Acids and Bases-Part I

Do the topics appropriate for your course

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Nuggets: Definitions; Strong and Weak Acids/Bases; Conjugates; Amphiprotic; Autoionization; pH Calculations

Strong Acids/Bases; pH Calculations
Weak Acids/Bases; %ionization; Stronger/Weaker Acids/Bases

DEFINITIONS

Acids: taste sour; dissolve metals; react with carbonates; change indicator colors; neutralize bases

Bases: taste bitter; feel slippery; precipitate metal ions; change indicator colors; neutralize acids

Arrhenius: Acid increases \([H_3O^+]\) in water; base increases \([OH^-]\) in water

Bronsted-Lowry: Acid donates H\(^+\); base accepts H\(^+\)

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

IDENTIFYING ACIDS/BASES in Chem 104 – STRONG and WEAK

Acids: 1. Start with H (except for H\(_2\)O which can act as an acid or a base) or
2. Contains -COOH, carboxylic acid functional group; e.g., CH\(_3\)COOH \(\rightleftharpoons\) CH\(_3\)COO\(^-\) + H\(^+\)

Bases: 1. Contains OH\(^-\) (e.g., NaOH)
2. Many of the Chem 104 weak bases contain N with 3 groups (e.g., (CH\(_3\))\(_2\)NH\(_2\)) or other elements like N with a lone pair of e\(^-\)

(Note: The above guidelines cover many Bronsted-Lowry acids/bases used in Chem 104, but the guidelines are just that: guidelines that may not work for acids or bases covered in other classes!)

Strong acids/Strong bases dissociate/dissolve completely (100%); weak acids/bases dissociate only a little

Strong acids: HCl, HBr, HI, HNO\(_3\), H\(_2\)SO\(_4\), HClO\(_4\) (best to memorize)

Strong bases: Group IA hydroxides: LiOH, NaOH, KOH, RbOH, CsOH (best to memorize)
Group IIA hydroxides: Ca(OH)\(_2\), Sr(OH)\(_2\), and Ba(OH)\(_2\) (best to memorize)

Weak acids: Any acid that isn't strong (common: HF, CH\(_3\)COOH, H\(_3\)PO\(_4\), NH\(_4^+\), HCN, etc.);

Carboxylic acids: R-CO\(_2\)H; acetic acid: CH\(_3\)COOH (bold H is acidic; best to memorize acetic acid)

Weak bases: NH\(_3\) and other amine compounds; ammonia, NH\(_3\) (best to memorize ammonia)

CONJUGATE ACIDS AND BASES: the conjugate acids/bases are on the product side of a reaction

CH\(_3\)COOH + NH\(_3\) \(\rightleftharpoons\) CH\(_3\)COO\(^-\) + NH\(_4^+\)

add H\(^+\) \(\Rightarrow\) conjugate acid; remove H\(^+\) \(\Rightarrow\) conjugate base; e.g., HF (acid) \(\Rightarrow\) remove H\(^+\) \(\Rightarrow\) F\(^-\) (conj base)

AMPHIPROTIC: Can donate or accept H\(^+\); e.g., H\(_2\)O, HS\(^-\), H\(_2\)PO\(_4^-\), HPO\(_4^{2-}\), etc.;

Example: HS\(^-\)(aq) + OH\(^-\)(aq) \(\rightleftharpoons\) S\(^2-\)(aq) + H\(_2\)O(l) (HS\(^-\) acting as an acid; donating H\(^+\))
HS\(^-\)(aq) + H\(_2\)O(aq) \(\rightleftharpoons\) H\(_2\)S(g) + OH\(^-\)(aq) (HS\(^-\) acting as a base; accepting H\(^+\))
WATER – AUTOIONIZATION (self-ionizing)

\[
H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \quad \text{and} \quad K_w = [H_3O^+][OH^-]
\]

\[
H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)
\]

I \quad \text{-}x \quad \text{-}x \quad 0 \quad 0
C \quad -x \quad +x \quad +x
E \quad \text{-}x \quad x \quad x

K_w = [H_3O^+][OH^-]; K_w = 1.0 \times 10^{-14} at 25^\circ C; 1.0 \times 10^{-14} = (x)(x) = x^2; x = 1.0 \times 10^{-7} = [H_3O^+] = [OH^-] in pure water

(Sometimes this is presented as: \(H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)\) and \(K_w = [H^+][OH^-]\))

WRITING REACTIONS

Molecular Rxn: Compounds are left in molecular form

\[CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)\]

Complete Ionic Rxn: Strong acids, strong bases, and soluble ionic compounds are written as ions(aq);

\[CH_3COOH(aq) + Na^+(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + Na^+(aq) + H_2O(l)\]

Net Ionic Rxn: Ions that "react" are included; spectator ions are not included

\[CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)\]

pH EQUATIONS

\[pH = -\log[H_3O^+] \quad \text{and} \quad pOH = -\log[OH^-];\]

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

\[K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \quad \text{and} \quad pH + pOH = 14\]

\[\text{pH} < 7 \rightarrow \text{acidic } ([H_3O^+] > [OH^-])\]

\[\text{pH} > 7 \rightarrow \text{basic } ([OH^-] > [H_3O^+])\]

\[\text{pH} = 7 \rightarrow \text{neutral } ([OH^-] = [H_3O^+])\]

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

STRONG ACID pH Calculation

Example 1 a. What is the pH of a 0.0010M HBr solution? \(b. \text{What is the } pOH \text{ and } [OH^-]?\)

Answer 1 part a: Step 1. Write reaction.

\[\text{HBr} + H_2O \rightleftharpoons H_3O^+ + Br^-\]

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

Step 2. Write an ICE Table.

\[
\begin{array}{c|cccc}
 & \text{HBr} & H_2O & H_3O^+ & Br^- \\
\hline
I & 0.0010 & \text{---} & 0 & 0 \\
C & -0.0010 & -0.0010 & +0.0010 & +0.0010 \\
E & 0 & \text{---} & 0.0010 & 0.0010 \\
\end{array}
\]

(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. \(pH = -\log[H_3O^+]\)

\[pH = -\log(0.0010) = 3.00\]

Note: Since HBr is a strong acid, it is not necessary to write an ICE Table since all of the acid dissociates. (e.g., \(0.0010M \text{ HBr} \rightarrow 0.0010M \text{ H}_3\text{O}^+)\)

Answer 1 part b:

\[\text{pOH} = 14.00 - pH; \text{pOH} = 14.00 - 3.00 = 11.00\]

\[\text{[OH}^-\text{]} = 10^{-pOH}; \text{[OH}^-\text{]} = 10^{-11} = 1.00 \times 10^{-11}\]
IONIZATION CONSTANT EXPRESSION

WEAK ACIDS – $K_a$ (acid ionization constant)

$$\text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq) \quad K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$$ (always written with acid as a reactant)

WEAK BASES – $K_b$ (base ionization constant) for weak bases; when writing the reaction of a weak base

$$\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq) \quad K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$ (always written with base as a 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 below: If $100(K_a) < [\text{initial concentration}]_o →$ then approximation is valid.

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:** Step 1. Write reaction.

$$\text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq)$$

Note: Weak acid (recognized because it is not a strong acid which are memorized or because a $K_a$ was given)

→ dissociates a little

Step 4. Substitute values from ICE Table into $K_a$.

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

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

Check if $(100)K_a < [\text{HF}]_o$ is true;

$(100)6.8 \times 10^{-4} < 0.15$M $→$ true $→$

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

**Step 6. Solve for x.**

$x^2 = 0.000102; \ x = 0.0101 = [\text{H}_3\text{O}^+]$;

$pH = -\log[\text{H}_3\text{O}^+];$

$pH = -\log(0.0101) = 1.996; pH = 2.00$

**Example 3:** What is the pH of a 0.75M NH₃ solution? ($K_b = 1.8 \times 10^{-5}$)

**Answer 3:** Step 1. Write reaction.

$$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$$

Note: Weak base (recognized because it is not a strong base which are memorized or because a $K_b$ was given)

→ reaction proceeds a little

Step 4. Substitute values into $K_b$.

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

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

Check if $(100)K_b < [\text{NH}_3]_o$ is true;

$(100)1.8 \times 10^{-5} < 0.75$M $→$ true $→$

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

**Step 6. Solve for x.**

$x = 0.00367 = [\text{OH}^-];$ determine $pOH$

$pOH = -\log[\text{OH}^-] = -\log(0.00367) = 2.438$

$pH = 14.00 - 2.438 = 11.57$

PERCENT IONIZATION

$$\%\text{ionization} = \frac{\text{amount that reacts}}{\text{initial concentration}} \times 100\%; \ \%\text{ionization} = \frac{x}{[I]}_0 \times 100\% \quad ("x" \text{ from ICE table; } x = [\text{H}_3\text{O}^+] \text{ or } x = [\text{OH}^-])$$

- 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 NH₃ (Example 3 above)?

**Answer 4:** 0.49% \{ $\%\text{ionization} = \frac{[\text{OH}^-]}{[\text{base}]}_0 \times 100\%; \ [\text{OH}^-] = x = 0.00367M \text{ (from Example 3); } [\text{base}]_0 = 0.75M; \ \%\text{ionization} = \frac{0.00367}{0.75} \times 100\% = 0.489\% \}$

**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 $\%\text{ionization}$ increases\}

b. 0.5M HF \{for different acids with the same concentrations the stronger acid will have a greater percent ionization\}
**THE STRONGER THE ACID → THE WEAKER THE CONJUGATE BASE**

\[ K_a \uparrow \Rightarrow \text{acid strength} \uparrow \Rightarrow \text{pH} \downarrow ; \quad K_b \uparrow \Rightarrow \text{base strength} \uparrow \Rightarrow \text{pH} \uparrow \]

Example 6: Given 1.0M solutions of CH₃COOH (\(K_a = 1.8 \times 10^{-5}\)), HNO₂ (\(K_a = 4.5 \times 10^{-4}\)), and HF solution (\(K_a = 6.8 \times 10^{-4}\)), answer the following questions.

a. Which acid is strongest?
   *Answer 6 part a:* Strongest acid has largest \(K_a\). HF is strongest acid.

b. Which conjugate base is strongest?
   *Answer 6 part b:* Weakest acid will have the strongest conjugate base. CH₃COOH is the weakest acid since it has smallest \(K_a\) and therefore CH₃COO⁻ will be the strongest conjugate base.

c. Which solution will have the highest pH?
   *Answer 6 part c:* Highest pH means most basic solution which means weakest acid. CH₃COOH is weakest acid because it has the smallest \(K_a\) and therefore the highest pH.

**SIG FIGS (skip if not covered!)**

- pH, pOH, pKₐ, and pKₐ are logarithmic values. *Only the numbers past the decimal point are significant* (e.g., pH = 11.52 has 2 sig figs, the “52” are significant, not the “11”)
- [H₃O⁺], [OH⁻], \(K_a\), and \(K_b\) values are often given as scientific notation. All numbers of the significand or mantissa are significant (e.g., 2.304 x 10⁻² = 4 sig figs)
- *Calculation:* If the \([H_3O^+] = 1.25 \times 10^{-11}\) M (3 sigs), what is the pH? *Answer:* pH = 10.903 (3 sigs).

1. a. What is the conjugate base of OH⁻?
   b. What is the conjugate acid of HC₂O₄⁻?
   c. What is the conjugate base of HC₂O₄⁻?
   d. What is the conjugate base of HCOOH?
   e. What is the conjugate base of CH₃COOH?

2. Identify the acid, base, conjugate acid, and conjugate base in each reaction.
   I. CH₃COOH(aq) + H₂O(l) ⇌ H₃O⁺(aq) + CH₃COO⁻(aq)
   II. HF(aq) + NH₃(aq) ⇌ F⁻(aq) + NH₄⁺(aq)
3. Complete the reaction and identify the acid, base, conjugate acid, and conjugate base.
   \[(\text{CH}_3)_2\text{NH}(aq) + \text{C}_6\text{H}_5\text{COOH}(aq) \rightleftharpoons \]
   a. acid = \text{C}_6\text{H}_5\text{COOH}(aq), base = (\text{CH}_3)_2\text{NH}(aq), conjugate acid = (\text{CH}_3)_2\text{NH}_2^+(aq), conjugate base = \text{C}_6\text{H}_4\text{COO}^-(aq)
   b. acid = \text{C}_6\text{H}_5\text{COOH}(aq), base = (\text{CH}_3)_2\text{NH}(aq), conjugate acid = \text{C}_6\text{H}_5\text{COO}^-(aq), conjugate base = (\text{CH}_3)_2\text{NH}_2^+(aq)
   c. acid = (\text{CH}_3)_2\text{NH}(aq), base = \text{HCOOH}(aq), conjugate acid = \text{C}_6\text{H}_5\text{COOH}^+(aq), conjugate base = \text{C}_6\text{H}_4\text{COO}^-(aq)
   d. acid = \text{C}_6\text{H}_5\text{COOH}(aq), base = (\text{CH}_3)_2\text{NH}(aq), conjugate acid = (\text{CH}_3)_2\text{NH}_2^+(aq), conjugate base = \text{C}_6\text{H}_5\text{COO}^-(aq)
   e. acid = (\text{CH}_3)_2\text{NH}_2^+(aq), base = \text{C}_6\text{H}_5\text{COO}^-(aq), conjugate acid = \text{C}_6\text{H}_5\text{COOH}(aq), conjugate base = (\text{CH}_3)_2\text{NH}(aq)

4. Which of the four chemicals are amphiprotic? (Not a multiple choice question.)
   a. H$_2$O
   b. H$_3$PO$_4^-$
   c. CN$^-$
   d. HS$^-$

5. a. Write the autoionization reaction for water.
   b. Liquid ammonia, NH$_3$(l), can also undergo autoionization. Write the autoionization reaction for NH$_3$(l).

6. Calculate the pH, [H$_3$O$^+$], [OH$^-$], and pOH for each of the following.
   I. A 1.0 x 10$^{-3}$M solution of HCl.
   II. A solution contains 2.75g Ba(OH)$_2$(aq) in 2.50 liters of water.

7. 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

8. Complete the following reactions.
   a. HNO$_2$(aq) + NH$_3$(aq) →
   b. HF(aq) + LiOH(aq) →
   c. HCl(aq) + Ba(OH)$_2$(aq) →
   d. CH$_3$COOH(aq) + \text{C}_6\text{H}_5\text{NH}_2(aq) →

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

10. I. If 20. ml 0.050M HCl(aq) is mixed with 50. ml of a 0.075M HI(aq), what is the [H$_3$O$^+$] and pH of the solution?
    II. If 18.5ml 0.0200M HNO$_3$(aq) is mixed with 37.5ml of a 0.0250M HBr(aq), what is the [H$_3$O$^+$] and pH of the solution?
11. Use the following $K_a$’s to answer the questions below assuming all solutions are 1.0M.

<table>
<thead>
<tr>
<th>Acid</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HF</td>
<td>$6.8 \times 10^{-4}$</td>
</tr>
<tr>
<td>CH$_3$COOH</td>
<td>$1.8 \times 10^{-5}$</td>
</tr>
<tr>
<td>HOCl</td>
<td>$6.8 \times 10^{-8}$</td>
</tr>
<tr>
<td>CH$_2$(OH)COOH</td>
<td>$1.5 \times 10^{-4}$</td>
</tr>
</tbody>
</table>

I. Which solution is the strongest acid?
   a. HF  
   b. CH$_3$COOH  
   c. HOCl  
   d. CH$_2$(OH)COOH

II. Which solution would have the highest pH?
   a. HF  
   b. CH$_3$COOH  
   c. HOCl  
   d. CH$_2$(OH)COOH

III. Which solution will have the strongest conjugate base.
   a. HF  
   b. CH$_3$COOH  
   c. HOCl  
   d. CH$_2$(OH)COOH

IV. Which solution will have the greatest conductivity.
   a. HF  
   b. CH$_3$COOH  
   c. HOCl  
   d. CH$_2$(OH)COOH

V. Which solution will have the greatest percent ionization.
   a. HF  
   b. CH$_3$COOH  
   c. HOCl  
   d. CH$_2$(OH)COOH

12. Calculate the pH and percent ionization for each of the acid solutions.
I. 1.0M CH$_3$COOH ($K_a = 1.8 \times 10^{-5}$)
II. 0.10M HF ($K_a = 6.8 \times 10^{-4}$)
III. 0.50M HCN ($K_a = 4.0 \times 10^{-10}$).

13. If a 0.075M solution of a weak acid, HA(aq), has a percent ionization of 2.8%, what is the $K_a$ for this acid?

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

15. I. In a solution containing 0.1M HClO$_4$(aq), what is the order of concentrations of the following species: [H$_3$O$^+$], [OH$^-$], [ClO$_4^-$], and [HClO$_4$]? Include: “=”, “>”, or “negligible” when comparing the relative concentrations.
   II. In a solution containing 0.1M HNO$_2$(aq), what is the order of concentrations of the following species: [H$_3$O$^+$], [OH$^-$], [NO$_2^-$], and [HNO$_2$]? Include: “=”, “>”, or “negligible” when comparing the relative concentrations.

16. Calculate the pH and percent ionization for each of the base solutions.
I. 0.010M NH$_3$ ($K_b = 1.8 \times 10^{-5}$)
II. 0.50M CH$_3$NH$_2$ ($K_b = 5.0 \times 10^{-4}$)

17. 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?
18. Write the following chemicals in the order of decreasing acid strength:

HF (K_a = 6.8 x 10^{-4}), NH_4^+ (K_a = 5.6 x 10^{-6}), HCN (K_a = 4.0 x 10^{-10}), HSO_4^- (K_a = 1.2 x 10^{-2}), HCl, NaOH

**strongest acid** > **weakest acid**

19. Five solutions are prepared: a 1.5M HCOOH, 0.15M HCOOH, 1.5M HOCl, 0.15M HOCl, and 1.5M HNO_3.<br>
HCOOH (K_a = 3.0 x 10^{-4}) and HOCl (K_a = 6.8 x 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 HCOOH < 1.5M HOCl < 1.5M HCOOH < 1.5M HNO_3

**ANSWERS**

1. In these questions: remove H^+ to form conjugate base; add H^+ to form conjugate acid
a. O^2-  b. H_2C_2O_4  c. C_2O_4^{2-}  d. HCOO^- (not COOH^-)  e. CH_3COO^- (not CH_2COOH^-)

2. I. CH_3COOH, acid {it donates H^+}; H_2O, base {it accepts H^+}; H_3O^+, conjugate acid; CH_3COO^-, conjugate base
II. HF, acid {it donates H^+}; NH_3, base {it accepts H^+}; F^-, conjugate base; NH_4^+, conjugate acid

3. d {write reaction: (CH_3)_2NH(aq) + C_6H_5COOH(aq) ⇌ (CH_3)_2NH_2^+(aq) + C_6H_5COO^-(aq);
acid/base = reactants; conjugate acid/conjugate base = products; C_6H_5COOH is acid because of the –COOH group; (CH_3)_2NH is ammonia and is a base because of the N having 3 groups (–CH_3, –CH_3, and –H attached in a similar manner to NH_3); after acid gives up H^+ what remains behind is the conjugate base = C_6H_5COO^-; after the base accepts the H^+ what remains is an acid = (CH_3)_2NH_2^+}

4. I. a. conjugate acid = H_3O^+ {add H^+ to create conjugate acid}, conjugate base = OH^- {remove H^+ to create conjugate base}
b. conjugate acid = H_2PO_4^-, conjugate base = HPO_4^{2-}
c. conjugate acid = HCN, conjugate base = none
d. conjugate acid = H_2S, conjugate base = S^{2-}
II. H_2O, H_2PO_4^-, HS^- {the chemicals that can accept or donate an H^+ are amphiprotic or another way to think of this, a chemical that has both a conjugate acid and conjugate base}

5. a. H_2O(l) + H_2O(l) → H_3O^+(aq) + OH^-(aq)
b. NH_3(l) + NH_3(l) → NH_4^+(solvated) + NH_2^-(solvated)
6. I. pH = 3.00, [H\textsubscript{3}O\textsuperscript{+}] = 1.0 \times 10^{-3} \text{ M}, [\text{OH}^{-}] = 1.0 \times 10^{-11} \text{ M}, pOH = 11.00
   \{HCl + H_{2}O \rightarrow H_{3}O^{+} + Cl^{-} (100\%-SA); [H_{3}O^{+}] = 1.0 \times 10^{-3} \text{ M}; pH = -\log(1.0 \times 10^{-3}) = 3.00; pOH = 14.00 - 3 = 11.00; [\text{OH}^{-}] = 10^{-11} = 1.0 \times 10^{-11} \text{ M}\}
   II. pH = 12.107, [H_{3}O^{+}] = 7.81 \times 10^{-13} \text{ M}, [\text{OH}^{-}] = 0.0128 \text{ M}, pOH = 1.893
   \{2.75g \text{ Ba(OH)}_{2}(\text{aq}) \rightarrow \frac{1mol \text{ Ba(OH)}_{2}}{171.3g \text{ Ba(OH)}_{2}} \left(\frac{2mol \text{ OH}^{-}}{1mol \text{ Ba(OH)}_{2}}\right) = 0.0321mol \text{ OH}^{-}; [\text{OH}^{-}] = \frac{0.0321mol \text{ OH}^{-}}{2.5L} = 0.0128 \text{ M}
   pOH = -\log(0.0128) = 1.893; pH = 14.00 - 1.893 = 12.107; [H_{3}O^{+}] = 10^{-12.107} = 7.81 \times 10^{-13} \text{ M}\}
   III. pH = 1.893, [H_{3}O^{+}] = 7.81 \times 10^{-13} \text{ M}, [\text{OH}^{-}] = 0.0128 \text{ M}, pOH = 1.893
   \{1.0 \times 10^{-3} \text{ M} HCl; pH = 3; pOH = 11\}

7. I. a. write reaction: Sr(OH)\textsubscript{2} \rightarrow Sr^{+2} + 2OH\textsuperscript{-}; write ICE Table; Sr(OH)\textsubscript{2}(s) \rightarrow Sr^{+2}(aq) + 2OH\textsuperscript{-}(aq)
   \begin{array}{c|c|c}
   \text{Sr(OH)}_{2}(s) & \text{Sr}^{2+}(aq) & 2\text{OH}^{-}(aq) \\
   \hline
   I & 0.065 & 0 \\
   C & -0.065 & +0.065 +2(0.065) \\
   E & 0 & 0.065 & 0.13 \\
   \end{array}
   II. c. (dilution: M\textsubscript{1}V\textsubscript{1} = M\textsubscript{2}V\textsubscript{2}; M\textsubscript{1} = [OH\textsuperscript{-}] = 0.13M; V\textsubscript{1} = 100.0ml; V\textsubscript{2} = 500.0ml;
   (0.13M)(100.0ml) = M\textsubscript{2}(500.0ml); M\textsubscript{2} = [OH\textsuperscript{-}]\textsubscript{diluted} = 0.026M;
   pOH = -\log(0.026) = 1.59; pH = 12.41
   III. b. greatest acidity = smallest pH; “a”: pH = 2.65; “b”: pH = -\log(7.5 \times 10^{-3}) = 2.12;
   “c”: pOH = -\log(1.5 \times 10^{-12}) = 11.82; pH = 14.00 - 11.82 = 2.18; “d”: [H\textsubscript{3}O\textsuperscript{+}] = 0.0015M 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
   IV. a. HNO\textsubscript{2}(aq) + NH\textsubscript{3}(aq) \rightarrow NO\textsubscript{2}^{-}(aq) + NH\textsubscript{4}^{+}(aq)
   b. HF(aq) + LiOH(aq) \rightarrow LiF(aq) + H\textsubscript{2}O(l)
   c. HCl(aq) + Ba(OH)\textsubscript{2}(aq) \rightarrow BaCl\textsubscript{2}(aq) + 2H\textsubscript{2}O(l)
   d. CH\textsubscript{3}COOH(aq) + C\textsubscript{6}H\textsubscript{5}NH\textsubscript{2}(aq) \rightarrow CH\textsubscript{3}COO\textsuperscript{-}(aq) + C\textsubscript{6}H\textsubscript{5}NH\textsubscript{3}\textsuperscript{+}(aq)

9. b. greatest acidity = smallest pH; “a”: pH = 2.65; “b”: pH = -\log(7.5 \times 10^{-3}) = 2.12;
   “c”: pOH = -\log(1.5 \times 10^{-12}) = 11.82; pH = 14.00 - 11.82 = 2.18; “d”: [H\textsubscript{3}O\textsuperscript{+}] = 0.0015M 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

10. I. [H\textsubscript{3}O\textsuperscript{+}] = 0.068M; pH = 1.17 \{mol HCl = M\textsubscript{HCl}(0.020L) = 0.001M HCl; since HCl is a strong acid \rightarrow 0.001mol H\textsuperscript{+} = 0.001mol H\textsubscript{3}O\textsuperscript{+}; mol HI = M\textsubscript{HI}(0.050L) = 0.00375mol HI; since HI is a strong acid \rightarrow 0.00375mol H\textsuperscript{+} = 0.00375mol H\textsubscript{3}O\textsuperscript{+}; total volume = 20 + 50 = 70ml; \}
   \frac{0.00475mol \text{ H}_{3}\text{O}^{+}}{0.070L} = 0.0679M H\textsubscript{3}O\textsuperscript{+}; pH = -\log(0.0679) = 1.17\}
   II. [H\textsubscript{3}O\textsuperscript{+}] = 0.0234M; pH = 1.631 \{mol HNO\textsubscript{3} = M\textsubscript{HNO3}(0.0185L) = 0.000370mol HNO\textsubscript{3}; since HNO\textsubscript{3} is a strong acid \rightarrow 0.000370mol H\textsuperscript{+} = 0.000370mol H\textsubscript{3}O\textsuperscript{+}; mol HBr = M\textsubscript{HBr}(0.0375L) = 0.0009375mol HBr; since HBr is a strong acid \rightarrow 0.0009375mol H\textsuperscript{+} = 0.0009375mol H\textsubscript{3}O\textsuperscript{+}; total volume = 18.5 + 37.5 = 56.0ml; \}
   \frac{0.00131mol \text{ H}_{3}\text{O}^{+}}{0.056L} = 0.0234M H\textsubscript{3}O\textsuperscript{+}; pH = -\log(0.0234) = 1.631\}

11. I. a. {larger K\textsubscript{a} \rightarrow stronger acid}
   II. c. {highest pH means most basic or weakest acid; the weakest acid has the smallest K\textsubscript{a}}
   III. c. {the weakest acid has the strongest conjugate base; the weakest acid has the smallest K\textsubscript{a}}
   IV. a. {the stronger the acid the more dissociates and the greater the conductivity of the solution (i.e., greater amount of ions); since the concentrations are the same, the stronger acid will have a greater conductivity and the stronger acid has a larger K\textsubscript{a}}
   V. a. {the stronger the acid the more dissociates; since concentrations are the same (that’s key!), the stronger acid will have a greater ionization and the stronger acid has a larger K\textsubscript{a}}
12. I. pH = 2.37; %ionization = 0.42%  {CH$_3$COOH + H$_2$O $\rightleftharpoons$ H$_3$O$^+$ + CH$_3$COO$^-$}; write ICE Table;

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}; \quad 1.8 \times 10^{-5} = \frac{x(x)}{(1.0-x)}; \quad 1.8 \times 10^{-5} = \frac{x^2}{(1.0-x)};$$

(100)$K_a$ < [HA]$_0$; 100(1.8 x 10$^{-5}$) = 1.8 x 10$^{-3}$ < 1.0 →

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

$$\%\text{ionization} = \left( \frac{x}{[H_3O^+]} \right) \times 100\%; \quad \%\text{ionization} = \left( \frac{0.00424}{1.0} \right) \times 100\% = 0.424\% \}$$

II. pH = 2.08; %ionization = 8.3%  {HF + H$_2$O $\rightleftharpoons$ H$_3$O$^+$ + F$^-$}; write ICE Table;

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}; \quad 6.8 \times 10^{-4} = \frac{x(x)}{(0.1-x)}; \quad 6.8 \times 10^{-4} = \frac{x^2}{(0.1-x)}; \quad (100)K_a < [HA]$; 0$;$

100(6.8 x 10$^{-4}$) = 6.8 x 10$^{-2}$ < 0.10 → make approximation:

$$6.8 \times 10^{-4} = \frac{x^2}{0.1}; x = 0.00825 = [H_3O^+]; \) pH = -log(0.00825) = 2.08;

$$\%\text{ionization} = \left( \frac{x}{[H_3O^+]} \right) \times 100\%; \quad \%\text{ionization} = \left( \frac{0.00825}{0.1} \right) \times 100\% = 8.25\% \}$$

III. pH = 4.85; %ionization = 2.8 x 10$^{-3}$%  {HCN + H$_2$O $\rightleftharpoons$ H$_3$O$^+$ + CN$^-$};

write ICE Table; $K_a = \frac{[H_3O^+][CN^-]}{[HCN]}; \quad 4.0 \times 10^{-10} = \frac{x(x)}{(0.50-x)}; \quad 4.0 \times 10^{-10} = \frac{x^2}{(0.50-x)}; \quad (100)K_a < [HA]$; 0$;$

100(4.0 x 10$^{-10}$) = 4.0 x 10$^{-8}$ < 0.50 → make approximation:

$$4.0 \times 10^{-10} = \frac{x^2}{0.50}; x = 1.41 \times 10^{-5} = [H_3O^+]; \) pH = -log(1.41 x 10$^{-5}$) = 4.85;

$$\%\text{ionization} = \left( \frac{x}{[H_3O^+]} \right) \times 100\%; \quad \%\text{ionization} = \left( \frac{1.41 \times 10^{-5}}{0.50} \right) \times 100\% = 0.00282\% \}$$

13. 6.1 x 10$^{-5}$  {HA + H$_2$O $\rightleftharpoons$ H$_3$O$^+$ + A$^-$}; write ICE Table;

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}; \quad K_a = \frac{(x)(x)}{(0.075-x)}; \quad K_a = \frac{x^2}{(0.075-x)}; \quad \text{need “x”};$$

$$\%\text{ionization} = \left( \frac{x}{[H_3O^+]} \right) \times 100\%; \quad \text{amount that dissociates = x in an ICE Table};$$

$$2.8 = \left( \frac{x}{0.075} \right) \times 100\%; x = 0.0021; \quad K_a = \left( \frac{(0.0021)^2}{0.075-0.0021} \right) = 6.05 \times 10^{-5} \}$$

14. 3.6 x 10$^{-4}$  {HCNO + H$_2$O $\rightleftharpoons$ CNO$^- + H_3O^+$}; write ICE Table;

$$K_a = \frac{[H_3O^+][CNO^-]}{[HCNO]}; \quad K_a = \frac{(x)(x)}{(0.75-x)}; \quad K_a = \frac{x^2}{(0.75-x)}; \quad \text{now the pH was given = 1.79; [H}_3\text{O}^+) = 10^{-1.79} = 0.01620M = x; \quad K_a = \left( \frac{(0.01620)^2}{0.75-0.01620} \right) = 3.58 \times 10^{-4} \}$$
15. I. $[\text{H}_2\text{O}^+] = [\text{ClO}_4^-] > [\text{OH}^-]$, $[\text{HClO}_4]$ is negligible \{HClO$_4$ is a strong acid and therefore dissociates 100%; hence, there is no HClO$_4$ left at the end\} if every H$^+$ that comes off it forms $\text{H}_2\text{O}^+$ and a ClO$_4^-$ also was created in the same amount as the H$^+$; there is always some OH$^-$ present but since this is a strong acid not very much\}

II. $[\text{HNO}_2] > [\text{H}_2\text{O}^+] > [\text{NO}_2^-] > [\text{OH}^-]$ \{HNO$_2$ is a weak acid and therefore dissociates only a little; hence, most of the acid remains as HNO$_2$; a small amount dissociates and creates one NO$_2^-$ for every H$^+$; for every H$^+$ that comes off it forms $\text{H}_2\text{O}^+$; since the solution is acidic, the amount of $\text{H}_2\text{O}^+$ is greater than OH$^-$\}

16. I. pH = 10.62 $\{\text{NH}_3 + \text{H}_2\text{O} ~\rightleftharpoons ~\text{NH}_4^+ + \text{OH}^-\}$ write ICE Table;

\[
\begin{align*}
K_b &= \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}; K_b = \frac{(x)(x)}{(0.010 - x)}; \quad K_b = \frac{x^2}{(0.010 - x)}; (100)K_b < [\text{Base}]_o; \\
100(1.8 \times 10^{-5}) &= 1.8 \times 10^{-3} < 0.010 \rightarrow \text{make approximation:} \quad 1.8 \times 10^{-5} = \frac{x^2}{0.010}; \\
x &= 4.24 \times 10^{-4} = [\text{OH}^-]; \quad pOH = -\log(4.24 \times 10^{-4}) = 3.37; \\
pH = 14.00-pOH = 14.00-3.37 = 10.63
\end{align*}
\]

$\%_{\text{ion}} = 4.2\%$ \{\%_{\text{ionization}} = \left(\frac{x}{[\text{I}_o]}\right) \times 100\% ; \quad \text{\%}_{\text{ionization}} = \left(\frac{4.24 \times 10^{-4}}{0.010}\right) \times 100\% = 4.24\%\}

II. pH = 12.20 $\{\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} ~\rightleftharpoons ~\text{CH}_3\text{NH}_3^+ + \text{OH}^-\}$ write ICE Table;

\[
\begin{align*}
K_b &= \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}; K_b = \frac{(x)(x)}{(0.50 - x)}; K_b = \frac{x^2}{(0.50 - x)}; \\
(100)K_b < [\text{Base}]_o; 100(5.0 \times 10^{-4}) &= 5.0 \times 10^{-2} < 0.50 \rightarrow \text{make approximation:} \\
5.0 \times 10^{-4} &= \frac{x^2}{0.50}; \quad x = 1.58 \times 10^{-2} = [\text{OH}^-]; pOH = -\log(1.58 \times 10^{-2}) = 1.80; \\
pH = 14.00-pOH = 14.00-1.80 = 12.20
\end{align*}
\]

$\%_{\text{ion}} = 3.2\%$ \{\%_{\text{ionization}} = \left(\frac{x}{[\text{I}_o]}\right) \times 100\% ; \quad \text{\%}_{\text{ionization}} = \left(\frac{1.58 \times 10^{-2}}{0.50}\right) \times 100\% = 3.16\%\}

17. I. $7.4 \times 10^{-5}$ $\{\text{CH}_3\text{N}_3 + \text{H}_2\text{O} ~\rightleftharpoons ~\text{CH}_3\text{N}_3^+ + \text{OH}^-\}$ write ICE Table;

\[
\begin{align*}
K_b &= \frac{[\text{CH}_3\text{N}_3^+][\text{OH}^-]}{[\text{CH}_3\text{N}_3]}; K_b = \frac{(x)(x)}{(0.25 - x)}; K_b = \frac{x^2}{(0.25 - x)}; \\
\text{the pH was given} &= 11.63;\quad pOH = 14.00-pH = 14.00-11.63 = 2.37 \rightarrow \\
[\text{OH}^-] &= 10^{-2.37} = 0.00427M = x \text{ since I have assigned x to be [OH$^-]}; K_b = \frac{(0.00427)^2}{(0.25-0.00427)} = 7.42 \times 10^{-5}
\end{align*}
\]

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;

\[
\begin{align*}
K_b &= \frac{[\text{C}_6\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_6\text{H}_5\text{NH}_2]}; K_b = \frac{(x)(x)}{(1.15 - x)}; K_b = \frac{x^2}{(1.15 - x)}; \\
\text{the pH was given} &= 9.35;\quad pOH = 14.00-pH = 14.00-9.35 = 4.65 \rightarrow [\text{OH}^-] = 10^{-4.65} = 2.22 \times 10^{-5}M = x \text{ since} \\
\text{I have assigned x to be [OH$^-]}; K_b = \frac{(2.22 \times 10^{-5})^2}{(1.15-2.22 \times 10^{-5})} = 4.30 \times 10^{-10}
\end{align*}
\]
18. (strongest acid) HCl > HSO₄⁻ > HF > NH₄⁺ > HCN > NaOH (weakest acid)

19. d  {HNO₃ is a strong acid and therefore has %ion = 100% and the greatest %ion; the other solutions should be divided into the stronger weak acid, HCOOH, and the weaker weak acid, HOCl; at this point: HOCl < HCOOH < HNO₃; to distinguish between the 2 concentrations containing the same weak acid, recall as [HCOOH] ↓ → %ion ↑ and [HOCl] ↓ → %ion ↑; therefore: smallest %ion 1.5M HOCl < 0.15M HOCl < 1.5M HCOOH < 0.15M HCOOH < 1.5M HNO₃ greatest %ion}