Chapter 11 – Acids and Bases – Practice Problems

Section 11.1 – Acids and Bases

Goal: Describe and name acids and bases.

Summary:
An Arrhenius acid produces H⁺ and an Arrhenius base produces OH⁻ in aqueous solutions.
Acids taste sour, may sting, and neutralize bases.
Bases taste bitter, feel slippery, and neutralize acids.

Naming acids:

Binary acids contain a single anion: \( H_nX \). To name:
- hydro [anion with -ic end] acid
  
  \[ HBr \text{ – hydrobromic acid} \]
  
  \[ H_2S \text{ – hydrosulfuric acid} \]

Polyatomic acids contain a polyatomic ion: \( H_nXO_m \) (\( XO_m \) = polyatomic ion)
- [polyatomic ion] acid
  
  if the polyatomic ion ends in -ate change to -ic
  if the polyatomic ion ends in -ite change to -ous

  \[ H_2SO_4 \text{ (sulfate) – sulfuric acid} \]
  
  \[ H_2SO_3 \text{ (sulfite) – sulfous acid} \]

Practice Problems

1. Indicate whether each of the following statements is characteristic of an Arrhenius acid, Arrhenius base, or both:
   a. has a sour taste
   b. neutralizes bases
   c. produces H⁺ ions in water
   d. is named barium hydroxide
   e. is an electrolyte

2. Indicate whether each of the following statements is characteristics of an Arrhenius acid, Arrhenius base, or both:
   a. neutralizes acids
   b. produces OH⁻ ions in water
   c. has a slippery feel
   d. conducts an electrical current in solution

3. Identify each of the following as an Arrhenius acid, Arrhenius base, or none:
   a. CsOH
   b. Mg(NO₃)₂
   c. HClO₄
   d. HNO₂
   e. Neither
4. Name each of the following acids or bases:
   a. HCl \text{ hydrochloric acid }
   b. Ca(OH)_2 \text{ calcium hydroxide }
   c. HClO_4 \text{ perchlorate } \rightarrow \text{ perchloric acid }
   d. HNO_3 \text{ NO}_3^- = \text{ nitrate } \rightarrow \text{ nitric acid }
   e. H_2SO_3 \text{ SO}_3^2- = \text{sulfite } \rightarrow \text{sulfous acid }

5. Name each of the following acids or bases:
   a. Al(OH)_3 \text{ aluminum hydroxide }
   b. H_2SO_4 \text{ SO}_4^{2-} = \text{sulfate } \rightarrow \text{sulfuric acid }
   c. HBr \text{ hydrobromic acid }
   d. KOH \text{ potassium hydroxide }
   e. HNO_2 \text{ NO}_2^- = \text{ nitrite } \rightarrow \text{ nitrous acid }
   f. HClO_2 \text{ ClO}_2^- = \text{ chlorite } \rightarrow \text{ chlorous acid }

6. Write formulas for each of the following acids or bases:
   a. rubidium hydroxide \quad \text{Rb}_2^+ \text{OH}^- \rightarrow \text{Rb}_3\text{OH}
   b. hydrofluoric acid 
   \quad \text{flouric} = \text{flourine} \quad \text{H}^+ \text{F}^- \rightarrow \text{HF}
   c. phosphoric acid 
   \quad \text{phosphoric} = \text{phosphate} \equiv \text{PO}_4^{3-} \quad \text{H}^+ \text{PO}_4^{3-} \rightarrow \text{H}_3\text{PO}_4
   d. lithium hydroxide 
   \quad \text{Li}^+ \text{OH}^- \rightarrow \text{LiOH}
   e. ammonium hydroxide 
   \quad \text{NH}_4^+ \text{OH}^- \rightarrow \text{NH}_4\text{OH}

7. Write formulas for each of the following acids or bases:
   a. barium hydroxide \quad \text{Ba}^{2+} \text{OH}^- \rightarrow \text{Ba(OH)}_2
   b. hydroiodic acid 
   \quad \text{iodic} = \text{iodine} \quad \text{H}^+ \text{I}^- \rightarrow \text{HI}
   c. nitric acid 
   \quad \text{nitric} = \text{nitrate} \equiv \text{NO}_3^- \quad \text{H}^+ \text{NO}_3^- \rightarrow \text{HNO}_3
   d. acetic acid 
   \quad \text{acetic} = \text{acetate} \equiv \text{CH}_3\text{COO}^- \quad \text{CH}_3\text{COO}^- \rightarrow \text{H}^+ \text{CH}_3\text{COO}_3
   e. hypochlorous acid 
   \quad \text{hypochlorite} = \text{ClO}^- 
   \quad \text{H}^+ \text{ClO}^- \rightarrow \text{HClO}
Section 11.2 – Bronsted-Lowery Acids

Goal: Identify conjugate acid-base pairs for Bronsted-Lowry acids and bases.

Summary:
A Bronsted-Lowry acid donates $\text{H}^+$ and a Bronsted-Lowry base accepts $\text{H}^+$.

Identifying Conjugate Acid-Base Pairs

- According to Bronsted-Lowry theory, a conjugate acid-base pair consists of molecules or ions related by the loss of one $\text{H}^+$ by an acid, an the gain of one $\text{H}^+$ by a base.
- Every acid-base reaction contains two conjugate acid-base pairs because an $\text{H}^+$ is transferred in both the forward and reverse directions.

When an acid such as HF loses one $\text{H}^+$, the conjugate base $\text{F}^-$ is formed. When $\text{H}_2\text{O}$ acts as a base, it gains one $\text{H}^+$, which forms its conjugate acid, $\text{H}_3\text{O}^+$.

Example

Identify the conjugate acid-base pairs in the following reaction:

$$\text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HSO}_4^-(aq) + \text{H}_3\text{O}^+(aq)$$

Answer

Conjugate acid–base pairs: $\text{H}_2\text{SO}_4/\text{HSO}_4^-$ and $\text{H}_3\text{O}^+$/H$_2$O

Practice Problems

8. Write the formula for the conjugate base for each of the following acids:
   a. $\text{HCO}_3^-$  
   b. $\text{HPO}_4^{2-}$  
   c. $\text{HBrO}$

9. Write the formula for the conjugate acid for each of the following bases:
   a. $\text{CO}_3^{2-}$  
   b. $\text{H}_2\text{O}$  
   c. $\text{H}_2\text{PO}_4^-$
10. In the following reaction, the acid-conjugate base pair is \_i\_ and the base-conjugate acid pair is \_ii\_

\[ \text{H}_3\text{PO}_4(aq) + H_2O(l) \rightleftharpoons H_2\text{PO}_4(aq) + H_3\text{O}^+(aq) \]

a. (i) H$_3$PO$_4$/H$_2$O$^+$
   (ii) H$_2$O/H$_2$PO$_4^-$

b. (i) H$_3$PO$_4$/H$_2$O
   (ii) H$_2$PO$_4^-$/H$_3$O$^-$

c. (i) H$_2$O/H$_3$O$^+$
   (ii) H$_3$PO$_4$/H$_2$PO$_4^-$

d. (i) H$_2$O/H$_2$PO$_4$ 
   (ii) H$_3$PO$_4$/H$_3$O$^+$

e. (i) H$_3$PO$_4$/H$_2$PO$_4^-$ 
   (ii) H$_2$O/H$_3$O$^+$

11. In the following reaction, the acid-conjugate base pair is \_i\_ and the base-conjugate acid pair is \_ii\_

\[ \text{CO}_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq) \]

a. (i) CO$_3^{2-}$/H$_2$O
   (ii) HCO$_3^-$/OH$^-$

b. (i) H$_2$O/HCO$_3^-$
   (ii) CO$_3^{2-}$/OH$^-$

c. (i) H$_2$O/OH$^-$
   (ii) CO$_3^{2-}$/HCO$_3^-$

d. (i) CO$_3^{2-}$/OH$^-$
   (ii) H$_2$O/HCO$_3^-$

e. (i) CO$_3^{2-}$/HCO$_3^-$
   (ii) H$_2$O/OH$^-$

12. In the following reaction, the acid-conjugate base pair is \_i\_ and the base-conjugate acid pair is \_ii\_

\[ \text{H}_3\text{PO}_4(aq) + \text{NH}_3(aq) \rightleftharpoons H_2\text{PO}_4(aq) + \text{NH}_4^+(aq) \]

a. (i) H$_3$PO$_4$/H$_2$PO$_4^-$
   (ii) NH$_3$/NH$_4^+$

b. (i) H$_3$PO$_4$/NH$_3$
   (ii) NH$_3$/H$_2$PO$_4^-$

c. (i) H$_3$PO$_4$/NH$_3$
   (ii) H$_2$PO$_4^-$/NH$_4^+$

d. (i) H$_2$PO$_4^-$/H$_3$PO$_4$
   (ii) NH$_4^+$/NH$_3$

e. (i) NH$_3$/NH$_4^+$
   (ii) H$_3$PO$_4$/H$_2$PO$_4^-$

13. Complete the following table:

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>HI</td>
<td>I$^-$</td>
</tr>
<tr>
<td>HCl</td>
<td>Cl$^-$</td>
</tr>
<tr>
<td>NH$_4^+$</td>
<td>NH$_3$</td>
</tr>
<tr>
<td>H$_2$S</td>
<td>HS$^-$</td>
</tr>
</tbody>
</table>

14. Complete the following table:

<table>
<thead>
<tr>
<th>Base</th>
<th>Conjugate Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>F$^-$</td>
<td>HF</td>
</tr>
<tr>
<td>C$_2$H$_5$O$_2^-$</td>
<td>HC$_2$H$_3$O$_2$</td>
</tr>
<tr>
<td>SO$_3^{2-}$</td>
<td>HSO$_3^-$</td>
</tr>
<tr>
<td>ClO$^-$</td>
<td>HClO</td>
</tr>
</tbody>
</table>
15. When ammonium chloride dissolves in water, the ammonium ion \( \text{NH}_4^+ \) donates an \( H^+ \) to water. Write a balanced equation for the reaction of the ammonium ion with water.

- \( \text{NH}_4^+ + H_2O \leftrightarrow \text{NH}_3 + H_3O^+ \)
- \( \text{NH}_4^+ + H_2O \leftrightarrow \text{NH}_3 + H_2O + H^+ \)
- \( \text{NH}_4^+ + H_2O \leftrightarrow \text{NH}_3 + H_2O + OH^- \)
- \( \text{NH}_3 + H_2O \leftrightarrow \text{NH}_2^- + H_3O^+ \)
- \( \text{NH}_3 + H_2O \leftrightarrow \text{NH}_4^+ + H_2O \)

Read the question carefully; it describes the equation.

16. When sodium carbonate dissolves in water, the carbonate ion \( \text{CO}_3^{2-} \) acts as a base. Write a balanced equation for the reaction of the carbonate ion with water.

- \( \text{CO}_3^{2-} + H_2O \leftrightarrow \text{CO}_2 + H_3O^+ \)
- \( \text{CO}_3^{2-} + H_2O \leftrightarrow \text{HCO}_3^- + OH^- \)
- \( \text{CO}_3^{2-} + H_2O \leftrightarrow \text{H}_2\text{CO}_3 \)
- \( \text{CO}_3^{2-} + OH^- \leftrightarrow \text{CO}_3^{2-} + H_2O \)
- \( \text{CO}_3^{2-} + H_2O \leftrightarrow \text{H}_2\text{CO}_3 \)

Section 11.3 – Strengths of Acids and Bases

Goal: Write equations for the dissociation of strong and weak acids; identify the direction of reaction.

Summary:

Strong acids dissociate completely in water, and the \( H^+ \) is accepted by \( H_2O \) acting as a base.

\[ \text{HCl} + H_2O \rightarrow H_3O^+ + \text{Cl}^- \]

A weak acid dissociates only slightly in water, producing only a small amount of \( H^+ \) and therefore a small amount of \( H_3O^+ \)

\[ \text{HI} + H_2O \leftrightarrow H_3O^+ + \text{I}^- \]

Strong bases are hydroxides with metals from Groups 1 and 2 and dissociate completely in water. (NaOH, Ca(OH)\(_2\), …)

An important weak base is ammonia, \( \text{NH}_3 \).

Understanding the Concepts:

In diagrams A and B, determine if the diagram represents a strong acid or a weak acid. The acid has the formula \( \text{HX} \).
### TABLE 11.3 Relative Strengths of Acids and Bases

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Strong Acids</strong></td>
<td></td>
</tr>
<tr>
<td>Hydroiodic acid</td>
<td>I(^{-})</td>
</tr>
<tr>
<td>Hydrobromic acid</td>
<td>Br(^{-})</td>
</tr>
<tr>
<td>perchloric acid</td>
<td>ClO(_4)(^{-})</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>Cl(^{-})</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>HSO(_4)(^{-})</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>NO(_3)(^{-})</td>
</tr>
<tr>
<td>Hydronium ion</td>
<td>H(_3)O(^{+})</td>
</tr>
<tr>
<td><strong>Weak Acids</strong></td>
<td></td>
</tr>
<tr>
<td>Hydrogen sulfate ion</td>
<td>SO(_4)(^{2-})</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H(_2)PO(_4)(^{-})</td>
</tr>
<tr>
<td>Nitrous acid</td>
<td>NO(_2)(^{-})</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>F(^{-})</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>C(_2)H(_3)O(_2)(^{-})</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>HCO(_3)(^{-})</td>
</tr>
<tr>
<td>Hydrosulfuric acid</td>
<td>HS(^{-})</td>
</tr>
<tr>
<td>Dihydrogen phosphate ion</td>
<td>HPO(_4)(^{2-})</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH(_3)(^{-})</td>
</tr>
<tr>
<td>Hydrocyanic acid</td>
<td>CN(^{-})</td>
</tr>
<tr>
<td>Bicarbonate ion</td>
<td>CO(_3)(^{2-})</td>
</tr>
<tr>
<td>Methylammonium ion</td>
<td>CH(_3)---NH(_3)(^{+})</td>
</tr>
<tr>
<td>Hydrogen phosphate ion</td>
<td>PO(_4)(^{3-})</td>
</tr>
<tr>
<td>Water</td>
<td>OH(^{-})</td>
</tr>
</tbody>
</table>

**Practice Problems**

17. Using Table 11.3, identify the **stronger** acid in each of the following pairs:
   a. NH\(_4\)\(^{+}\) or H\(_2\)O\(^{+}\)
   b. H\(_2\)SO\(_4\) or HCl
   c. H\(_2\)O or H\(_2\)CO\(_3\)

18. Using Table 11.3, identify the **weaker** acid in each of the following pairs:
   a. HCl or HSO\(_4\)\(^{-}\)
   b. HNO\(_2\) or HF
   c. HCO\(_3\)\(^{-}\) or NH\(_4\)\(^{+}\)

19. Predict whether the following reaction contains mostly reactants or products at equilibrium:

\[ \text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_3(aq) + \text{H}_3\text{O}^+(aq) \]

   a. mostly products
   b. mostly reactants

- **Acids:** H\(_2\)CO\(_3\) + H\(_3\)O\(^{+}\)
- **Bases:** H\(_2\)O + HCO\(_3\)
20. Predict whether the following reaction contains mostly reactants or products at equilibrium:

\[ \text{NH}_4^+ (aq) + H_2O(l) \rightleftharpoons \text{NH}_3(aq) + H_3O^+(aq) \]

\[ \text{Acids: NH}_4^+ \text{ and H}_3\text{O}^+ \]
\[ \text{weaker} \quad \text{stronger} \]
\[ \text{Bases: H}_2\text{O and NH}_3 \]
\[ \text{weaker} \quad \text{stronger} \]

a. mostly products
b. mostly reactants

21. Predict whether the following reaction contains mostly reactants or products at equilibrium:

\[ \text{HNO}_3(aq) + \text{NH}_3(aq) \rightleftharpoons \text{NO}_2^-(aq) + \text{NH}_4^+(aq) \]

\[ \text{Acids: HNO}_3 \text{ and NH}_4^+ \]
\[ \text{stronger} \quad \text{weaker} \]
\[ \text{Bases: NH}_3 \text{ and NO}_2^- \]
\[ \text{Stronger} \quad \text{weaker} \]

a. mostly products
b. mostly reactants

22. Predict whether the following reaction contains mostly reactants or products at equilibrium:

\[ \text{H}_3\text{PO}_4(aq) + H_2O(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{H}_2\text{PO}_4^-(aq) \]

\[ \text{Acids: H}_3\text{PO}_4 \text{ and H}_3\text{O}^+ \]
\[ \text{weaker} \quad \text{stronger} \]
\[ \text{Bases: H}_2\text{O and H}_2\text{PO}_4^- \]
\[ \text{weaker} \quad \text{stronger} \]

a. mostly products
b. mostly reactants

Challenge Problems

23. a. Write the formula for the conjugate base of H_2S \[ \text{HS}^- \]
b. Write the formula for the conjugate base of H_3PO_4 \[ \text{H}_2\text{PO}_4^- \]
c. Which is the weaker acid: (H_2S) or H_3PO_4? ____________

\[ \text{from table 11.3} \]

24. a. Write the formula for the conjugate base of HCO_3^- \[ \text{CO}_3^{2-} \]
b. Write the formula for the conjugate base of HC_2H_3O_2 \[ \text{C}_2\text{H}_5\text{O}_2^- \]
c. Which is the stronger acid: HCO_3^- or HC_2H_3O_2? ____________
Section 11.4 – Dissociation Constants for Acids and Bases

Goal: Write the expression for the dissociation constant of a weak acid or weak base.

Summary:
In water, weak acids and weak bases produce only a few ions when equilibrium is reached.

Weak acids have small $K_a$ values, whereas strong acids, which are essentially 100% dissociated, have very large $K_a$ values.

The reaction for a weak acid can be written as $HA + H_2O \rightleftharpoons H_3O^+ + A^-$. The acid dissociation constant expression is written as

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

For a weak base the equation is $B + H_2O \rightleftharpoons BH^+ + OH^-$, and the base dissociation constant expression is written as

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Practice Problems

25. Answer true or false for each of the following: A strong acid...
   a. is completely dissociated in aqueous solution True
   b. has a small value of $K_a$ False
   c. has a strong conjugate base True
   d. has a weak conjugate base False
   e. is slightly dissociated in aqueous solution False

26. Answer true or false for each of the following: A weak acid...
   a. is completely dissociated in aqueous solution False
   b. has a small value of $K_a$ True
   c. has a strong conjugate base True
   d. has a weak conjugate base False
   e. is slightly dissociated in aqueous solution True

27. Consider the following acids and their dissociation constants:
   \[ H_2SO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_3^-(aq) \]
   \[ K_a = 1.2 \times 10^{-2} \]
   \[ HS^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + S^{2-}(aq) \]
   \[ K_a = 1.3 \times 10^{-19} \]
   a. Which is the stronger acid, $H_2SO_3$ or HS? $H_2SO_3$
   b. What is the conjugate base of $H_2SO_3$? $HSO_3^-$
   c. Which acid has the weaker conjugate base? $HS^-$
   d. Which acid has the stronger conjugate base? $H_2SO_3$
   e. Which acid produces more ions? $H_2SO_3$

the bigger the $\#$ the stronger the acid
the stronger the acid, the weaker the conj. B and vice versa.
28. Consider the following acids and their dissociation constants:

\[
\text{HPO}_4^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{PO}_4^{3-}(aq) \quad K_a = 2.2 \times 10^{-13} \text{ weaker}
\]

\[
\text{HCHO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CHO}_2^-(aq) \quad K_a = 1.8 \times 10^{-4} \text{ stronger}
\]

a. Which is the weaker acid, \(\text{HPO}_4^{2-}\) or \(\text{HCHO}_2\)?

b. What is the conjugate base of \(\text{HPO}_4^{2-}\)?

c. Which acid has the weaker conjugate base?

d. Which acid has the stronger conjugate base?

e. Which acid produces more ions?

29. Aniline, \(\text{C}_6\text{H}_5\text{NH}_2\), a weak base with a \(K_b\) of \(4.0 \times 10^{-10}\), reacts with water to form \(\text{C}_6\text{H}_5\text{NH}_3^+\) and hydroxide ion. Write the equation for the reaction and the base dissociation constant expression for aniline:

\[
\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^- \quad K_b = 4.0 \times 10^{-10}
\]

Section 11.5 – Dissociation of Water

Goal: Use the water dissociation constant expression to calculate the \([\text{H}_3\text{O}^+]\) and \([\text{OH}^-]\) in an aqueous solution.

Summary:

- In pure water, a few water molecules transfer \(\text{H}^+\) to other water molecules, producing small, but equal amounts of \([\text{H}_3\text{O}^+]\) and \([\text{OH}^-]\).

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

- In pure water, the molar concentrations of \(\text{H}_3\text{O}^+\) and \(\text{OH}^-\) are each \(1.0 \times 10^{-7}\)M.
- The water dissociation constant expression, \(K_w\):

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25°C
\]

- In acidic solutions, the \([\text{H}_3\text{O}^+]\) is greater than the \([\text{OH}^-]\).
- In basic solutions, the \([\text{OH}^-]\) is greater than the \([\text{H}_3\text{O}^+]\).

Neutral solution

Calculating \([\text{H}_3\text{O}^+]\) and \([\text{OH}^-]\) in solutions

If we know the \([\text{H}_3\text{O}^+]\) of a solution, we can use the \(K_w\) expression to calculate the \([\text{OH}^-]\).

If we know the \([\text{OH}^-]\) of a solution, we can calculate the \([\text{H}_3\text{O}^+]\) using the \(K_w\) expression.

\[
[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}
\]

Example What is the \([\text{OH}^-]\) in a solution that has \([\text{H}_3\text{O}^+] = 2.4 \times 10^{-11}\)M? Is the solution acidic or basic?

Answer We solve the \(K_w\) expression for \([\text{OH}^-]\) and substitute in the known values of \(K_w\) and \([\text{H}_3\text{O}^+]\).

\[
[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{2.4 \times 10^{-11}} = 4.2 \times 10^{-4}\text{M}
\]

Because the \([\text{OH}^-]\) is greater than the \([\text{H}_3\text{O}^+]\), this is a basic solution.
Understanding the Concepts:

Why are the concentrations of $H_3O^+$ and $OH^-$ equal in pure water?

One of each are produced every time a $H^+$ is transferred from one water molecule to another.

$$2H_2O \leftrightarrow OH^- + H_3O^+$$

What is the meaning and value of $K_w$ at 25°C? $K_w = 1.0 \times 10^{-14}$ is the $K_a$ of...

In an acidic solution, how does the concentration of $H_3O^+$ compare to the concentration of $OH^-$?

$[H_3O^+]$ is greater than $[OH^-]$

If a base is added to pure water, why does the $[H_3O^+]$ decrease?

Recall $H_3O^+ = H_2O + H^+$. The $H^+$ is the acid part. A base is an $H^+$ acceptor so it steals $H^+$ from $H_3O^+$ which decreases the amount of $H_3O^+$.

Practice Problems

30. Indicate whether each of the following solutions is acidic, basic, or neutral?

a. $[H_3O^+] = 6.0 \times 10^{-12}$ M
   
   Basic

b. $[H_3O^+] = 1.4 \times 10^{-4}$ M
   
   Acidic

c. $[OH^-] = 5.0 \times 10^{-12}$ M
   
   Acidic

d. $[OH^-] = 4.5 \times 10^{-2}$ M
   
   Basic

31. Calculate the $[H_3O^+]$ of each aqueous solution with the following $[OH^-]$: The next section will give the equation for pH. These based on:

- $[H_3O^+] > 1.0 \times 10^{-7}$ M acidic
- $[H_3O^+] = 1.0 \times 10^{-7}$ M neutral
- $[H_3O^+] < 1.0 \times 10^{-7}$ M basic
- $[OH^-]$ has the opposite trend

a. NaOH, $1.0 \times 10^{-2}$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12} \text{ M}$$

b. milk of magnesia, $1.0 \times 10^{-5}$ M

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} \text{ M}$$

c. aspirin, $1.8 \times 10^{-11}$ M

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-11}} = 5.56 \times 10^{-4} \text{ M}$$

d. seawater, $2.5 \times 10^{-6}$ M

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-6}} = 4.0 \times 10^{-9} \text{ M}$$

32. Calculate the $[OH^-]$ of each aqueous solution with the following $[H_3O^+]$:

- $[H_3O^+] = \frac{K_w}{[OH^-]}$

a. stomach acid, $4.0 \times 10^{-2}$ M

$$[OH^-] = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-2}} = 2.5 \times 10^{-13} \text{ M}$$

b. urine, $5.0 \times 10^{-6}$ M

$$[OH^-] = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-6}} = 2.0 \times 10^{-9} \text{ M}$$

c. orange juice, $2.0 \times 10^{-4}$ M

$$[OH^-] = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-4}} = 5.0 \times 10^{-10} \text{ M}$$

d. bile, $7.9 \times 10^{-9}$ M

$$[OH^-] = \frac{1.0 \times 10^{-14}}{7.9 \times 10^{-9}} = 1.27 \times 10^{-6} \text{ M}$$
Section 11.6 – The pH Scale

Goal: Calculate pH from [H₃O⁺]; given the pH, calculate the [H₃O⁺] and [OH⁻] of a solution.

Summary:
The **pH scale** is a range of numbers typically from 0 to 14, which represents the [H₃O⁺] of the solution.

A neutral solution has a pH of 7.0.
In acidic solutions, the pH is below 7.0.
In basic solutions, the pH is above 7.0.

Mathematically, pH is the negative logarithm of the hydronium ion concentration,

\[
\text{pH} = -\log[\text{H}_3\text{O}^+]\]

Other useful equations:

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}
\]

\[
[\text{H}_3\text{O}^+] = 10^{-\text{pH}}
\]
Understanding the Concepts

Why does a neutral solution have a pH of 7.0?

If you know the [OH⁻], how can you determine the pH of a solution?

Use \([\text{OH}^-]\) to get \([H_3O^+]\) then use \(\text{pH} = -\log [H_3O^+]\).

State whether each of the following solutions is acidic, basic, or neutral:

a. blood plasma, pH 7.38 basic
b. vinegar, pH 2.8 acidic
c. coffee, pH 5.52 acidic
d. tomatoes, pH 4.2 acidic
e. chocolate cake, pH 7.6 basic

State whether each of the following solutions is acidic, basic, or neutral:

a. soda, pH 3.22 acidic
b. shampoo, pH 5.7 acidic
c. rain, pH 5.8 acidic
d. honey, pH 3.9 acidic
e. cheese, pH 5.2 acidic

A solution with a pH of 3 is 10 times more acidic than a solution with pH 4. Explain.

The pH scale is logarithmic. So every increase in 1.0 is ten times stronger.

A solution with a pH of 10 is 100 times more basic than a solution with pH 8. Explain.

8 \rightarrow 9 \rightarrow 10 \text{ 2 steps. 10} \times 10 = 100\text{ times stronger.}

Practice Problems

33. Calculate the pH of each solution given the following:

a. \([H_3O^+] = 1 \times 10^{-8} \text{M}\) \(\text{pH} = -\log (1 \times 10^{-8}) = 8.00\)

b. \([H_3O^+] = 5 \times 10^{-6} \text{M}\) \(\text{pH} = -\log (5 \times 10^{-6}) = 5.30\)

1.0 \times 10^{-14} = K_w = [OH^-][H_3O^+]

c. \([\text{OH}^-] = 1 \times 10^{-2} \text{M}\) \(\frac{K_w}{[\text{OH}^-]} = \frac{[H_3O^+]}{1 \times 10^{-2}} = 1.0 \times 10^{-12}\)

\(\text{pH} = -\log (1.0 \times 10^{-12}) = 12.00\)

d. \([\text{OH}^-] = 8.0 \times 10^{-3} \text{M}\)

\([H_3O^+] = \frac{1.0 \times 10^{-14}}{8.0 \times 10^{-3}} = 1.25 \times 10^{-12}\)

\(\text{pH} = -\log (1.25 \times 10^{-12}) = 11.90\)

e. \([H_3O^+] = 4.7 \times 10^{-2} \text{M}\) \(\text{pH} = -\log (4.7 \times 10^{-2}) = 1.33\)

f. \([\text{OH}^-] = 3.9 \times 10^{-6} \text{M}\)

\([H_3O^+] = \frac{1.0 \times 10^{-14}}{3.9 \times 10^{-6}} = 2.56 \times 10^{-9}\)

\(\text{pH} = -\log (2.56 \times 10^{-9}) = 8.59\)
34. Complete the following table for a selection of foods:

<table>
<thead>
<tr>
<th>Food</th>
<th>$[\text{H}_3\text{O}^+]$</th>
<th>$[\text{OH}^-]$</th>
<th>pH</th>
<th>Acidic, Basic, or Neutral</th>
</tr>
</thead>
<tbody>
<tr>
<td>Rye bread</td>
<td>$1.59 \times 10^{-9}$</td>
<td>$6.3 \times 10^{-6}$</td>
<td>8.80</td>
<td>Basic</td>
</tr>
<tr>
<td>Tomatoes</td>
<td>$2.29 \times 10^{-5}$</td>
<td>$4.37 \times 10^{-10}$</td>
<td>4.64</td>
<td>Acidic</td>
</tr>
<tr>
<td>Peas</td>
<td>$6.2 \times 10^{-7}$</td>
<td>$1.61 \times 10^{-8}$</td>
<td>6.20</td>
<td>Acidic</td>
</tr>
</tbody>
</table>

\[
[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = 1.0 \times 10^{-14}
\]

\[
\text{pH} = -\log (1.59 \times 10^{-9}) = 8.80
\]

35. Complete the following table for a selection of foods:

<table>
<thead>
<tr>
<th>Food</th>
<th>$[\text{H}_3\text{O}^+]$</th>
<th>$[\text{OH}^-]$</th>
<th>pH</th>
<th>Acidic, Basic, or Neutral</th>
</tr>
</thead>
<tbody>
<tr>
<td>Strawberries</td>
<td>$1.26 \times 10^{-4}$</td>
<td>$7.94 \times 10^{-11}$</td>
<td>3.90</td>
<td>Acidic</td>
</tr>
<tr>
<td>Soy milk</td>
<td>$1.0 \times 10^{-7}$</td>
<td>$1.0 \times 10^{-7}$</td>
<td>7.00</td>
<td>Neutral</td>
</tr>
<tr>
<td>Canned tuna fish</td>
<td>$6.3 \times 10^{-7}$</td>
<td>$1.59 \times 10^{-8}$</td>
<td>6.20</td>
<td>Acidic</td>
</tr>
</tbody>
</table>

\[
[\text{H}_3\text{O}^+] = 10^{-3.90} = 1.26 \times 10^{-4} \text{ M}
\]

\[
[\text{OH}^-] = \frac{1 \times 10^{-14}}{1.26 \times 10^{-4}} = 7.94 \times 10^{-11} \text{ M}
\]

36. Calculate the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for a solution with each of the following pH values:

a. $3.00$

\[
[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.00} = 1.0 \times 10^{-3} \text{ M}
\]

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} \text{ M}
\]

b. $6.2$

\[
[\text{H}_3\text{O}^+] = 10^{-6.2} = 6.31 \times 10^{-7} \text{ M}
\]

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{6.31 \times 10^{-7}} = 1.58 \times 10^{-8} \text{ M}
\]

c. $8.85$

\[
[\text{H}_3\text{O}^+] = 10^{-8.85} = 1.41 \times 10^{-9} \text{ M}
\]

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.41 \times 10^{-9}} = 7.08 \times 10^{-6} \text{ M}
\]
37. Solution A has a pH of 4.5, and solution B has a pH of 6.7.
   a. Which solution is more acidic? **Solution A**
   b. What is the \([\text{H}_3\text{O}^+]\) in each?
      \[
      \text{Sol. A: } [\text{H}_3\text{O}^+] = 10^{-4.5} = 3.16 \times 10^{-5} \text{ M}
      \]
      \[
      \text{Sol. B: } [\text{H}_3\text{O}^+] = 10^{-6.7} = 2.0 \times 10^{-7} \text{ M}
      \]
   c. What is the \([\text{OH}^-]\) in each?
      \[
      \text{Sol. A: } [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{3.16 \times 10^{-5}} = 3.16 \times 10^{-10} \text{ M}
      \]
      \[
      \text{Sol. B: } [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-7}} = 5.00 \times 10^{-8} \text{ M}
      \]

**Section 11.7 – Reactions of Acids and Bases**

**Goal:** Write balanced equations for reactions of acids with metals, carbonates or bicarbonates, and bases.

**Summary:**

An acid reacts with a metal to produce hydrogen gas (\(\text{H}_2\)) and a salt.

\[
\text{Mg(s) + 2HCl(aq) } \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq)
\]

Metal \quad Acid \quad Hydrogen \quad Salt

The reaction of an acid with a carbonate (\(\text{CO}_3^{2-}\)) or bicarbonate (\(\text{HCO}_3^-\)) produces carbon dioxide, water, and a salt.

\[
2\text{HCl(aq) + Na}_2\text{CO}_3(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O(l)} + 2\text{NaCl(aq)}
\]

Acid \quad Carbonate \quad Carbon \quad Water \quad Salt

dioxide

In neutralization, a strong or weak acid reacts with a strong base to produce water and a salt.

\[
\text{HCl(aq) + NaOH(aq) } \rightarrow \text{H}_2\text{O(l) + NaCl(aq)}
\]

Acid \quad Base \quad Water \quad Salt

**Example** Write the balanced chemical equation for the reaction of \(\text{ZnCO}_3(s)\) and hydrobromic acid \(\text{HBr}(aq)\).

**Answer**

\[
\text{ZnCO}_3(s) + 2\text{HBr}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O(l)} + \text{ZnBr}_2(aq)
\]

**Practice Problems**

38. Complete and balance the equation for each of the following reactions:
   a. \(\text{ZnCO}_3(s) + 2\text{HBr}(aq) \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{ZnBr}_2\)
      balance the \(\text{CO}_2 + \text{H}_2\text{O}\) 1st, whatever is left goes into the salt.
   b. \(\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2\)
c. \( \text{HCl(aq)} + \text{NaHCO}_3(s) \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{NaCl} \)

d. \( \text{H}_2\text{SO}_4(aq) + \text{Mg(OH)}_2(s) \rightarrow 2\text{H}_2\text{O} + \text{MgSO}_4 \)

39. Complete and balance the equation for each of the following reactions:

a. \( \text{KHCO}_3(s) + \text{HBr(aq)} \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{KBr} \)

b. \( \text{Ca(s)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{H}_2 + \text{CaSO}_4 \)

c. \( \text{H}_2\text{SO}_4(aq) + \text{Ca(OH)}_2(s) \rightarrow 2\text{H}_2\text{O} + \text{CaSO}_4 \)

d. \( \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{Na}_2\text{SO}_4 \)

40. Balance each of the following neutralization reactions:

a. \( 2\text{HCl(aq)} + \text{Mg(OH)}_2(s) \rightarrow 2\text{H}_2\text{O(l)} + \text{MgCl}_2(aq) \)

41. Write a balanced equation for the neutralization of each of the following:

a. \( \text{H}_2\text{SO}_4(aq) \) and \( \text{NaOH(aq)} \)  
\( \text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4 \)

b. \( \text{HCl(aq)} \) and \( \text{Fe(OH)}_3(s) \)  
\( 3\text{HCl} + \text{Fe(OH)}_3 \rightarrow 3\text{H}_2\text{O} + \text{FeCl}_3 \)

c. \( \text{H}_2\text{CO}_3(aq) \) and \( \text{Mg(OH)}_2(s) \)  
\( \text{H}_2\text{CO}_3 + \text{Mg(OH)}_2(s) \rightarrow 2\text{H}_2\text{O} + \text{MgCO}_3 \)

Section 11.8 – Buffers
Goal: Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.
Summary:

A buffer solution maintains pH by neutralizing small amounts of an acid or base.

Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base such as acetic acid, \( \text{HC}_2\text{H}_3\text{O}_2 \), and its salt \( \text{NaC}_2\text{H}_3\text{O}_2 \).

The \( [\text{H}_3\text{O}^+] \) is calculated by solving the \( K_a \) expression for \( [\text{H}_3\text{O}^+] \), then substituting the values of \( [\text{H}_3\text{O}^+] \), \( [\text{HA}] \), and \( K_a \) into the equation.

\[
K_a = \frac{[\text{H}_3\text{O}^+] [\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} 
\]

Solving for \( [\text{H}_3\text{O}^+] \) gives:

\[
[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]} \quad \text{Weak acid} \quad \text{Conjugate base}
\]

The pH of the buffer is calculated from the \( [\text{H}_3\text{O}^+] \).

\[
\text{pH} = -\log[\text{H}_3\text{O}^+] 
\]

Example

What is the pH of a buffer prepared with 0.40 M \( \text{HC}_2\text{H}_3\text{O}_2 \) and 0.20 M \( \text{C}_2\text{H}_3\text{O}_2^- \), if the \( K_a \) of acetic acid is \( 1.8 \times 10^{-5} \)?

Answer

\[
[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 1.8 \times 10^{-5} \times \frac{[0.40]}{[0.20]} = 3.6 \times 10^{-5} \text{ M} 
\]

\[
\text{pH} = -\log[3.6 \times 10^{-5}] = 4.44
\]

Understanding the Concepts:
43. Consider the buffer system of hydrofluoric acid, HF, and its salt, NaF:
\[ \text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq) \]

(i) The purpose of this buffer system is to:
   a. maintain [HF]
   b. maintain [F-]
   c. maintain pH

(ii) The salt of the weak acid is needed to:
   a. provide the conjugate base
   b. neutralize added \( \text{H}_3\text{O}^+ \)
   c. provide the conjugate acid

(iii) If \( \text{OH}^- \) is added, it is neutralized by:
   a. the salt
   b. \( \text{H}_2\text{O} \)
   c. \( \text{H}_3\text{O}^+ \)

(iv) When \( \text{H}_3\text{O}^+ \) is added, the equilibrium shifts in the direction of the:
   a. reactants
   b. products
   c. does not change

44. Consider the buffer system of nitrous acid, HNO\(_2\), and its salt, NaNO\(_2\):
\[ \text{HNO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NO}_2^-(aq) \]

(i) The purpose of this buffer system is to:
   a. maintain [HNO\(_2\)]
   b. maintain [NO\(_2^-\)]
   c. maintain pH

(ii) The weak acid is needed to:
   a. provide the conjugate base
   b. neutralize added OH\(^-\)
   c. provide the conjugate acid

(iii) If \( \text{H}_3\text{O}^+ \) is added it is neutralized by:
   a. the salt
   b. \( \text{H}_2\text{O} \)
   c. OH\(^-\)

(iv) When OH\(^-\) is added, the equilibrium shifts in the direction of the:
   a. reactants
   b. products
   c. does not change
44. Consider the buffer system of nitrous acid, HNO₂, and its salt, NaNO₂:
\[
\text{HNO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NO}_2^-(aq)
\]
(i) The purpose of this buffer system is to:
   a. maintain [HNO₂]
   b. maintain [NO₂⁻]
   c. maintain pH

(ii) The weak acid is needed to:
   a. provide the conjugate base
   b. neutralize added OH⁻
   c. provide the conjugate acid

(iii) If H₃O⁺ is added it is neutralized by:
   a. the salt
   b. H₂O
   c. OH⁻

(iv) When OH⁻ is added, the equilibrium shifts in the direction of the:
   a. reactants
   b. products
   c. does not change

\[
\text{OH}^- \text{reacts with H}_3\text{O}^+ \text{which decreases the overall amount of H}_3\text{O}^+ \text{. Since it is a product in the above reaction, the reaction shifts toward the products. (decrease product concentration)}
\]

45. Nitrous acid has a \( K_a \) of \( 4.5 \times 10^{-4} \). What is the pH of a buffer solution containing 0.10M HNO₂ and 0.010M NO₂⁻?
   
   - \( \text{c. } 3.35 \)

\[
[\text{H}_3\text{O}^+] = (4.5 \times 10^{-4}) \left( \frac{0.1 \text{ M}}{0.01 \text{ M}} \right)
\]

\[
[\text{H}_3\text{O}^+] = 0.045 \text{ M}
\]

\[
\text{pH} = -\log(0.045) = 2.35
\]

46. Acetic acid has a \( K_a \) of \( 1.8 \times 10^{-5} \). What is the pH of a buffer solution containing 0.25M H₃C₂H₅O₂ and 0.15M C₂H₅O₂⁻?

   a. 4.74
   b. 4.52
   c. 4.97
   d. 3.00 \times 10^{-5}
   e. 1.08 \times 10^{-5}

\[
[\text{H}_3\text{O}^+] = (1.8 \times 10^{-5}) \left( \frac{0.25}{0.15} \right)
\]

\[
[\text{H}_3\text{O}^+] = 3.0 \times 10^{-5} \text{ M}
\]

\[
\text{pH} = -\log(3.0 \times 10^{-5}) = 4.52
\]

Chlorine Problems
44. A buffer solution is made by dissolving H₂PO₄ and NaH₂PO₄ in water.