Today’s agenda

- Equilibrium in acid-base systems
  - Finish comparing strong vs. weak acids (and strong vs. weak bases)
- Buffers
  - When approximately equal amounts of HA and A⁻ are present in solution
- Titration
  - Predict what happens to pH as you add an acid to a base, or vice-versa

Announcements

- There is a graded extra credit assignment during discussion sections on Thursday Nov 20 and Tuesday Nov 25. You must attend discussion to do it.
- 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
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

Comparing strong and weak acids

<table>
<thead>
<tr>
<th>Strong acid</th>
<th>Weak acid</th>
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<tbody>
<tr>
<td>0.020 M HCl solution</td>
<td>0.020 M CH₃COOH solution</td>
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- Acid dissociates completely
  - [H⁺] is equal to [HCl]
  - [H⁺] = 0.020 M
  - pH = 1.70

- Acid does not dissociate completely
  - Need to know \( K_a \) to solve
  - Must use equilibrium calculation to solve
  - \([H^+] \approx \sqrt{C_{ac} \cdot K_a} = 0.00060 \text{ M}\)
  - 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

Acids and Bases in general:
What you (will) need to be able to do

- Identify conjugate acid-base pairs and predict reactions
- Equilibrium
- Titration
- 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)
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.

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

- \( K_a \) for HBr is very large, so \( K_b \) for Br⁻ is very small.
- Equilibrium lies so strongly to the left that OH⁻ does not get produced in significant enough quantity to rival 1.0 \( \times \) 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.

\[
\text{BrO}^- + \text{H}_2\text{O} \rightleftharpoons \text{HBrO} + \text{OH}^-
\]

- \( K_a \) for HBrO is 2.5 \( \times \) 10⁻⁹, so \( K_b \) for BrO⁻ 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) NH₄NO₃
- c) AlCl₃
- d) Na₂HPO₄
Acid-base equilibrium

**Acid**

HA or HB+

- Strong: HCl, HNO₃, HClO₄, H₂SO₄, a few others
- Weak

**Base**

A⁻ or B

- Strong: has an OH⁻ ion
- Weak

**Equilibrium**

- HA + H₂O ⇌ A⁻ + H₃O⁺
- B + H₂O ⇌ HB⁺ + OH⁻

**Calculations**

- To figure out [H⁺]:
  - Determine how many moles of strong acid dissolved in how many liters of water
  - Use quadratic formula or make approximations to get
    \[ [H^+] = \sqrt{C_A \cdot K_a} \]

- To figure out [OH⁻]:
  - Determine how many moles of strong base dissolved in how many liters of water
  - Use the equation
    \[ [OH^-] = \sqrt{C_B \cdot K_b} \]

**pH, pOH**

- pH + pOH = 14

**Buffer**

- If conjugate weak partners are present in approx equal molar quantities

**Conjugate Partner**

- HA or HB⁺

**Weak acids**

- Dissociates partially

**Strong acids**

- Dissociates completely