

16.1 Acids and Bases: A Brief Review

- Acids taste sour and cause certain dyes to change color.
- Bases taste bitter and feel soapy.
- 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 $\text{H}^+(\text{aq})$ ion is simply a proton with no surrounding valence electrons.
- In water, clusters of hydrated $\text{H}^+(\text{aq})$ ions form.
- The simplest cluster is $\text{H}_3\text{O}^+(\text{aq})$.
 - We call this a **hydronium ion**.
 - Larger clusters are also possible (such as H_5O_2^+ and H_9O_4^+).
- Generally we use $\text{H}^+(\text{aq})$ and $\text{H}_3\text{O}^+(\text{aq})$ interchangeably.

Proton-Transfer Reactions

- We will focus our attention on $\text{H}^+(\text{aq})$.
- According to the Arrhenius definitions, an acid increases $[\text{H}^+]$ and a base increases $[\text{OH}^-]$.
- 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 $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$:
 - 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 $\text{HX}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{X}^-(\text{aq})$:
 - 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_3O^+ (acid).
 - H_3O^+ is the **conjugate acid**.
 - Therefore, H_2O and H_3O^+ are a conjugate acid-base pair.

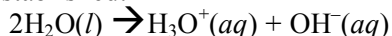
Relative Strengths of Acids and Bases

- 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:



- 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:

$$K_c = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

- 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 [OH⁻] = [H₃O⁺].
 - If the [H₃O⁺] > [OH⁻], the solution is *acidic*.
 - If the [H₃O⁺] < [OH⁻], the solution is *basic*.

16.4 The pH Scale

- In most solutions [H⁺] is quite small.
- We express the [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 [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, [H⁺] > 1.0 × 10⁻⁷, so pH < 7.00.
 - As the pH decreases, the acidity of the solution increases.
 - In basic solutions, [H⁺] < 1.0 × 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 [OH⁻].

$$\text{pOH} = -\log[\text{OH}^-]$$
- Recall that the value of K_w at 25°C is 1.0 × 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.

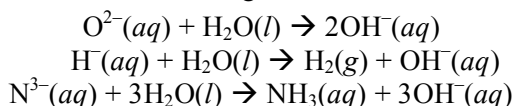
$$\text{HNO}_3(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{NO}_3^-(aq)$$
 - Since H⁺ and H₃O⁺ are used interchangeably, we write

$$\text{HNO}_3(aq) \rightarrow \text{H}^+(aq) + \text{NO}_3^-(aq)$$
- 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⁻.



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:

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

- 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 >> 1, then the acid is completely ionized and the acid is a strong acid.

Calculating K_a from pH

- 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 .
- Substitute into the equilibrium constant expression and solve.
 - Remember to convert x to pH if necessary.
- What do we do if we are faced with having to solve a quadratic equation in order to determine the value of x ?
 - Often this cannot be avoided.
 - However, if the K_a value is quite small, we find that 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.
- Weak acids are only partially ionized.
- Percent ionization is another method used to assess acid strength.
- For the reaction, $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

$$\% \text{ ionization} = \frac{[H^+]_{\text{equilibrium}}}{[HA]_{\text{initial}}} \times 100$$

- 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_2SO_3 (sulfurous acid):
$$H_2SO_3(aq) \rightleftharpoons H^+(aq) + HSO_3^-(aq) \quad K_{a1} = 1.7 \times 10^{-2}$$
$$HSO_3^-(aq) \rightleftharpoons H^+(aq) + SO_3^{2-}(aq) \quad K_{a2} = 6.4 \times 10^{-8}$$
 - 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 10^3 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

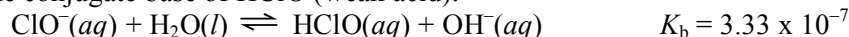
- Weak bases remove protons from substances.
- There is an equilibrium between the base and the resulting ions:
$$\text{Weak base} + H_2O(l) \rightleftharpoons \text{conjugate acid} + OH^-(aq)$$
- Example: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.
- The **base-dissociation constant**, K_b , is defined as

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

- The larger K_b , the stronger the base.

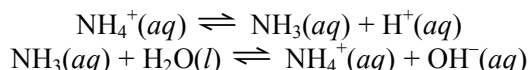
Types of Weak Bases

- 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., CH_3NH_2 is methylamine).
 - Anions of weak acids are also weak bases.
- Example: ClO^- is the conjugate base of HClO (weak acid):



16.8 Relationship Between K_a and K_b

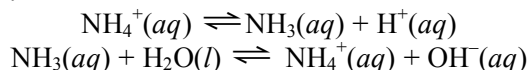
- We can quantify the relationship between the strength of an acid and the strength of its conjugate base.
- Consider the following equilibria:



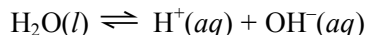
- We can write equilibrium expressions for these reactions:

$$K_a = \frac{[\text{NH}_3][\text{H}^+]}{[\text{NH}_4^+]} \quad K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

- If we add these equations together:



- The net reaction is the autoionization of water.



- 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:

$$\text{p}K_a + \text{p}K_b = \text{p}K_w = 14.00 \text{ (at } 25^\circ\text{C)}$$

- Thus, the larger K_a (and the smaller $\text{p}K_a$), the smaller K_b (and the larger $\text{p}K_b$).
 - 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., HSO_4^-) 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 $\text{Ca}(\text{NO}_3)_2$.
 - Salts derived from a strong base and a weak acid are basic.
 - Examples are NaClO and $\text{Ba}(\text{C}_2\text{H}_3\text{O}_2)_2$.
 - Salts derived from a weak base and a strong acid are acidic.
 - An example is NH_4Cl .
 - 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_4CN .
 - 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.

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 $\text{H}^{\delta+}$ and $\text{X}^{\delta-}$.
- In ionic hydrides, the bond polarity is reversed.
 - The H–X bond is polar with $\text{H}^{\delta-}$ and $\text{X}^{\delta+}$.
 - 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: $\text{HClO} > \text{HBrO} > \text{HIO}$
 - For oxyacids with the same central atom, Y:
 - acid strength increases as the number of oxygen atoms attached to Y increases.
 - Example: $\text{HClO}_4 > \text{HClO}_3 > \text{HClO}_2 > \text{HClO}$

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:
$$\text{H}_2\text{O}(l) + \text{CO}_2(g) \rightarrow \text{H}_2\text{CO}_3(aq)$$
 - 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:
$$\text{Fe}(\text{H}_2\text{O})_6^{3+}(aq) \rightleftharpoons \text{Fe}(\text{H}_2\text{O})_5(\text{OH})^{2+}(aq) + \text{H}^+(aq) \quad K_a = 2 \times 10^{-3}$$
- 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 6.8×10^{-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_2SO_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_2SO_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_4OH

4) The conjugate base of HSO_4^- is

- A) OH^- B) H_2SO_4 C) SO_4^{2-}
D) HSO_4^+ E) H_3SO_4^+

5) What is the pH of an aqueous solution at 25.0°C in which $[\text{H}^+]$ 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 $[\text{OH}^-]$ 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 3.98×10^{-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) 1.92×10^{-10}
D) 5.22×10^{-5} E) 1.66×10^4

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) 2.0×10^{-9} B) 1.1×10^{-9} C) 6.0×10^{-5}
D) 3.3×10^{-5} E) 3.0×10^4

11) Calculate the pH of a 0.500 M aqueous solution of NH_3 .

The K_b of NH_3 is 1.77×10^{-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_3PO_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) 2.0×10^{-19} B) 9.1×10^{-5}
C) 0.13 D) 2.5×10^{-5}
E) 8.2×10^{-9}

13) The acid-dissociation constant, K_a , for gallic acid is

4.57×10^{-3} . What is the base-dissociation constant, K_b , for the gallate ion?

- A) 4.57×10^{-3} B) 2.19×10^{-12}
C) 5.43×10^{-5} D) 7.81×10^{-6}
E) 2.19×10^2

14) K_a for HF is 7.0×10^{-4} . K_b for the fluoride ion is ____.

- A) 2.0×10^{-8} B) 1.4×10^{-11}
C) 7.0×10^{-18} D) 7.0×10^{-4}
E) 1.4×10^3

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 ($\text{C}_6\text{H}_5\text{NH}_3\text{Cl}$) solution in water, given that K_b for aniline is 3.83×10^{-4} .

- A) 1.77 B) 12.2 C) 5.36
D) 8.64 E) 12.4

17) The K_a for formic acid (HCO_2H) is 1.8×10^{-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) 4.7×10^{-6} B) 8.67 C) 2.1×10^{-9}
D) 5.3×10^{-14} E) 8.7×10^{-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 H_2O
- B) decreases $[H^+]$ when placed in H_2O
- C) increases $[OH^-]$ when placed in H_2O
- D) acts as a proton acceptor
- E) acts as a proton donor

21) The molar concentration of hydronium ion in pure water at $25^\circ C$ is

- A) 0.00
- B) 1.0×10^{-7}
- C) 1.0×10^{-14}
- D) 1.00
- E) 7.00

22) The magnitude of K_w 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) HNO_2
- B) H_2SO_4
- C) HNO_3
- D) $HClO_4$
- E) HCl

24) Which one of the following is the weakest acid?

- A) HF ($K_a = 6.8 \times 10^{-4}$)
- B) $HClO$ ($K_a = 3.0 \times 10^{-8}$)
- C) HNO_2 ($K_a = 4.5 \times 10^{-4}$)
- D) HCN ($K_a = 4.9 \times 10^{-10}$)
- E) Acetic acid ($K_a = 1.8 \times 10^{-5}$)

25) In which of the following aqueous solutions does the weak acid exhibit the highest percentage ionization?

- A) 0.01 M $HC_2H_3O_2$ ($K_a = 1.8 \times 10^{-5}$)
- B) 0.01 M HNO_2 ($K_a = 4.5 \times 10^{-4}$)
- C) 0.01 M HF ($K_a = 6.8 \times 10^{-4}$)
- D) 0.01 M $HClO$ ($K_a = 3.0 \times 10^{-8}$)
- E) These will all exhibit the same percentage ionization.

26) Classify the following compounds as weak acids (W) or strong acids (S):

- | | | |
|--------------|-------------------|-------------------|
| nitrous acid | hydrochloric acid | hydrofluoric acid |
| A) W W W | B) S S S | C) S W W |
| D) W S S | E) W S W | |

27) Classify the following compounds as weak acids (W) or strong acids (S):

- | | | |
|-------------------|-----------------|--------------|
| hypochlorous acid | perchloric acid | chloric acid |
| A) W S S | B) S S S | C) S W W |
| D) W W W | E) W S W | |

28) Classify the following compounds as weak bases (W) or strong bases (S):

- | | | |
|----------|--------------|------------------|
| ammonia | fluoride ion | sodium hydroxide |
| A) W W S | B) S S S | C) S W W |
| D) W S S | E) W S W | |

29) Which solution has the highest $[OH^-]$?

- A) a solution with a pH of 3.0
- B) a 1×10^{-4} M solution of HNO_3
- C) a solution with a pOH of 12.0
- D) pure water
- E) a 1×10^{-3} M solution of NH_4Cl

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.

- | | | | |
|----------------------------|-------------------------|-----------|-------|
| NH_4Cl | $Cu(NO_3)_2$ | K_2CO_3 | NaF |
| A) NH_4Cl , $Cu(NO_3)_2$ | B) K_2CO_3 , NH_4Cl | | |
| C) NaF only | D) NaF , K_2CO_3 | | |
| E) NH_4Cl only | | | |

32) Of the compounds below, a 0.1 M aqueous solution of _____ will have the highest pH.

- A) KCN , K_a of $HCN = 4.0 \times 10^{-10}$
- B) NH_4NO_3 , K_b of $NH_3 = 1.8 \times 10^{-5}$
- C) $NaOAc$, K_a of $HOAc = 1.8 \times 10^{-5}$
- D) $NaClO$, K_a of $HClO = 3.2 \times 10^{-8}$
- E) $NaHS$, K_b of $HS^- = 1.8 \times 10^{-7}$

33) A 0.1 M solution of _____ has a pH of 7.0.

- A) Na_2S
- B) KF
- C) $NaNO_3$
- D) NH_4Cl
- E) NaF

34) Of the following, which is the strongest acid?

- A) HIO
- B) HIO_4
- C) HIO_2
- D) HIO_3
- E) The acid strength of all of the above is the same.

35) Which of the following acids will be the strongest?

- A) H_2SO_4
- B) HSO_4^-
- C) H_2SO_3
- D) H_2SeO_4
- E) HSO_3^-

36) A solution of acetic acid is 2.0% dissociated at $25.0^\circ C$. What was the original concentration (in M) of the acetic acid solution? The K_a for acetic acid is 1.8×10^{-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 $25.0^\circ C$? The K_a at $25.0^\circ C$ for acetic acid is 1.8×10^{-5} .

38) TRUE or FALSE

A Lewis acid is an electron-pair acceptor, and a Lewis base is an electron-pair donor.

1. B
2. B
3. D
4. C
5. B
6. C
7. A
8. D
9. C
10. A
11. B
12. A
13. B
14. B
15. C
16. C
17. D
18. A
19. B
20. D
21. B
22. C
23. A
24. D
25. C
26. E
27. A
28. A
29. D
30. C
31. D
32. A
33. C
34. B
35. A
36. 0.045
37. 8.87
38. TRUE