsted acids and bases there must be an H^+ being transferred – there is no H^+ visible in this reaction so none of the chemicals are following the Bronsted definition of an acid or base!\}

18. e \{The interaction in this coordination compound is a Lewis acid-Lewis base interaction; the Cr^{3+} with its +3 charge is acting as a Lewis acid accepting lone pairs; the O in H_2O (\(\text{H}^-\text{O}\)) is donating a pair of electrons so it is acting as a Lewis base; the \(\text{NH}_3\) \(\text{H}^-\text{N}^-\text{H}\) replaces a H_2O (a Lewis base) so it too must be donating a pair of electrons to Cr^{3+} and therefore is also acting as a Lewis base; with Bronsted acids and bases there must be an H^+ being transferred – there is no H^+ visible in this reaction so none of the chemicals are following the Bronsted definition of an acid or base!\}