Chapter 14: Acids and Bases

Check MasteringChemistry Deadlines

Acids and Bases:
The sour taste of lemons and lime, the bite of sourdough bread, and the tang of a tomato are all caused by acids. Citric acid, acetic acid, and tartaric acid are examples. The slippery feel of soap and some household cleaning solutions, such as ammonia have the typical slippery feel of a base. Bases feel slippery because they react with oils on your skin to form soap-like substances.

Review Electrolytes in Aqueous Solution:
Strong Electrolytes will largely dissociate into its ions in an aqueous solution and are written as separated ions in the ionic reactions.

Examples: Strong acids, Strong Bases, Soluble salts.

Strong Acids: HCl, HBr, HI, HNO₃, H₂SO₄, HClO₄, HClO₃

Strong Bases: soluble hydroxides from Group 1A (not including H) and Group 11A, not including the top two) LiOH, NaOH, KOH, RbOH, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂

Soluble Salts: Ionic compounds that contain the cations from Group 1A; Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, or ammonium ion, NH₄⁺. A compound is probably soluble if it has the anion; Cl⁻¹, Br⁻¹, I⁻¹ except with Ag⁺, Hg₂⁺², or Pb⁺², and most compounds that include NO₃⁻¹, ClO₄⁻¹, C₂H₃O₂⁻¹, Soluble with most SO₄⁻² except Ba⁺², Hg₂⁺², or Pb⁺².

Weak Electrolytes partially dissociate into its ions in an aqueous solution, but are written as compounds in an ionic equation. Weak electrolytes are weak acids and weak bases such as HC₂H₃O₂ or NH₃
Acids:
- Sour taste
- Dissolve many metals
- Turn litmus red

**Hydrochloric Acid:**

Hydrochloric acid is the main component of stomach acid. Hydrochloric acid is found in most chemistry laboratories. It is used in industry to clean metals, to prepare and process foods, and to refine metal ores.

Hydrochloric acid helps break down food. It kills harmful bacteria that might enter the body through food. The sour taste associated with indigestion is caused by the stomach’s hydrochloric acid refluxing up into the esophagus.

**Sulfuric Acid:**

- Sulfuric acid is the most widely produced chemical in the United States; annual U.S. production of sulfuric acid exceeds 36 million tons.
- Sulfuric acid is used in the manufacture of fertilizers, explosives, dyes, and glue.
- Sulfuric acid is contained in most automobile batteries.

**Acetic Acid:**

- Acetic acid forms in improperly stored wines. The word vinegar originates from the French vin aigre, which means “sour wine.”
- Acetic acid is an example of a carboxylic acid, an acid containing the COOH grouping of atoms, known as the carboxylic acid group.
Bases:

- Bitter taste  
- Slippery feel  
- Turn litmus blue

Examples: NaOH, Ba(OH)$_2$, NH$_3$

Some common bases:

- Sodium hydroxide and potassium hydroxide are found in most chemistry laboratories. They are used in processing petroleum and cotton, and in soap and plastic manufacturing. NaOH is the active ingredient in products such as Drano that work to unclog drains.

- Sodium bicarbonate (NaHCO$_3$) can be found in most homes as baking soda and is an active ingredient in many antacids. When taken as an antacid, sodium bicarbonate neutralizes stomach acid, relieving heartburn and sour stomach.

- Bases are less common in foods than acids because of their bitter taste. Our aversion to the taste of bases is probably an adaptation to protect us against alkaloids, organic bases found in plants, which are often poisonous. The active component of hemlock is the alkaloid coniine. The bitter taste warns us against eating them.

- Coffee is acidic overall, but bases present in coffee—such as caffeine—impart a bitter flavor.

Definitions:

**Arrhenius Definition (1884):** Most limited, requires water.

**Acid:** Substance that will increase the H$^{+1}$ ion concentration in an aqueous solution.

\[ H^+ \text{ (acid or proton)} = H(H_2O)_n^{+1} = H_3O^+ \text{ (hydronium ion)} \]

**Base:** Substance that will increase the OH$^{-1}$ ion concentration in an aqueous solution.
**Bronsted Lowry Definitions** : In 1923, Johannes Brønsted, working in Denmark, and Thomas Lowry, working in England, developed the concept of proton transfer in acid–base behavior independently and simultaneously. This is a broader definition, more possibilities

**Acid:** Donates one \( H^{+} \) ion. \((HA), \text{NH}_{4}^{+} \)

**Base:** Accepts one \( H^{+} \) ion. \((A^{-}), \text{NH}_{3} \)

The Brønsted–Lowry definition works well with bases such as \( \text{NH}_{3} \) that do not inherently contain \( \text{OH}^{-} \) ions but still produce \( \text{OH}^{-} \) ions in solution.

**Conjugate Acid/Base Pairs:** Any two substances related to each other by the transfer of a proton can be considered a conjugate acid–base pair.

**Bronsted Lowry acid/base conjugates** are different by only a single \( H^{+} \).

The acid has one more \( H^{+} \) compared to the base in a conjugate pair

- \( \text{NH}_{4}^{+} \) is the conjugate acid for \( \text{NH}_{3} \) the conjugate base
- \( \text{HF} \) is the conjugate acid for \( \text{F}^{-} \) the conjugate base.

For the reaction...

\[
\text{NH}_{3}(aq) + \text{H}_{2}O(l) \rightleftharpoons \text{NH}_{4}^{+}(aq) + \text{OH}^{-}(aq)
\]

\( \text{NH}_{3} \) is a **Bronsted–Lowry base** because it accepts a proton

\( \text{H}_{2}O \) is a **Bronsted–Lowry acid** because it donates a proton

- \( \text{NH}_{4}^{+} \) is the conjugate acid of the \( \text{NH}_{3} \) base
- \( \text{OH}^{-} \) is the conjugate base of the \( \text{H}_{2}O \) acid

**Amphoteric** substances can act as an acid or a base.
Example 1:
   a) Write the formulas for the conjugate bases given the acids:
      \[ \text{NH}_4^+1, \text{HF}, \text{HNO}_2, \text{H}_2\text{SO}_3, \text{HSO}_3^{-1}, \text{H}_2\text{O} \]
   b) Write the formulas for the conjugate acids given the bases:
      \[ \text{C}_2\text{H}_3\text{O}_2^{-1}, \text{OH}^{-1}, \text{ClO}^{-1}, \text{SO}_3^{2-}, \text{HSO}_3^{-1}, \text{H}_2\text{O} \]
   c) Which of the species above are amphoteric?

Relative strengths:
A stronger acid will have a weaker conjugate base. Strong acids have a negligible (spectator ion) conjugate base. Stronger bases have weaker conjugate acids. Reactions always favor making more of the weaker acid or base in the reaction.

Example 2:
Which acid - HI (strong acid) or HF (weaker acid) - will have the weaker conjugate base?

Acid and Base Reactions:
Example 3: Complete and balance the reactions below...

Neutralization is the double displacement reaction of an acid with a base to form salt and water. \( \text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)} \)

\[ \text{H}_2\text{SO}_4 (aq) + \text{KOH} \rightarrow \_________________________ \]

Acid with metal is the combination of an acid with a metal to form salt and hydrogen gas, as long as the metal is more active than hydrogen.
\( 2\text{HCl (aq)} + \text{Zn (s)} \rightarrow \text{ZnCl}_2 (aq) + \text{H}_2 (g) \)
\( 2\text{HI (aq)} + \text{Mg (s)} \rightarrow \_________________________ \)
**Acid with metal oxides** is the combination of an acid with a metal oxide to form salt and water.

\[ 2\text{HCl (aq)} + \text{MgO (s)} \rightarrow \text{MgCl}_2 \text{(aq)} + \text{H}_2\text{O (l)} \]

\[ 2\text{HBr (aq)} + \text{ZnO (s)} \rightarrow \text{________________________} \]

**Acid Base Titrations:**

- **Titration** is a common lab procedure that applies solution stoichiometry. In titration, a substance in a solution of known concentration is reacted with another substance in a solution of unknown concentration. The end point is a visible change, such as color change of an indicator (phenolphthalein) that occurs near the stoichiometric equivalence point where neither reactant is present in excess, and both are limiting.

- After adding a few drops of indicator to the flask, slowly measure and add a solution of known concentration from a buret to the solution of unknown concentration in the flask.

- When you reach stoichiometric proportions (1 mol of \( \text{OH}^- \) for every 1 mol of \( \text{H}^+ \)), the indicator (phenolphthalein) changes to pink, signaling the equivalence point of the titration.

- Indicators are intense colored organic dyes that change color at different pH values. **Phenolphthalein** is a common indicator that is colorless in acidic solution and pink in basic solution.
**Water:**

Pure water ionizes just slightly. It is generally considered a nonelectrolyte since the amount it ionizes is tiny. Water is also considered amphoteric since it can act as an acid and gives up H\(^+\) and as a base and accepts an H\(^+\).

\[
\text{H}_2\text{O (l)} + \text{H}_2\text{O (l)} \rightarrow \text{H}_3\text{O}^{+1} (\text{aq}) + \text{OH}^{-1} (\text{aq})
\]

The hydronium ion (H\(_3\)O\(^{+1}\)) is often written as a proton in water, H\(^+1\) (aq), even though the H\(^+1\) is really chemically bonded to one or more water molecules in an aqueous solution connected by hydrogen bonding.

\[
\text{H(H}_2\text{O}_n)^{+1}
\]

**Ion product constant for water (K\(_w\))**

\[
\text{H}_2\text{O (l)} + \text{H}_2\text{O (l)} \rightarrow \text{H}_3\text{O}^{+1} (\text{aq}) + \text{OH}^{-1} (\text{aq})
\]

\[
K_w = [H^{+1}][OH^{-1}]; \text{ the numerical value of } K_w = 1.0 \times 10^{-14} \text{ at } 25°C
\]

[M], square parenthesis represents the substances concentration is measured in units of Molarity.

As the temperature changes, so will K\(_w\), only at 25°C is to be remembered.

- 0°C \[ K_w = 1.1 \times 10^{-15} \]
- **25°C** \[ K_w = 1.0 \times 10^{-14} \]
- 37°C \[ K_w = 2.5 \times 10^{-14} \]

**Example 4:** use \( K_w = [H^{+1}][OH^{-1}] = 1.0 \times 10^{-14} \) at 25°C

a) Solve for the Molarity of [H\(^+\)] and [OH\(^{-1}\)] in neutral water

b) Solve for the Molarity of [H\(^+\)] if [OH\(^{-1}\)] = 5.0 \times 10^{-4}M

c) Solve for the Molarity of [OH\(^{-1}\)] if [H\(^+\)] = 4.0 \times 10^{-3}M
**pH scale:**

pH is a value that helps in determining the acidity of a solution. pH can be determined using pH meter, pH paper, or indicators

**Formulas:**

\[ \text{K}_w = 1.0 \times 10^{-14} = [H^+] [OH^-] \]

\[ \text{pH} = -\log[H^+] \quad [H^+] = 10^{-\text{pH}} \]

\[ \text{pOH} = -\log[OH^-] \quad [OH^-] = 10^{-\text{pOH}} \]

\[ \text{pK}_w = -\log[\text{K}_w] \quad \text{pK}_w = \text{pH} + \text{pOH} = 14 \]

- Acid has a pH < 7
- Neutral has pH = 7
- Base has a pH > 7
Example 5:
Given that the pH of a solution is 4.88, Decide if the solution is neutral, acidic, or basic and solve for the $[H^+]$, $[OH^-]$, and pOH.

Buffered Solutions:
Buffered solutions resist change in pH even after a strong acid or a strong base is added to the solution. This is extremely important in many situations. Fish tanks, medicines, enzymes, biological applications all generally require a specific pH range to function properly. It is possible to buffer a solution at any pH. Buffers do not mean neutral solution?

Human Blood is buffered around a pH of 7.4, too high or too low will cause sickness and death. In healthy individuals, blood pH is between 7.36 and 7.40. If blood pH were to drop below 7.0 or rise above 7.8, death would result.

Buffers: A buffer contains significant amounts of a weak acid and its conjugate base. The weak acid consumes any added strong base, and the conjugate base consumes any added strong acid. In this way, a buffer resists pH change.
Composition of Effective Buffers:

a) *Weak acid and weak base conjugate pairs:* Buffers eliminate strong acids or strong bases when they are added. They do this because a reaction favors creating a weaker acid or weaker base. The weak conjugate base in the buffer accepts the $\text{H}^+$ that comes from an addition of a strong acid, this process forms its weak conjugate acid. Adding a strong base causes the weak conjugate acid in the buffer to donate its $\text{H}^+$, creating its weak conjugate base.

b) *Ratio of Conjugate acid base pairs should be within 1 to 10 or 10 to 1:* Generally, a buffer solution will have close to equivalent concentrations of each weak acid/weak base conjugate pair.

c) *Relatively high concentrations of each:* A buffer will generally have a minimum concentration of 0.05 M of each conjugate pair. The greater the concentrations, the better the buffer capacity (amount of acid or base it is able to take on before losing its buffering capabilities).

*Example 6:*
Identify which of the following pairs can function as a buffer. For those pairs that cannot function as a buffer, explain why not.

a) $\text{HNO}_3 ; \text{KNO}_3$

e) $\text{CH}_3\text{NH}_2 ; \text{CH}_3\text{NH}_3\text{Br}$

b) $\text{NH}_4\text{Cl} ; \text{NH}_3$

f) $\text{H}_2\text{CO}_3 ; \text{HCO}_3^{-1}$

c) $\text{NaC}_2\text{H}_3\text{O}_2 ; \text{HC}_2\text{H}_3\text{O}_2$

g) $\text{NaCl} ; \text{NaNO}_3$

d) $\text{HF} ; \text{HNO}_2$

h) $\text{KOH} ; \text{H}_2\text{O}$
Acid Rain:

All rain is a bit acidic due to naturally occurring CO₂ in the air…

\[ \text{H}_2\text{O} (l) + \text{CO}_2 (g) \rightarrow \text{H}_2\text{CO}_3 (aq) \text{ in the rain } \text{pH} \approx 5.6 \]

Burning fossil fuels, especially coal, cause rain to be more acidic due to sulfur and nitrogen impurities oxidizing. Legislation such as the Clean Air Act, limit amounts emitted into the air from smokestacks. Scrubbers absorb much of the pollution. ACID RAIN is the result of sulfur oxides and nitrogen oxides emitted by fossil fuel combustion mixing with water (moisture) in the air to form the acids like sulfuric acid and nitric acid, which fall as acid rain. The left photo was taken in 1935, the right 60 years later in New York’s Washington Square Park.

Example 7:

Predict the acid products…

4 \( \text{NO}_2 (g) + \text{O}_2 (g) + 2 \text{H}_2\text{O} (l) \rightarrow 4 \) ______ (an acid)

\( \text{SO}_2 (g) + \text{H}_2\text{O} (l) \rightarrow \) ______ (an acid)

2 \( \text{SO}_2 (g) + \text{O}_2 (g) + 2 \text{H}_2\text{O} (l) \rightarrow 2 \) ______ (an acid)