CHEM 116
Acid-Base Models and Equilibrium

Lecture 18
Prof. Sevian

Note: We probably will not get all the way through these lecture slides at Lecture 18, but I want you to have them as a resource as you are working on the homework, so I have included them all.

Today’s agenda

- Acids & bases (ch. 16)
  - Arrhenius model
  - Bronsted-Lowry model
- Mathematical relationships
  - Converting between pH and [H⁺]
  - Using equilibrium constant for water (Kw) to calculate [OH⁻]
  - Relationship between K_a, K_b, and Kw
- Equilibrium in acid-base systems
  - Strong acid + strong base
  - Weak acid + strong base
  - Strong acid + weak base
  - Weak acid + weak base

- Announcements
  - Reminder: Re-take of problems section of Exam 2 will be on Thursday, Nov 13, at lecture
  - Next exam is Exam 3 on Tuesday, Dec 2 (first lecture after Thanksgiving break) – note date change from syllabus
  - Final exam has been scheduled for Tuesday, Dec 16, 3:00pm

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Some reactions that are considered acid-base reactions in the Arrhenius model of acids and bases

1. $\text{HCl (aq)} + \text{NaOH (aq)} \rightleftharpoons \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
2. $2 \text{HNO}_3 (aq) + \text{Ba(OH)}_2 (aq) \rightleftharpoons \text{Ba(NO}_3)_2 (aq) + 2 \text{H}_2\text{O (l)}$
3. $\text{CH}_3\text{COOH (aq)} + \text{KOH (aq)} \rightleftharpoons \text{KCH}_3\text{COO (aq)} + \text{H}_2\text{O (l)}$
4. $\text{H}_2\text{SO}_4 (aq) + \text{NaOH (aq)} \rightleftharpoons \text{NaHSO}_4 (aq) + \text{H}_2\text{O (l)}$
5. $\text{H}_2\text{C}_2\text{O}_4 (aq) + \text{NaOH (aq)} \rightleftharpoons \text{NaHC}_2\text{O}_4 (aq) + \text{H}_2\text{O (l)}$

This is the general Arrhenius acid-base reaction.

- What do all Arrhenius acids have in common? A removable H$^+$ in the acid
- What do all Arrhenius bases have in common? A removable OH$^-$ in the base

**General Arrhenius acid-base reaction**

$$\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$$
$$\text{B(OH)} \rightleftharpoons \text{B}^+ + \text{OH}^-$$
$$\text{HA} + \text{B(OH)} \rightleftharpoons \text{H}^+ + \text{A}^- + \text{B}^+ + \text{OH}^-$$

However, we also know that $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$ has

$$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$$

Therefore, nearly all the H$^+$ + OH$^-$ converts to H$_2$O, so

$$\text{HA} + \text{B(OH)} \rightleftharpoons (\text{B}^+)(\text{A}^-) + \text{H}_2\text{O}$$

**Acid + Base $\rightleftharpoons$ Salt + Water**
Some reactions that are considered acid-base reactions in the Bronsted-Lowry model of acids and bases

1. \( \text{H}_3\text{O}^+ (aq) + \text{NH}_3 (g) \rightarrow \text{NH}_4^+ (aq) + \text{H}_2\text{O} (l) \)
2. \( \text{H}_2\text{SO}_3 (aq) + \text{HS}^- (aq) \rightarrow \text{H}_2\text{S} (g) + \text{HSO}_3^- (aq) \)
3. \( \text{CH}_3\text{COOH} (aq) + \text{NH}_3 (aq) \rightarrow \text{NH}_4^+ (aq) + \text{CH}_3\text{COO}^- (aq) \)
4. \( \text{HNO}_2 (aq) + \text{HPO}_4^{2-} (aq) \rightarrow \text{H}_2\text{PO}_4^- (aq) + \text{NO}_2^- (aq) \)
5. \( \text{NH}_4^+ (aq) + \text{CN}^- (aq) \rightarrow \text{HCN} (aq) + \text{NH}_3 (g) \)

This is the general Bronsted-Lowry acid-base reaction.

- What do all Bronsted acids have in common?
- What do all Bronsted bases have in common?
Compare the two models of acids and bases

**Arrhenius definition**
- Acid is a substance that produces H\(^+\) when added to water
  - HCl, H\(_2\)SO\(_4\), HNO\(_3\)
  - CH\(_3\)COOH
- Base is a substance that produces OH\(^-\) when added to water
  - NaOH, Ba(OH)\(_2\)

**Bronsted-Lowry definition**
- Acid is a substance that donates protons (H\(^+\))
  - HCl, H\(_2\)SO\(_4\), HNO\(_3\)
  - CH\(_3\)COOH
  - H\(_3\)O\(^+\), NH\(_4^+\)
- Base is a substance that accepts protons (H\(^+\))
  - OH\(^-\)
  - NH\(_3\), H\(_2\)O

**Bronsted-Lowry conjugate acid-base pairs**

Pair #1: an acid on one side and its conjugate base on the other side
- Acid #1
  - Proton donor
- Base #1
  - After Acid #1 (H\(_3\)O\(^+\)) donated the proton

Pair #2: a base on one side and its conjugate acid on the other side
- Acid #2
  - Proton acceptor
- Base #2
  - After Base #2 (NH\(_3\)) accepted the proton
Recognizing reactions by their patterns

Arrhenius acid-base reaction

\[
\begin{align*}
\text{HCl (aq)} & + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)} \\
\text{H}_2\text{SO}_4 (aq) & + \text{Ba(OH)}_2 (aq) \rightarrow \text{BaSO}_4 (aq) + 2 \text{H}_2\text{O (l)} \\
\text{Acid} & + \text{Base} \rightarrow \text{Salt} + \text{Water}
\end{align*}
\]

Bronsted-Lowry acid-base reaction

\[
\begin{align*}
\text{H}_3\text{O}^+ (aq) & + \text{NH}_3 (aq) \rightarrow \text{H}_2\text{O (l)} + \text{NH}_4^+ (aq) \\
\text{CH}_3\text{COOH (aq)} & + \text{NH}_3 (aq) \rightarrow \text{CH}_3\text{COO}^- (aq) + \text{NH}_4^+ (aq) \\
\text{Acid #1} & + \text{Base #2} \rightarrow \text{Base #1} + \text{Acid #2}
\end{align*}
\]

Amphiprotic substances

- Some chemicals can be an acid or a base
- Water is the most important amphiprotic substance (water can be the proton donor and its own proton acceptor)
- Autoionization of water is a fast equilibrium reaction that always is occurring wherever there is water present

\[
\begin{align*}
\text{Net:} & \quad 2 \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^- \\
\text{Shorthand:} & \quad \text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-
\end{align*}
\]

- Equilibrium constant
  \[
  K_c = [\text{H}_3\text{O}^+] \cdot [\text{OH}^-] \quad \text{(or } K_c = [\text{H}^+] \cdot [\text{OH}^-] \text{ in shorthand)}
  \]
  \[
  = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}
  \]
You practice it

A Precursor to Practice Exercises on p. 673

Identify the acid on the left and its conjugate base on the right. Similarly, identify the base on the left and its conjugate acid on the right.

1. \( \text{HNO}_3 (aq) + \text{NH}_3 (g) \rightleftharpoons \text{NH}_4^+ (aq) + \text{NO}_3^- (aq) \)

2. \( \text{CN}^- (aq) + \text{HClO} (aq) \rightleftharpoons \text{HCN} (aq) + \text{ClO}^- (aq) \)

3. \( \text{HCO}_3^- (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{CO}_3^{2-} (aq) + \text{H}_3\text{O}^+ (aq) \)

A = (a) is an acid, (c) is its conjugate base
   (b) is a base, (d) is its conjugate acid

B = (b) is an acid, (c) is its conjugate base
   (a) is an acid, (d) is its conjugate base

C = (a) is a base, (c) is its conjugate acid
   (b) is a base, (d) is its conjugate acid

D = (b) is an acid, (d) is its conjugate base
   (a) is a base, (c) is its conjugate acid
Acid (or base) ionization constant

**$K_a$** (or $K_b$)

Acid ionization (with water acting as the base):
- General form: $HA + H_2O \rightleftharpoons A^- + H_3O^+$
- Examples
  - $HF + H_2O \rightleftharpoons F^- + H_3O^+$
  - $H_2S + H_2O \rightleftharpoons HS^- + H_3O^+$
  - $HSO_4^- + H_2O \rightleftharpoons SO_4^{2-} + H_3O^+$

Base ionization (with water acting as the acid):
- General form: $B^- + H_2O \rightleftharpoons HB^+ + OH^-$
- Examples
  - $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
  - $F^- + H_2O \rightleftharpoons HF + OH^-$
  - $HPO_4^{2-} + H_2O \rightleftharpoons H_2PO_4^- + OH^-$

Using the $K_a$ values to predict the position of a proton-transfer equilibrium

*Similar to Practice Exercise, p. 676*

Use the table of ionization constants for acids and bases to answer the following questions.

1. Which is the stronger acid, $H_2SO_4$ or $H_2SO_3$?
2. To which direction does this equilibrium lie?
   
   $H_2SO_4 + HSO_3^- \rightleftharpoons H_2SO_3^- + HSO_4^-$
3. Is benzoic acid ($C_6H_5COOH$) stronger or weaker than acetic acid?
4. Which has the stronger conjugate base, acetic acid or boric acid?
5. Which is the stronger base, ammonia or the acetate ion?
6. Which has the stronger conjugate acid, ammonia or the acetate ion?
Key points about acids and bases so far

- Arrhenius definition is acids have H⁺ while bases have OH⁻
- Bronsted-Lowry definition encompasses Arrhenius definition, plus more
- Bronsted-Lowry definition focuses on transfer of proton (H⁺) from acid to base
- Conjugate acid base pairs differ by an H⁺
- General acid-base reaction has acid1 becoming base1, while simultaneously base2 becomes acid2: in other words, acid1 gives an H⁺ to base1
- Acid ionization constant (\(K_a\)) is a measure of acid strength, when \(K_a\) is larger, the acid is stronger (it goes more to products, which include H⁺)
- The stronger the acid, the weaker its conjugate base, and vice versa

What is \(K_b\)? How is it related to \(K_a\)?

- General acid reaction
  \[
  \text{Acid}1 + \text{H}_2\text{O} \rightleftharpoons \text{Base}1 + \text{H}_3\text{O}^+ \quad K_a = \frac{[\text{Base}1][\text{H}_3\text{O}^+]}{[\text{Acid}1]}
  \]

- General base reaction involving the same pair
  \[
  \text{Base}1 + \text{H}_2\text{O} \rightleftharpoons \text{Acid}1 + \text{OH}^- \quad K_b = \frac{[\text{Acid}1][\text{OH}^-]}{[\text{Base}1]}
  \]

- These ought to be related to each other
  \[
  \text{Acid}1 + \text{H}_2\text{O} \rightleftharpoons \text{Base}1 + \text{H}_3\text{O}^+
  \]
  \[
  \text{Base}1 + \text{H}_2\text{O} \rightleftharpoons \text{Acid}1 + \text{OH}^-
  \]
  Sum: \[
  2 \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-
  \]

*Therefore, \(K_w = K_a K_b\)*

but this is just \(K_w = [\text{H}_3\text{O}^+][\text{OH}^-]\)

and \(K_w\) is always equal to \(1.0 \times 10^{-14}\)
What we measure in the laboratory is pH

- What is pH?
- How is it related to the $\text{H}_3\text{O}^+$ (aka, $\text{H}^+$) concentration?
- Why is it useful? What other things can you infer if you know the pH?

What is the pH scale?

- **Macro**
  - Provides information about whether a material (aqueous solution) is acid or base, and how strong it is

- **Particle**
  - Actually measures hydronium ion ($\text{H}^+$ or $\text{H}_3\text{O}^+$) concentration
    - Hydrogen ions ($\text{H}^+$) do not actually exist in solution
    - Instead, $\text{H}^+$ ions attach to water molecules and form $\text{H}_3\text{O}^+$ ions
    - $\text{H}_3\text{O}^+$ ions are called hydronium ions
    - Now that you know this, they are often abbreviated $\text{H}^+$

- **Symbol**
  - Acid strength increases to left
  - Base strength increases to right
  - pH 7 neutral

What is pH?

The letter "p" stands for “the negative logarithm base-10 of”

Some symbols that you will see include pH, pOH, p$K_a$, p$K_b$ and p$K_w$

\[
pH = -\log_{10}[H^+]
\]

\[
pOH = -\log_{10}[OH^-]
\]

\[
pK_a = -\log_{10}[K_a]
\]

\[
pK_b = -\log_{10}[K_b]
\]

\[
pK_w = -\log_{10}[K_w]
\]

Review of Powers of 10 Math

- All numbers can be written as powers of 10. Most important to remember where the decimal point is. (Key: is it a big or small number? If big then exponent is positive. If small, then exponent is negative.)

  \[
  \begin{align*}
  10,000 &= 10^4 \\
  100 &= 10^2 \\
  10 &= 10^1 \\
  1 &= 10^0 \\
  0.1 &= 10^{-1} \\
  0.001 &= 10^{-3} \\
  0.000001 &= 10^{-6}
  \end{align*}
  \]

- When multiplying two powers of 10, add the exponents. When dividing, subtract the exponents.

  \[
  10^5 \times 10^{-2} = 10^{5+(-2)} = 10^3
  \]

  \[
  10^{-3} / 10^1 = 10^{(-3)-1} = 10^{-4}
  \]
What does pH mean mathematically?

pH scale is a logarithmic scale

<table>
<thead>
<tr>
<th>pH scale</th>
<th>Acids</th>
<th>Neutral</th>
<th>Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$[\text{H}^+] &gt; [\text{OH}^-]$</td>
<td>$[\text{H}^+] = [\text{OH}^-]$</td>
<td>$[\text{H}^+] &lt; [\text{OH}^-]$</td>
</tr>
</tbody>
</table>

- **Scale A**
  - H^+ concentration in molarity (in exponential notation)
  - $10^0$, $10^{-1}$, $10^{-2}$, $10^{-3}$, $10^{-4}$, $10^{-5}$, $10^{-6}$, $10^{-7}$, $10^{-8}$, $10^{-9}$, $10^{-10}$, $10^{-11}$, $10^{-12}$, $10^{-13}$, $10^{-14}$

- **Scale B**
  - H^+ concentration in molarity (in decimal notation)
  - 0.1, 0.01, 0.001, etc.

- **Scale C**
  - a, b, c

from Sevian et al, Active Chemistry (2006)

There are only three type of acid-base problems

1. **Predict the pH**: given the amount of acid, base and/or salt added to some water, predict the pH of the solution.
   - How you approach calculating the pH depends on what you added to the water, so that’s the first thing you have to figure out
   - Variables are: strong or weak, acid or base
   - “Salts” in the Bronsted-Lowry scheme are actually acids or bases – their conjugates are more familiar to you

2. **Equilibrium**: given the measured pH of a solution, figure out how much acid, base or salt must have been added to some water to make the pH be that value.
   - These are always equilibrium problems

3. **Titration**: given a solution of unknown (acid or base) concentration, neutralize it with a known amount of (base or acid) to figure out the unknown concentration.
   - Involves stoichiometry since a neutralization reaction is occurring
   - If a weak acid or base is involved, it will also involve equilibrium calculations
What you need to be good at

Concepts
- Acid vs. base
- Figuring out conjugates
- Strong vs. weak
- Writing acid + water, and base + water, reactions and recognizing them as $K_a$ and $K_b$ reactions, respectively
- Applying strong/weak arguments to determine whether reactions go to completion

Calculation skills
- Molarity
- Comparing magnitudes of numbers to figure out when certain approximations will work
- Setting up an ICE table

Titration involving acid + base

You need an indicator to be able to tell when you reach the equivalence point
- Strong acid + strong base will have equivalence point of pH 7 (neutral)
- Strong acid + weak base will have equivalence point more acidic than pH 7 (i.e., pH < 7)
- Weak acid + strong base will have equivalence point more basic than pH 7 (i.e., pH > 7)

Basic idea: at equivalence point, moles of acid = moles of base
Adding strong acid or strong base to water

Not complicated. It just dissociates completely.

1 mole of HCl turns into 1 mole of H⁺ and 1 mole of Cl⁻

1 mole of H₂SO₄ turns into 1 mole of H⁺ and 1 mole of HSO₄⁻
  (note that HSO₄⁻ is a weak acid, so you cannot assume it dissociates completely)

1 mole of NaOH turns into 1 mole of Na⁺ and 1 mole of OH⁻

1 mole of Ba(OH)₂ turns into 1 mole of Ba²⁺ and 2 moles of OH⁻

So, to figure out [H⁺] or [OH⁻], calculate moles of it per liter of solution

Adding a strong acid to water

If you added 0.020 mol of HCl to 1.0 L of water, what would be the pH?

\[
[H^+] = 1.0 \times 10^{-7} M + 2.0 \times 10^{-2} M = 2.0 \times 10^{-2} M
\]

[OH⁻] can be calculated from \(K_w\) equation
A neat math trick to make it easier

- We know that $K_w$ always equals $1.0 \times 10^{-14}$
- We also know that $K_w = [H^+] [OH^-]$

\[
\begin{align*}
K_w &= [H^+] \cdot [OH^-] \\
1.0 \times 10^{-14} &= [H^+] \cdot [OH^-] \\
- \log(1.0 \times 10^{-14}) &= - \log([H^+] \cdot [OH^-]) \\
14 &= - \log([H^+]) + - \log([OH^-]) \\
14 &= \text{pH} + \text{pOH}
\end{align*}
\]

Example of strong base calculations

*Similar to Practice Exercises, pp. 683-684*

a) What is the pH of a 0.0012 M NaOH solution?

b) If the pH of a solution of the strong base Sr(OH)$_2$ is 10.46, what is the concentration of Sr(OH)$_2$ in mol/L?
Important distinction between strong and weak

- **Strong acids (and bases) dissociate completely**, so if you know the moles of acid (or base) you can determine the \([H^+]\) concentration (or \([OH^-]\))

- **Weak acids do not dissociate completely**, so you can’t figure out their \([H^+]\) concentration from knowing how much acid you added. Must use equilibrium calculation with \(K_a\)

- **Same idea for weak bases, but use** \(K_b\) **to get** \([OH^-]\)

Key points about acid-base equilibria

- **The difference between a strong acid and a weak acid** is that a strong acid dissociates completely into \(H^+\) and \(A^-\), while a weak acid dissociates only partially (and similar idea for bases)

- **The general acid reaction** has equilibrium constant \(K_a\) and is of the form \(HA + H_2O \rightleftharpoons A^- + H_3O^+\)

- **The general base reaction** has equilibrium constant \(K_b\) and is of the form \(B + H_2O \rightleftharpoons HB^+ + OH^-\)

- If you have an acid-base reaction, you can determine whether equilibrium lies to the left or right by comparing strengths of the acids via their \(K_a\) values

- **pH** is a logarithmic scale used for:
  - Reporting the \([H^+]\)
  - Making calculations involving \([H^+]\) simpler

- \(K_w = [H^+] [OH^-] = 1.0 \times 10^{-14}\)

- Because of this, \(pH + pOH = 14\)

- Also, \(K_w = K_a K_b\)
Acids and Bases in general:
What you (will) need to be able to do

- Identify conjugate acid-base pairs and predict reactions
- Equilibrium
- Titration
- Lewis acids and bases
- Buffers
  - Equations to use as shortcuts for solving problems

Strategies to master:
- Using the math tricks to solve problems
- Deciding on the right approach to solving a problem: recognizing acid-base equilibrium problems
- Recognizing hydrolysis reactions – “hydrolysis” is a fancy name for adding a weak acid or weak base to water (unfortunately referred to as a “salt” because it’s the conjugate that happens to be more familiar)

What happens when you add a weak acid (or base) to water?

It dissociates, but only partially, because \( K_a \) (or \( K_b \)) is small. In other words, not all the HA reacts.

\[
\begin{align*}
[H^+] &= 1.0 \times 10^{-7} M \\
[OH^-] &= 1.0 \times 10^{-7} M \\
[\text{pure water}] \\
\text{Acidic solution} \\
\text{Add 0.020 mol CH}_3\text{COOH} \\
\text{Effect is to add only a small amount of } H^+ \\
[\text{acid}]
\end{align*}
\]
How much does a weak acid or base dissociate?

You must use equilibrium calculations to solve this. You need to know $K_a$ to solve the problem.

**Example:** If you added 0.020 moles of CH$_3$COOH to water, what would be the pH? $K_a = 1.8 \times 10^{-5}$

\[
\text{CH}_3\text{COOH} (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{CH}_3\text{COO}^- (aq)
\]

<table>
<thead>
<tr>
<th>Initial</th>
<th>0.020 M</th>
<th>1.0$\times 10^{-7}$ M</th>
<th>0 M</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change</td>
<td>- $x$</td>
<td>+ $x$</td>
<td>+ $x$</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.020 - $x$</td>
<td>1.0$\times 10^{-7}$ + $x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5} = \frac{x(1.0 \times 10^{-7} + x)}{0.020 - x}
\]

Solving this exactly yields $x = 6.0 \times 10^{-4}$, which means pH = 3.22

There are two important approximations that can simplify the calculations…

**Approximations**

The amount of H$^+$ present in neutral water (1.0$\times 10^{-7}$ M) is negligible compared to the amount of H$^+$ contributed by adding the weak acid (or base) to water.

That is, the value of $x$ is much larger than 1.0$\times 10^{-7}$

The amount that the weak acid dissociates is negligible compared to the amount of acid that remains.

That is, the starting amount of acid is much greater than the value of $x$.
Comparing strong and weak acids

**Strong acid**
0.020 M HCl solution
- Acid dissociates completely
- \([\text{H}^+]\) is equal to \([\text{HCl}]\)
- \([\text{H}^+] = 0.020 \text{ M}\)
- \(\text{pH} = 1.70\)

**Weak acid**
0.020 M CH₃COOH solution
- Acid does not dissociate completely
- Need to know \(K_a\) to solve
- Must use equilibrium calculation to solve
- \([\text{H}^+] = \sqrt{C_d \cdot K_a}\)
- \([\text{H}^+] = 0.00060 \text{ M}\)
- \(\text{pH} = 3.22\)

How to recognize strong vs. weak acids

Memorize the strongest acids
- All halides except fluoride: HCl, HBr, HI
- Nitric acid: HNO₃
- Sulfuric acid (only the first H⁺): H₂SO₄
- Perchloric acid: HClO₄

Weak acids are listed in the \(K_a\) table
Adding a “salt” to water

- Is the salt a conjugate of a strong acid/base or of a weak acid/base?
- If it is a salt of a strong acid or base, then nothing will happen (like adding table salt to water – no change in pH).
- If it is a conjugate of a weak acid or base, then the “salt” is itself also a weak base or acid. So it hydrolyzes and makes some H⁺ or OH⁻, which changes the pH.

Acid-base properties of salt solutions: hydrolysis

When you add a salt to water, if it is soluble to any extent, it breaks apart into its constituent + and – ions. These ions can be weak acids or weak bases themselves. If they are, they “hydrolyze” to form either H⁺ or OH⁻, which changes the pH away from neutral pH 7 of the water.
Hydrolysis of a salt: comparing weak vs. strong

Salt of a strong acid
- What is the pH of a 0.020 M solution of NaBr?
- Is Na\(^+\) a conjugate of anything? No.
- Is Br\(^-\) a conjugate of anything? Yes. Of HBr.
- Is HBr strong or weak?
- HBr is a strong acid, so Br\(^-\) is a very weak base.

\[
\text{Br}^- + \text{H}_2\text{O} \rightleftharpoons \text{HBr} + \text{OH}^- \\
K_a \text{ for HBr is very large, so } K_b \text{ for Br}^- \text{ is very small.}
\]
- Equilibrium lies so strongly to the left that OH\(^-\) does not get produced in significant enough quantity to rival 1.0×10\(^{-7}\) M that exists in water.

Salt of a weak acid
- What is the pH of a 0.020 M solution of NaBrO?
- Is Na\(^+\) a conjugate of anything? No.
- Is BrO\(^-\) a conjugate of anything? Yes. Of HBrO.
- Is HBrO strong or weak?
- HBrO is a weak acid, so BrO\(^-\) is a weak base, but not very weak.

\[
\text{BrO}^- + \text{H}_2\text{O} \rightleftharpoons \text{HBrO} + \text{OH}^- \\
K_a \text{ for HBrO is } 2.5 \times 10^{-9}, \text{ so } K_b \text{ for BrO}^- \text{ is } 4.0 \times 10^{-6}.
\]
- Rxn occurs to enough extent that OH\(^-\) gets produced in significant enough quantity to make solution basic.

Hydrolysis example

*Exercise similar to 16.17, p. 701*
Which of the following salts, when added to water, would produce the most acidic solution?

a) KBr  
b) NH\(_4\)NO\(_3\)  
c) AlCl\(_3\)  
d) Na\(_2\)HPO\(_4\)