16.1 Acids and Bases: A Brief Review

- Acids taste sour and cause certain dyes to change color.
- Bases taste bitter and feel soapy.
- The Arrhenius concept of acids and bases:
  - An acid is a substance that, when dissolved in water, increases the concentration of \( H^+ \) ions.
    - Example: HCl is an acid.
  - An Arrhenius base is a substance that, when dissolved in water, increases the concentration of \( OH^- \) ions.
    - Example: NaOH is a base.
  - This definition is quite narrow in scope as it limits us to aqueous solutions.

16.2 Brønsted-Lowry Acids and Bases

- We can use a broader, more general definition for acids and bases that is based on the fact that acid-base reactions involve proton transfers.

The \( H^+ \) Ion in Water

- The \( H^+(aq) \) ion is simply a proton with no surrounding valence electrons.
- In water, clusters of hydrated \( H^+(aq) \) ions form.
- The simplest cluster is \( H_2O^-(aq) \).
  - We call this a hydronium ion.
  - Larger clusters are also possible (such as \( H_3O^+ \) and \( H_5O_2^+ \)).
- Generally we use \( H^+(aq) \) and \( H_3O^+(aq) \) interchangeably.

Proton-Transfer Reactions

- We will focus our attention on \( H^+(aq) \).
- According to the Arrhenius definitions, an acid increases \([H^+]\) and a base increases \([OH^-]\).
- Another definition of acids and bases was proposed by Brønsted and Lowry.
- In the Brønsted-Lowry system, a Brønsted-Lowry acid is a species that donates \( H^+ \) and a Brønsted-Lowry base is a species that accepts \( H^+ \).
  - Therefore a Brønsted-Lowry base does not need to contain \( OH^- \).
    - \( NH_3 \) is a Brønsted-Lowry base, but not an Arrhenius base.
- Consider \( NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq) \):
  - \( H_2O \) donates a proton to ammonia.
    - Therefore, water is acting as an acid.
  - \( NH_3 \) accepts a proton from water.
    - Therefore, ammonia is acting as a base.
  - An amphiprotic substance can behave either as an acid or as a base.
    - Thus, water is an example of an amphiprotic species.

Conjugate Acid-Base Pairs

- Whatever is left of the acid after the proton is donated is called its conjugate base.
- Similarly, a conjugate acid is formed by adding a proton to the base.
- Consider \( HX(aq) + H_2O(l) \rightarrow H_3O^+(aq) + X^-(aq) \):
  - \( HX \) and \( X^- \) differ only in the presence or absence of a proton.
    - They are said to be a conjugate acid-base pair.
  - \( X^- \) is called the conjugate base.
  - After \( HX \) (acid) loses its proton it is converted into \( X^- \) (base).
    - Therefore \( HX \) and \( X^- \) are a conjugate acid-base pair.
  - After \( H_2O \) (base) gains a proton it is converted into \( H_2O^+ \) (acid).
    - \( H_2O^+ \) is the conjugate acid.
    - Therefore, \( H_2O \) and \( H_2O^- \) are a conjugate acid-base pair.

Relative Strengths of Acids and Bases

- The stronger an acid is, the weaker its conjugate base will be.
- We can categorize acids and bases according to their behavior in water.
  1. Strong acids completely transfer their protons to water.
    - No undissociated molecules remain in solution.
    - Their conjugate bases have negligible tendencies to become protonated.
      - An example is HCl.
2. **Weak acids** only partially dissociate in aqueous solution.
   • They exist in solution as a mixture of molecules and component ions.
   • Their conjugate bases show a slight tendency to abstract protons from water.
   • These conjugate bases are weak bases.
     • Example: Acetic acid is a weak acid; acetate ion (conjugate base) is a weak base.

3. **Substances with negligible acidity** do not transfer a proton to water.
   • An example is CH₄.
   • In every acid-base reaction, the position of the equilibrium favors the transfer of a proton from the stronger acid to the stronger base.
     • H⁺ is the strongest acid that can exist in equilibrium in aqueous solution.
     • OH⁻ is the strongest base that can exist in equilibrium in aqueous solution.

16.3 The Autoionization of Water
• In pure water the following equilibrium is established:
  \[ 2\text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \]
• This process is called the **autoionization** of water.

The Ion Product of Water
• We can write an equilibrium constant expression for the autoionization of water.
• Because H₂O(l) is a pure liquid, we exclude it from the expression:
  \[ K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \]
  • \(K_w\) is called the **ion-product constant**.
    • At 25°C the ion-product of water is:
    \[ 1.0 \times 10^{-14} = K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \]
• This applies to pure water as well as to aqueous solutions.
  • A solution is **neutral** if \([\text{OH}^-] = [\text{H}_3\text{O}^+]\).
  • If the \([\text{H}_3\text{O}^+] > [\text{OH}^-]\), the solution is **acidic**.
  • If the \([\text{H}_3\text{O}^+] < [\text{OH}^-]\), the solution is **basic**.

16.4 The pH Scale
• In most solutions \([\text{H}^+]\) is quite small.
• We express the \([\text{H}^+]\) in terms of **pH**.
  \[ \text{pH} = -\log[\text{H}^+] = -\log[\text{H}_3\text{O}^+] \]
• Note that this is a logarithmic scale.
• Thus, a change in \([\text{H}^+]\) by a factor of 10 causes the pH to change by 1 unit.
• Most pH values fall between 0 and 14.
  • In neutral solutions at 25°C, pH = 7.00.
  • In acidic solutions, \([\text{H}^+] > 1.0 \times 10^{-7}\), so pH < 7.00.
    • As the pH decreases, the acidity of the solution increases.
  • In basic solutions, \([\text{H}^+] < 1.0 \times 10^{-7}\), so pH > 7.00.
    • As the pH increases, the basicity of the solution increases (acidity decreases).

Other “p” Scales
• We can use a similar system to describe the \([\text{OH}^-]\).
  \[ \text{pOH} = -\log[\text{OH}^-] \]
• Recall that the value of \(K_w\) at 25°C is 1.0 x 10^{-14}.
  • Thus, we can describe a relationship between pH and pOH:
    \[ -\log[\text{H}^+] + (-\log[\text{OH}^-]) = \text{pH} + \text{pOH} = -\log K_w = 14.00 \]

Measuring pH
• The most accurate method to measure pH is to use a pH meter.
  • However, certain dyes change color as pH changes.
    • They are called acid-base indicators.
  • Indicators are less precise than pH meters.
Many indicators do not have a sharp color change as a function of pH. Most acid-base indicators can exist as either an acid or a base. These two forms have different colors. The relative concentration of the two different forms is sensitive to the pH of the solution. Thus, if we know the pH at which the indicator turns color, we can use this color change to determine whether a solution has a higher or lower pH than this value. Some natural products can be used as indicators. (Tea is colorless in acid and brown in base; red cabbage extract is another natural indicator.)

16.5 Strong Acids and Bases

Strong Acids

The most common strong acids are HCl, HBr, HI, HNO₃, HClO₃, HClO₄, and H₂SO₄.

Strong acids are strong electrolytes. All strong acids ionize completely in solution. Example: Nitric acid ionizes completely in aqueous solution. 

\[
\text{HNO}_3(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{NO}_3^-(aq)
\]

Since H⁺ and H₃O⁺ are used interchangeably, we write 

\[
\text{HNO}_3(aq) \rightarrow \text{H}^+(aq) + \text{NO}_3^-(aq)
\]

In solution the strong acid is usually the only source of H⁺. Therefore, the pH of a solution of a monoprotic acid may usually be calculated directly from the initial molarity of the acid.

Caution: If the molarity of the acid is less than 10⁻⁶ M then the autoionization of water needs to be taken into account.

Strong Bases

The most common strong bases are ionic hydroxides of the alkali metals or the heavier alkaline earth metals (e.g., NaOH, KOH, and Ca(OH)₂ are all strong bases).

Strong bases are strong electrolytes and dissociate completely in solution.

For example: 

\[
\text{NaOH}(aq) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)
\]

The pOH (and thus the pH) of a strong base may be calculated using the initial molarity of the base.

Not all bases contain the OH⁻ ion. Ionic metal oxides, hydrides, and nitrides are basic.

The oxide, hydride and nitride ions are stronger bases than hydroxide.

They are thus able to abstract a proton from water and generate OH⁻. 

\[
\text{O}_2^-(aq) + \text{H}_2\text{O}(l) \rightarrow 2\text{OH}^-(aq)
\]

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

\[
\text{N}^3^-(aq) + 3\text{H}_2\text{O}(l) \rightarrow \text{NH}_3(aq) + 3\text{OH}^-(aq)
\]

16.6 Weak Acids

Weak acids are only partially ionized in aqueous solution.

There is a mixture of ions and un-ionized acid in solution.

Therefore, weak acids are in equilibrium: 

\[
\text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq)
\]

Or: 

\[
\text{HA}(aq) \rightleftharpoons \text{H}^+(aq) + \text{A}^-(aq)
\]

We can write an equilibrium constant expression for this dissociation: 

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \
\text{or } K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}
\]

\(K_a\) is called the acid-dissociation constant.

Note that the subscript “a” indicates that this is the equilibrium constant for the dissociation of an acid.

Note that [H₂O] is omitted from the \(K_a\) expression. (H₂O is a pure liquid.)

The larger the \(K_a\), the stronger the acid.

\(K_a\) is larger since there are more ions present at equilibrium relative to un-ionized molecules.

If \(K_a \gg 1\), then the acid is completely ionized and the acid is a strong acid.
16.7 Weak Bases

- In order to find the value of $K_a$, we need to know all of the equilibrium concentrations.
  - The pH gives the equilibrium concentration of $H^+$.
  - Thus, to find $K_a$ we use the pH to find the equilibrium concentration of $H^+$ and then use the stoichiometric
coefficients of the balanced equation to help us determine the equilibrium concentration of the other species.
  - We then substitute these equilibrium concentrations into the equilibrium constant expression and solve for $K_a$.

Using $K_a$ to Calculate pH

- Using $K_a$, we can calculate the concentration of $H^+$ (and hence the pH).
- Write the balanced chemical equation clearly showing the equilibrium.
- Write the equilibrium expression. Look up the value for $K_a$ (in a table).
- Write down the initial and equilibrium concentrations for everything except pure water.
  - We usually assume that the equilibrium concentration of $H^+$ is $x$.
  - However, if the $K_a$ value is quite small, we can make a simplifying assumption.
    - Assume that $x$ is negligible compared with the initial concentration of the acid.
    - This will simplify the calculation.
  - It is always necessary to check the validity of any assumption.
    - Once we have the value of $x$, check to see how large it is compared with the initial concentration.
    - If $x$ is $<5\%$ of the initial concentration, the assumption is probably a good one.
    - If $x>5\%$ of the initial concentration, then it may be best to solve the quadratic equation.
- Percent ionization relates the equilibrium $H^+$ concentration, $[H^+]_{\text{equilibrium}}$, to the initial HA concentration, $[HA]_{\text{initial}}$.
  - The higher the percent ionization is, the stronger the acid.
  - However, we need to keep in mind that percent ionization of a weak acid decreases as the molarity of the solution
increases.

Polyprotic Acids

- Polyprotic acids have more than one ionizable proton.
- The protons are removed in successive steps.
  - Consider the weak acid, H$_2$SO$_3$ (sulfurous acid):
    \[
    \text{H}_2\text{SO}_3(aq) \rightleftharpoons \text{H}^+(aq) + \text{HSO}_3^{-}(aq) \quad K_{a1} = 1.7 \times 10^{-2} \\
    \text{HSO}_3^{-}(aq) \rightleftharpoons \text{H}^+(aq) + \text{SO}_3^{2-}(aq) \quad K_{a2} = 6.4 \times 10^{-8}
    \]
  - where $K_{a1}$ is the dissociation constant for the first proton released, $K_{a2}$ is for the second, etc..
- It is always easier to remove the first proton than the second proton in a polyprotic acid.
  - Therefore, $K_{a1} > K_{a2} > K_{a3}$, etc..
- The majority of the $H^+$ (aq) at equilibrium usually comes from the first ionization (i.e., the $K_{a1}$ equilibrium).
  - If the successive $K_a$ values differ by a factor of $10^1$ or more, we can usually get a good approximation of the pH of a
solution of a polyprotic acid by considering the first ionization only.

16.7 Weak Bases

- Weak bases remove protons from substances.
- There is an equilibrium between the base and the resulting ions:
  \[
  \text{Weak base} + \text{H}_2\text{O}(l) \rightleftharpoons \text{conjugate acid} + \text{OH}^-(aq)
  \]
  - Example: $\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$.
  - The base-dissociation constant, $K_b$, is defined as
  \[
  K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}
  \]
  - The larger $K_b$, the stronger the base.
Types of Weak Bases

- Weak bases generally fall into one of two categories.
  - Neutral substances with a lone pair of electrons that can accept protons.
  - Most neutral weak bases contain nitrogen.
  - **Amines** are related to ammonia and have one or more N–H bonds replaced with N–C bonds (e.g., CH₃NH₂ is methylamine).
  - Anions of weak acids are also weak bases.
  - Example: ClO⁻ is the conjugate base of HClO (weak acid):
    \[ \text{ClO}^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HClO}(aq) + \text{OH}^- (aq) \]
    \[ K_b = 3.33 \times 10^{-7} \]

16.8 Relationship Between \( K_a \) and \( K_b \)

- We can quantify the relationship between the strength of an acid and the strength of its conjugate base.
  - Consider the following equilibria:
    \[ \text{NH}_4^+ (aq) \rightleftharpoons \text{NH}_3(aq) + \text{H}^+(aq) \]
    \[ \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^- (aq) \]
  - We can write equilibrium expressions for these reactions:
    \[ K_a = \frac{[\text{NH}_3][\text{H}^+]}{[\text{NH}_4^+]} \]
    \[ K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \]
  - If we add these equations together:
    \[ \text{NH}_4^+(aq) \rightleftharpoons \text{NH}_3(aq) + \text{H}^+(aq) \]
    \[ \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^- (aq) \]
  - The net reaction is the autoionization of water.
    \[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^- (aq) \]
  - Recall that:
    \[ K_w = [\text{H}^+][\text{OH}^-] \]
  - We use this information to write an expression that relates the values of \( K_a \), \( K_b \), and \( K_w \) for a conjugate acid-base pair:
    \[ K_a \times K_b = K_w \]
  - Alternatively, we can express this as:
    \[ pK_a + pK_b = pK_w = 14.00 \text{ (at 25}^\circ\text{C)} \]

- Thus, the larger \( K_a \) (and the smaller \( pK_a \)), the smaller \( K_b \) (and the larger \( pK_b \)).
  - The stronger the acid, the weaker its conjugate base and vice versa.

16.9 Acid-Base Properties of Salt Solutions

- Nearly all salts are strong electrolytes.
  - Therefore, salts in solution exist entirely of ions.
  - Acid-base properties of salts are a consequence of the reactions of their ions in solution.
  - Many salt ions can react with water to form OH⁻ or H⁺.
    - This process is called hydrolysis.

An Anion’s Ability to React with Water

- Consider an anion, \( X^- \), as the conjugate base of an acid.
  - Anions from weak acids are basic.
    - They will cause an increase in pH.
  - Anions from strong acids are neutral.
    - They do not cause a change in pH.
  - Anions with ionizable protons (e.g., HSO₄⁻) are amphiprotic.
    - They are capable of acting as an acid or a base.
      - If \( K_a > K_b \), the anion will tend to decrease the pH.
      - If \( K_b > K_a \), the anion will tend to increase the pH.

A Cation’s Ability to React with Water

- Polyatomic cations that have one or more ionizable protons are conjugate acids of weak bases.
  - They tend to decrease pH.
- Metal cations of Group 1A and heavy alkaline earth metals are cations of strong bases and do not alter pH.
- Other metal ions can cause a decrease in pH.
**Combined Effect of Cation and Anion in Solution**

- The pH of a solution may be qualitatively predicted using the following guidelines:
  - Salts derived from a strong acid and a strong base are neutral.
    - Examples are NaCl and Ca(NO$_3$)$_2$.
  - Salts derived from a strong base and a weak acid are basic.
    - Examples are NaClO and Ba(C$_2$H$_3$O$_2$)$_2$.
  - Salts derived from a weak base and a strong acid are acidic.
    - An example is NH$_4$Cl.
  - Salts derived from a weak acid and a weak base can be either acidic or basic.
    - Equilibrium rules apply!
    - We need to compare $K_a$ and $K_b$ for hydrolysis of the anion and the cation.
    - For example, consider NH$_4$CN.
      - Both ions undergo significant hydrolysis.
      - Is the salt solution acidic or basic?
      - The $K_a$ of NH$_4^+$ is smaller than the $K_b$ of CN$^-$, so the solution should be basic.

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**16.10 Acid-Base Behavior and Chemical Structure**

**Factors That Affect Acid Strength**

- Consider H–X.
- For this substance to be an acid:
  - The H–X bond must be polar with $H^+$ and $X^-$.
- In ionic hydrides, the bond polarity is reversed.
  - The H–X bond is polar with $H^-$ and $X^+$.
  - In this case, the substance is a base.
- Other important factors in determining acid strength include:
  - The strength of the bond.
    - The H–X bond must be weak enough to be broken.
  - The stability of the conjugate base, $X^-$.
    - The greater the stability of the conjugate base, the more acidic the molecule.

**Binary Acids**

- The H–X bond strength is important in determining relative acid strength in any group in the periodic table.
  - The H–X bond strength tends to decrease as the element X increases in size.
  - Acid strength increases down a group; base strength decreases down a group.
- H–X bond polarity is important in determining relative acid strength in any period of the periodic table.
  - Acid strength increases and base strength decreases from left to right across a period as the electronegativity of X increases.
- For example, consider the molecules HF and CH$_4$.
  - HF is a weak acid because the bond energy is high.
  - The electronegativity difference between C and H is so small that the C–H bond is nonpolar, and CH$_4$ is neither an acid nor a base.

**Oxyacids**

- Many acids contain one or more O–H bonds.
  - Acids that contain OH groups (and often additional oxygen atoms) bound to the central atom are called oxyacids.
    - All oxyacids have the general structure $Y$–O–H.
  - The strength of the acid depends on Y and the atoms attached to Y.
    - As the electronegativity of Y increases, so does the acidity of the substance.
    - The bond polarity increases and the stability of the conjugate base (usually an anion) increases.
- We can summarize how acid structure relates to the electronegativity of Y and the number of groups attached to Y:
  - For oxyacids with the same number of OH groups and the same number of oxygen atoms:
    - acid strength increases with increasing electronegativity of the central atom, Y.
    - Example: HClO > HBrO > HIO
  - For oxyacids with the same central atom, Y:
    - acid strength increases as the number of oxygen atoms attached to Y increases.
    - Example: HClO$_4$ > HClO$_3$ > HClO$_2$ > HClO
Carboxylic Acids
• There is a large class of acids that contain a –COOH group (a carboxyl group).
• Acids that contain this group are called carboxylic acids.
  • Examples are acetic acid, benzoic acid, and formic acid.
• Why are these molecules acidic?
  • The additional oxygen atom on the carboxyl group increases the polarity of the O–H bond and stabilizes the conjugate base.
  • The conjugate base (carboxylate anion) exhibits resonance.
    • This gives it the ability to delocalize the negative charge over the carboxylate group, further increasing the stability of the conjugate base.
• The acid strength also increases as the number of electronegative groups in the acid increases.
  • For example, acetic acid is much weaker than trichloroacetic acid.

16.11 Lewis Acids and Bases
• A Brønsted-Lowry acid is a proton donor.
• Focusing on electrons, a Brønsted-Lowry acid can be considered as an electron pair acceptor.
• Lewis proposed a new definition of acids and bases that emphasizes the shared electron pair.
  • A Lewis acid is an electron pair acceptor.
  • A Lewis base is an electron pair donor.
    • Note that Lewis acids and bases do not need to contain protons.
    • Therefore, the Lewis definition is the most general definition of acids and bases.
• What types of compounds can act as Lewis acids?
  • Lewis acids generally have an incomplete octet (e.g., BF$_3$).
  • Transition-metal ions are generally Lewis acids.
  • Lewis acids must have a vacant orbital (into which the electron pairs can be donated).
  • Compounds with multiple bonds can act as Lewis acids.
    • For example, consider the reaction:
      \[ \text{H}_2\text{O(l)} + \text{CO}_2(\text{g}) \rightarrow \text{H}_2\text{CO}_3(\text{aq}) \]
      • Water acts as the electron pair donor and carbon dioxide as the electron pair acceptor in this reaction.
      • Overall, the water (Lewis base) has donated a pair of electrons to the CO$_2$ (Lewis acid).

Hydrolysis of Metal Ions
• The Lewis concept may be used to explain the acidic properties of many metal ions.
• Metal ions are positively charged and attract water molecules (via the lone pairs on the oxygen atom of water).
  • This interaction is called hydration.
• Hydrated metal ions act as acids.
  • For example:
    \[ \text{Fe(H}_2\text{O)}_6^{2+}(\text{aq}) \rightleftharpoons \text{Fe(H}_2\text{O)}_5(\text{OH})^{2+}(\text{aq}) + \text{H}^+(\text{aq}) \quad K_a = 2 \times 10^{-3}. \]
  • In general:
    • the higher the charge is, the stronger the M–OH$_2$ interaction.
    • the smaller the metal ion is, the more acidic the ion.
      • Thus, the pH of an aqueous solution increases as the size of the ion increases (e.g., Ca$^{2+}$ vs. Zn$^{2+}$) and as the charge increases (e.g., Na$^+$ vs. Ca$^{2+}$ and Zn$^{2+}$ vs. Al$^{3+}$).

Homework
1. Read pages 669-678
   Pg 712 #1, 2, 15, 16, 17, 19, 21, 25, 27, 29, 31
2. Read pages 678-684
   Pg 714 #33, 35, 37, 39, 41, 43, 45, 47
3. Read pages 684-696
   Pg 712 #3, 4, 6, 53, 55, 57, 63, 65, 71, 75, 77
4. Read pages 696-702
   Pg 712 #5, 7, 79, 81, 83, 85, 87, 89
5. Read pages 702-712
   Pg 716 # 8, 9, 10, 91, 93, 95, 99, 101, 103
Chapter 16 Practice Test

1. The $K_a$ of hydrofluoric acid (HF) is $6.8 \times 10^{-4}$ at 25°C. What is the pH of a 0.35 M aqueous solution of HF?
   A) 3.2    B) 1.8    C) 3.6
   D) 0.46   E) 12

2. The acid-dissociation constants of sulfurous acid (H$_2$SO$_3$) are $K_{a1} = 1.7 \times 10^{-2}$ and $K_{a2} = 6.4 \times 10^{-8}$ at 25.0°C. Calculate the pH of a 0.163 M aqueous solution of H$_2$SO$_3$.
   A) 4.5    B) 1.4    C) 1.8
   D) 7.2    E) 1.3

3. What is the conjugate acid of NH$_3$?
   A) NH$_3$    B) NH$_2^+$    C) NH$_3^+$
   D) NH$_4^+$    E) NH$_4$OH

4. The conjugate base of HSO$_4^-$ is
   A) OH$^-$    B) H$_2$SO$_4$    C) SO$_4^{2-}$
   D) HSO$_4^-$    E) H$_3$SO$_4^+$

5. What is the pH of an aqueous solution at 25.0 °C in which [H$^+$] is 0.00250 M?
   A) 3.40    B) 2.60    C) -2.60
   D) -3.40   E) 2.25

6. What is the pH of an aqueous solution at 25.0 °C in which [OH$^-$] is 0.00250 M?
   A) +2.60   B) -2.60   C) +11.4
   D) -11.4   E) -2.25

7. What is the pH of an aqueous solution at 25.0 °C that contains $3.98 \times 10^{-9}$ M hydronium ion?
   A) 8.40    B) 5.60    C) 9.00
   D) 3.98    E) 7.00

8. What is the concentration (in M) of hydronium ions in a solution at 25.0 °C with pH = 4.282?
   A) 4.28    B) 9.71    C) $1.92 \times 10^{-10}$
   D) $5.22 \times 10^{-5}$    E) $1.66 \times 10^4$

9. What is the pOH of a 0.0150 M solution of barium hydroxide?
   A) 12.2    B) 12.5    C) 1.52
   D) 1.82    E) 10.4

10. The pH of a 0.55 M aqueous solution of hypobromous acid, HBrO, at 25.0 °C is 4.48. What is the value of $K_a$ for HBrO?
    A) $2.0 \times 10^{-9}$    B) $1.1 \times 10^{-9}$    C) $6.0 \times 10^{-5}$
    D) $3.3 \times 10^{-5}$    E) $3.0 \times 10^4$

11. Calculate the pH of a 0.500 M aqueous solution of NH$_3$. The $K_b$ of NH$_3$ is $1.77 \times 10^{-5}$.
    A) 8.95    B) 11.5    C) 2.52
    D) 5.05    E) 3.01

12. The acid-dissociation constants of phosphoric acid (H$_3$PO$_4$) are $K_{a1} = 7.5 \times 10^{-3}$, $K_{a2} = 6.2 \times 10^{-8}$, and $K_{a3} = 4.2 \times 10^{-13}$ at 25.0°C. What is the molar concentration of phosphate ion in a 2.5 M aqueous solution of phosphoric acid?
    A) $2.0 \times 10^{-19}$    B) $9.1 \times 10^{-5}$
    C) 0.13    D) $2.5 \times 10^{-5}$
    E) $8.2 \times 10^{-9}$

13. The acid-dissociation constant, $K_a$, for gallic acid is $4.57 \times 10^{-3}$. What is the base-dissociation constant, $K_b$, for the gallate ion?
    A) $4.57 \times 10^{-3}$    B) $2.19 \times 10^{-12}$
    C) $5.43 \times 10^{-5}$    D) $7.81 \times 10^{-6}$
    E) $2.19 \times 10^2$

14. $K_a$ for HF is $7.0 \times 10^{-4}$. $K_b$ for the fluoride ion is _____.
    A) $2.0 \times 10^{-8}$    B) $1.4 \times 10^{-11}$
    C) $7.0 \times 10^{-14}$    D) $7.0 \times 10^{-4}$
    E) $1.4 \times 10^3$

15. Determine the pH of a 0.15 M aqueous solution of KF. For hydrofluoric acid, $K_a = 7.0 \times 10^{-4}$.
    A) 12    B) 5.8    C) 8.2
    D) 2.3    E) 6.6

16. Calculate the pH of 0.726 M anilinium hydrochloride (C$_6$H$_5$NH$_3$Cl) solution in water, given that $K_b$ for aniline is $3.83 \times 10^{-4}$.
    A) 1.77    B) 12.2    C) 5.36
    D) 8.64    E) 12.4

17. The $K_a$ for formic acid (HCO$_2$H) is $1.8 \times 10^{-4}$. What is the pH of a 0.35 M aqueous solution of sodium formate (NaHCO$_2$)?
    A) 11    B) 5.4    C) 3.3
    D) 8.6    E) 4.2

18. Calculate the molarity of hydroxide ion in an aqueous solution that has a pOH of 5.33.
    A) $4.7 \times 10^{-6}$    B) 8.67    C) $2.1 \times 10^{-9}$
    D) $5.3 \times 10^{-14}$    E) $8.7 \times 10^{-14}$

19. According to the Arrhenius concept, an acid is a substance that _________.
    A) is capable of donating one or more H$^+$
    B) causes an increase in the concentration of H$^+$ in aqueous solutions
    C) can accept a pair of electrons to form a coordinate covalent bond
    D) reacts with the solvent to form the cation formed by autoionization of that solvent
    E) tastes bitter
20) A Brønsted-Lowry base is defined as a substance that _________.
A) increases [H⁺] when placed in H₂O
B) decreases [H⁺] when placed in H₂O
C) increases [OH⁻] when placed in H₂O
D) acts as a proton acceptor
E) acts as a proton donor

21) The molar concentration of hydronium ion in pure water at 25°C is
A) 0.001
B) 1.0 x 10⁻¹⁴
C) 1.0 x 10⁻¹⁰
D) 1.0
E) 7.00

22) The magnitude of Kₐ indicates that
A) water autoionizes very slowly
B) water autoionizes very quickly
C) water autoionizes only to a very small extent
D) the autoionization of water is exothermic

23) Of the following acids, ________ is not a strong acid.
A) HNO₂
B) H₂SO₄
C) HNO₃
D) HClO₄
E) HCl

24) Which one of the following is the weakest acid?
A) HF (Kₐ = 6.8 x 10⁻⁴)
B) HClO (Kₐ = 3.0 x 10⁻⁸)
C) HNO₂ (Kₐ = 4.5 x 10⁻⁴)
D) HCN (Kₐ = 4.9 x 10⁻¹⁰)
E) Acetic acid (Kₐ = 1.8 x 10⁻⁵)

25) In which of the following aqueous solutions does the weak acid exhibit the highest percentage ionization?
A) 0.01 M HC₂H₃O₂ (Kₐ = 1.8 x 10⁻⁵)
B) 0.01 M HNO₂ (Kₐ = 4.5 x 10⁻⁴)
C) 0.01 M HF (Kₐ = 6.8 x 10⁻⁴)
D) 0.01 M HClO (Kₐ = 3.0 x 10⁻⁸)
E) These will all exhibit the same percentage ionization.

26) Classify the following compounds as weak acids (W) or strong acids (S):
<table>
<thead>
<tr>
<th>Acids</th>
<th>Nitrous acid</th>
<th>Hydrochloric acid</th>
<th>Hydrofluoric acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>A) W W W</td>
<td>B) S S S</td>
<td>C) S W W</td>
<td></td>
</tr>
<tr>
<td>D) W S S</td>
<td>E) W S W</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

27) Classify the following compounds as weak acids (W) or strong acids (S):
<table>
<thead>
<tr>
<th>Acids</th>
<th>Hypochlorous acid</th>
<th>Perchloric acid</th>
<th>Chloric acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>A) W S S</td>
<td>B) S S S</td>
<td>C) S W W</td>
<td></td>
</tr>
<tr>
<td>D) W W W</td>
<td>E) W S W</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

28) Classify the following compounds as weak bases (W) or strong bases (S):
<table>
<thead>
<tr>
<th>Bases</th>
<th>Ammonia</th>
<th>Fluoride ion</th>
<th>Sodium hydroxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>A) W W S</td>
<td>B) S S S</td>
<td>C) S W W</td>
<td></td>
</tr>
<tr>
<td>D) W S S</td>
<td>E) W S W</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

29) Which solution has the highest [OH⁻]?
A) a solution with a pH of 3.0
B) a 1 x 10⁻⁴ M solution of HNO₃
C) a solution with a pOH of 12.0
D) pure water
E) a 1 x 10⁻³ M solution of NH₄Cl

30) A 0.0035 M aqueous solution of a particular compound has pH = 2.46. The compound is ________.
A) a weak base
B) a weak acid
C) a strong acid
D) a strong base
E) a salt

31) An aqueous solution of ________ is basic.
NH₄Cl Cu(NO₃)₂ K₂CO₃ NaF
A) NH₄Cl, Cu(NO₃)₂
B) K₂CO₃, NH₄Cl
C) NaF only
D) NaF, K₂CO₃
E) NH₄Cl only

32) Of the compounds below, a 0.1 M aqueous solution of ________ will have the highest pH.
A) KCN, Kₐ of HCN = 4.0 x 10⁻¹⁰
B) NH₄NO₃, Kₐ of NH₃ = 1.8 x 10⁻⁵
C) NaOAc, Kₐ of HOAc = 1.8 x 10⁻⁵
D) NaClO, Kₐ of HClO = 3.2 x 10⁻⁸
E) NaHS, Kₐ of HS⁻ = 1.8 x 10⁻⁷

33) A 0.1 M solution of ________ has a pH of 7.0.
A) Na₂S
B) KF
C) NaNO₃
D) NH₄Cl
E) NaF

34) Of the following, which is the strongest acid?
A) HIO
B) HIO₄
C) HIO₂
D) HIO₃
E) The acid strength of all of the above is the same.

35) Which of the following acids will be the strongest?
A) H₂SO₄
B) HSO₄⁻
C) H₂SO₃
D) H₂SeO₄
E) HSO₃⁻

36) A solution of acetic acid is 2.0% dissociated at 25.0 °C. What was the original concentration (in M) of the acetic acid solution? The Kₐ for acetic acid is 1.8 x 10⁻⁵.

37) What is the pH of a solution prepared by adding 0.820 grams of sodium acetate to 100.0 ml of water at 25.0 °C? The Kₐ at 25.0 °C for acetic acid is 1.8 x 10⁻⁵.

38) TRUE or FALSE
A Lewis acid is an electron-pair acceptor, and a Lewis base is an electron-pair donor.