Chapter 8 Acids and Bases

8.1 Acids and Bases

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Acids

Arrhenius acids
• produce H⁺ ions in water.

\[
\text{H}_2\text{O} \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq)
\]
• are electrolytes.
• have a sour taste.
• turn litmus red.
• neutralize bases.

Names of Acids

• Acids with H and a nonmetal are named with the prefix hydro and end with ic acid.
  
  HCl  hydrochloric acid

• Acids with H and a polyatomic ion are named by changing the end of the name of the polyatomic ion from ate to ic acid or ous to ous acid.

  ClO₃⁻  chlorate  HClO₃  chloric acid
  ClO₂⁻  chlorite  HClO₂  chlorous acid

Names of Some Common Acids

<table>
<thead>
<tr>
<th>Acid</th>
<th>Name of Acid</th>
<th>Anion</th>
<th>Name of Anion</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>Hydrochloric acid</td>
<td>Cl⁻</td>
<td>Chloride</td>
</tr>
<tr>
<td>HBr</td>
<td>Hydrobromic acid</td>
<td>Br⁻</td>
<td>Bromide</td>
</tr>
<tr>
<td>HNO₃</td>
<td>Nitric acid</td>
<td>NO₃⁻</td>
<td>Nitrate</td>
</tr>
<tr>
<td>HNO₂</td>
<td>Nitrous acid</td>
<td>NO₂⁻</td>
<td>Nitrite</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>Sulfuric acid</td>
<td>SO₄²⁻</td>
<td>Sulfate</td>
</tr>
<tr>
<td>H₂SO₃</td>
<td>Sulfurous acid</td>
<td>SO₃²⁻</td>
<td>Sulfite</td>
</tr>
<tr>
<td>HCO₃⁻</td>
<td>Bicarbonate</td>
<td></td>
<td>Carbonate</td>
</tr>
<tr>
<td>H₂PO₄⁻</td>
<td>Phosphoric acid</td>
<td>PO₄³⁻</td>
<td>Phosphate</td>
</tr>
<tr>
<td>HCO₃⁻</td>
<td>Borax</td>
<td></td>
<td>Borate</td>
</tr>
<tr>
<td>HBrO₂</td>
<td>Acid bromide</td>
<td>BrO₂⁻</td>
<td>Bromide</td>
</tr>
<tr>
<td>HBrO₃</td>
<td>Acid bromide</td>
<td>BrO₃⁻</td>
<td>Bromide</td>
</tr>
</tbody>
</table>

Learning Check

Select the correct name for each of the following acids.

A. HBr
  1. bromic acid
  2. bromous acid
  3. hydrobromic acid

B. H₂CO₃
  1. carbonic acid
  2. hydrocarbonic acid
  3. carbonous acid

C. HBrO₂
  1. bromic acid
  2. hydrobromous acid
  3. bromous acid

Solution

A. HBr  3. hydrobromic acid
The name of an acid with H and a nonmetal uses the prefix hydro and ends with ic acid.

B. H₂CO₃  1. carbonic acid
An acid with H and a polyatomic ion is named by changing the end of an ate ion to ic acid.

C. HBrO₂  3. bromous acid
This acid of bromite (BrO₂⁻) is bromous acid.
Bases

Arrhenius bases
• produce OH ions in water.
• taste bitter or chalky.
• are electrolytes.
• feel soapy and slippery.
• neutralize acids.

Some Common Bases

Bases with OH ions are named as the hydroxide of the metal in the formula.

- NaOH Sodium hydroxide
- KOH Potassium hydroxide
- Ba(OH)2 Barium hydroxide
- Al(OH)3 Aluminum hydroxide
- Fe(OH)3 Iron(III) hydroxide

Learning Check

Match the formulas with the names.

A. ___HNO2 1) Iodic acid
B. ___Ca(OH)2 2) Sulfuric acid
C. ___H2SO4 3) Sodium hydroxide
D. ___HIO3 4) Nitrous acid
E. ___NaOH 5) Calcium hydroxide

Solution

Match the formulas with the names.

A. 4) HNO2 Nitrous acid
B. 5) Ca(OH)2 Calcium hydroxide
C. 2) H2SO4 Sulfuric acid
D. 1) HIO3 Iodic acid
E. 3) NaOH Sodium hydroxide

Comparing Acids and Bases

<table>
<thead>
<tr>
<th>Characteristic</th>
<th>Acids</th>
<th>Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reaction: Acid-base</td>
<td>Produce H+</td>
<td>Produce OH</td>
</tr>
<tr>
<td>Reaction: Brønsted-Lowry</td>
<td>Donate H+</td>
<td>Accept H+</td>
</tr>
<tr>
<td>Electrolytes</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>Taste</td>
<td>Sour</td>
<td>Inert, chalky</td>
</tr>
<tr>
<td>Feel</td>
<td>May sting</td>
<td>Soapy, slippery</td>
</tr>
<tr>
<td>Litmus</td>
<td>Red</td>
<td>Blue</td>
</tr>
<tr>
<td>Phenolphthalein</td>
<td>Colorless</td>
<td>Pink</td>
</tr>
<tr>
<td>Neutralization</td>
<td>Neutralize bases</td>
<td>Neutralize acids</td>
</tr>
</tbody>
</table>

Learning Check

Identify each as a characteristic of an A) acid or B) base.

___1. has a sour taste
___2. produces OH in aqueous solutions
___3. has a chalky taste
___4. is an electrolyte
___5. produces H in aqueous solutions
Identify each as a characteristic of an A) acid or B) base.

A 1. has a sour taste
B 2. produces OH⁻ in aqueous solutions
B 3. has a chalky taste
A 4. is an electrolyte
A 5. produces H⁺ in aqueous solutions

According to the Brønsted-Lowry theory, acids donate a proton (H⁺), bases accept a proton (H⁺).

In the reaction of ammonia and water, NH₃ is the base that accepts H⁺. H₂O is the acid that donates H⁺.

In any acid-base reaction, there are two conjugate acid-base pairs.

- Each pair is related by the loss and gain of H⁺.
- One pair occurs in the forward direction.
- One pair occurs in the reverse direction.

In this acid-base reaction, an acid, HF, donates H⁺ to form its conjugate base, F⁻.
- a base, H₂O, accepts H⁺ to form its conjugate acid, H₃O⁺.
- there are two conjugate acid-base pairs.

In the reaction of HF and H₂O,
- one conjugate acid-base pair is HF/F⁻.
- the other conjugate acid-base pair is H₂O/H₃O⁺.
- each pair is related by a loss and gain of H⁺.
Conjugate Acid-Base Pairs

In the reaction of NH₃ and H₂O,
• one conjugate acid-base pair is NH₃/NH₄⁺,
• the other conjugate acid-base is H₂O/H₃O⁺.

Learning Check

A. Write the conjugate base of each of the following:
1. HBr
2. H₂S
3. H₂CO₃

B. Write the conjugate acid of each of the following:
1. NO₂⁻
2. NH₃
3. OH⁻

Solution

A. Remove one H⁺ to write the conjugate base.
1. HBr → H⁺ Br⁻
2. H₂S → H⁺ HS⁻
3. H₂CO₃ → H⁺ HCO₃⁻

B. Add one H⁺ to write the conjugate acid.
1. NO₂⁻ + H⁺ → HNO₂
2. NH₃ + H⁺ → NH₄⁺
3. OH⁻ + H⁺ → H₂O

Learning Check

Identify the sets that contain acid-base conjugate pairs.
1. HNO₂, NO₂⁻
2. H₂CO₃, CO₃²⁻
3. HCl, ClO₄⁻
4. HS⁻, H₂S
5. NH₃, NH₄⁺

Solution

Identify the sets that contain acid-base conjugate pairs.
1. HNO₂, NO₂⁻
4. HS⁻, H₂S
5. NH₃, NH₄⁺

Learning Check

A. The conjugate base of HCO₃⁻ is
1. CO₂⁻
2. HCO₃⁻
3. H₂CO₃

B. The conjugate acid of HCO₃⁻ is
1. CO₂⁻
2. HCO₃⁻
3. H₂CO₃

C. The conjugate base of H₂O is
1. OH⁻
2. H₂O
3. H₃O⁺

D. The conjugate acid of H₂O is
1. OH⁻
2. H₂O
3. H₃O⁺
Solution

A. The conjugate base of HCO₃⁻ is
   1. CO₃²⁻
B. The conjugate acid of HCO₃⁻ is
   3. H₂CO₃
C. The conjugate base of H₂O is
   1. OH⁻
D. The conjugate acid of H₂O is
   3. H₃O⁺

Strengths of Acids

• A strong acid completely ionizes (100%) in aqueous solutions.
  HCl(g) + H₂O(l) → H₃O⁺(aq) + Cl⁻(aq)

• A weak acid dissociates only slightly in water to form a few ions in aqueous solutions.
  H₂CO₃(aq) + H₂O(l) → H₃O⁺(aq) + HCO₃⁻(aq)

Weak Acids

In weak acids,
• only a few molecules dissociate.
• most of the weak acid remains as the undissociated (molecular) form of the acid.
• the concentrations of the H₃O⁺ and the anion (A⁻) are small.
  HA(aq) + H₂O(l) → H₃O⁺(aq) + A⁻(aq)

Strong Acids

In water, the dissolved molecules of a strong acid
• dissociate into ions.
• give large concentrations of H₃O⁺ and the anion (A⁻).

Strong and Weak Acids

• In an HCl solution, the strong acid HCl dissociates 100%.
• A solution of the weak acid HC₂H₃O₂ contains mostly molecules and a few ions.
Strong Acids

- Make up six of all the acids.
- Have weak conjugate bases.

<table>
<thead>
<tr>
<th>Strong Acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>Perchloric acid (HClO₄)</td>
</tr>
<tr>
<td>Sulfuric acid (H₂SO₄)</td>
</tr>
<tr>
<td>Hydrobromic acid (HBr)</td>
</tr>
<tr>
<td>Hydroiodic acid (HI)</td>
</tr>
<tr>
<td>Hydrochloric acid (HCl)</td>
</tr>
<tr>
<td>Nitric acid (HNO₃)</td>
</tr>
</tbody>
</table>

Weak Acids

- Make up most of the acids.
- Have strong conjugate bases.

Strong Bases

- Are formed from metals of Groups 1A (1) and 2A (2).
- Include LiOH, NaOH, KOH, and Ca(OH)₂.
- Dissociate completely in water.

\[ \text{KOH}(s) \rightarrow \text{K}^+(aq) + \text{OH}^-(aq) \]

Weak Bases

- Are most other bases.
- Dissociate only slightly in water.
- Form only a few ions in water.

\[ \text{NH}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq) \]

Learning Check

Identify each of the following as a strong or weak acid or base.

A. HBr
B. HNO₂
C. NaOH
D. H₂SO₄
E. Cu(OH)₂
Solution

Identify each of the following as a strong or weak acid or base.

A. HBr  strong acid
B. HNO₂  weak acid
C. NaOH  strong base
D. H₂SO₄  strong acid
E. Cu(OH)₂  weak base

Learning Check

Identify the stronger acid in each pair.

1. HNO₂ or H₂S
2. HCO₃⁻ or HBr
3. H₃PO₄ or H₂O⁺

Chapter 8  Acids and Bases

8.3 Ionization of Water

In water, H⁺ is transferred from one H₂O molecule to another. One water molecule acts as an acid, while another acts as a base.

H₂O + H₂O ⇌ H₃O⁺ + OH⁻

In pure water, the ionization of water molecules produces small, but equal quantities of H₃O⁺ and OH⁻ ions. Molar concentrations are indicated in brackets as [H₃O⁺] and [OH⁻].

[H₃O⁺] = 1.0 x 10⁻⁷ M
[OH⁻] = 1.0 x 10⁻⁷ M
Acidic Solutions

Adding an acid to pure water
• increases the \( [H_3O^+] \).
• causes the \( [H_3O^+] \) to exceed 1.0 x 10\(^{-7}\) M.
• decreases the \( [OH^-] \).

Basic Solutions

Adding a base to pure water
• increases the \( [OH^-] \).
• causes the \( [OH^-] \) to exceed 1.0 x 10\(^{-7}\) M.
• decreases the \( [H_3O^+] \).

Comparison of \( [H_3O^+] \) and \( [OH^-] \)

Ion Product of Water, \( K_w \)

The ion product constant, \( K_w \) for water
• is the product of the concentrations of the hydronium and hydroxide ions.
\[
K_w = [H_3O^+] [OH^-]
\]
• can be obtained from the concentrations in pure water.
\[
K_w = [H_3O^+] [OH^-] = (1.0 \times 10^{-7} M) \times (1.0 \times 10^{-7} M) = 1.0 \times 10^{-14}
\]

\([H_3O^+]\) and \([OH^-]\) in Solutions

In neutral, acidic, or basic solutions, the \( K_w \) is always 1.0 x 10\(^{-14}\).

<table>
<thead>
<tr>
<th>Type of Solution</th>
<th>([H_3O^+])</th>
<th>([OH^-])</th>
<th>( K_w )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutral</td>
<td>1.0 \times 10^{-7} M</td>
<td>1.0 \times 10^{-7} M</td>
<td>1.0 \times 10^{-14}</td>
</tr>
<tr>
<td>Acidic</td>
<td>1.0 \times 10^{-7} M</td>
<td>1.0 \times 10^{-7} M</td>
<td>1.0 \times 10^{-14}</td>
</tr>
<tr>
<td>Basic</td>
<td>2.5 \times 10^{-8} M</td>
<td>4.0 \times 10^{-8} M</td>
<td>1.0 \times 10^{-14}</td>
</tr>
<tr>
<td>Basic</td>
<td>1.0 \times 10^{-8} M</td>
<td>1.0 \times 10^{-8} M</td>
<td>1.0 \times 10^{-14}</td>
</tr>
<tr>
<td>Basic</td>
<td>5.0 \times 10^{-9} M</td>
<td>2.0 \times 10^{-8} M</td>
<td>1.0 \times 10^{-14}</td>
</tr>
</tbody>
</table>

Guide to Calculating \([H_3O^+]\)

1. Write the \( K_w \) constant for water.
2. Enter the \( K_w \) for the unknown \([H_3O^+]\) or \([OH^-]\) and calculate.
3. Substitute the known \([H_3O^+]\) or \([OH^-]\) and calculate.
Calculating $[\text{H}_3\text{O}^+]$

What is the $[\text{H}_3\text{O}^+]$ of a solution if $[\text{OH}^-]$ is $5.0 \times 10^{-8}$ M?

**STEP 1:** Write the $K_w$ for water.

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

**STEP 2:** Rearrange the $K_w$ expression.

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

**STEP 3:** Substitute $[\text{OH}^-]$.

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-8}} = 2.0 \times 10^{-7} \text{ M}$$

Learning Check

If lemon juice has $[\text{H}_3\text{O}^+]$ of $2 \times 10^{-3}$ M, what is the $[\text{OH}^-]$ of the solution?

1) $2 \times 10^{-11}$ M
2) $5 \times 10^{-11}$ M
3) $5 \times 10^{-12}$ M

Solution

3) $5 \times 10^{-12}$ M

Rearrange the $K_w$ to solve for $[\text{OH}^-]$

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

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{2 \times 10^{-3}} = 5 \times 10^{-12} \text{ M}$$

Learning Check

The $[\text{OH}^-]$ of an ammonia solution is $4.0 \times 10^{-2}$ M. What is the $[\text{H}_3\text{O}^+]$ of the solution?

1) $2.5 \times 10^{-11}$ M
2) $2.5 \times 10^{-12}$ M
3) $2.5 \times 10^{-13}$ M

Solution

3) $2.5 \times 10^{-13}$ M

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-2}} = 2.5 \times 10^{-13} \text{ M}$$

Chapter 8 Acids and Bases

8.4 The pH Scale
**pH Scale**

The pH of a solution

- is used to indicate the acidity of a solution.
- has values that usually range from 0 to 14.
- is acidic when the values are less than 7.
- is neutral with a pH of 7.
- is basic when the values are greater than 7.

**pH of Everyday Substances**

**Learning Check**

Identify each solution as

1) acidic  2) basic  3) neutral

A. ___ HCl with a pH = 1.5
B. ___ pancreatic fluid \([\text{H}_3\text{O}^+] = 1 \times 10^{-8} \text{ M}\)
C. ___ Sprite® soft drink  \(\text{pH} = 3.0\)
D. ___ \(\text{pH} = 7.0\)
E. ___ \([\text{OH}^-] = 3 \times 10^{-10} \text{ M}\)
F. ___ \([\text{H}_3\text{O}^+] = 5 \times 10^{-12}\)

**Solution**

A.  1  HCl with a pH = 1.5
B.  2  pancreatic fluid \([\text{H}_3\text{O}^+] = 1 \times 10^{-8} \text{ M}\)
C.  1  Sprite® soft drink  \(\text{pH} = 3.0\)
D.  3  \(\text{pH} = 7.0\)
E.  1  \([\text{OH}^-] = 3 \times 10^{-10} \text{ M}\)
F.  2  \([\text{H}_3\text{O}^+] = 5 \times 10^{-12}\)

**Testing the pH of Solutions**

The pH of solutions can be determined using

- a) pH meter.
- b) pH paper.
- c) indicators that have specific colors at different pH values.

**Calculating pH**

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

*Example:* For a solution with \([\text{H}_3\text{O}^+] = 1 \times 10^{-4}\)

\[\text{pH} = -\log [1 \times 10^{-4}]\]
\[\text{pH} = -[-4.0]\]
\[\text{pH} = 4.0\]

*Note:* The number of decimal places in the pH equals the significant figures in the coefficient of \([\text{H}_3\text{O}^+]\).
Guide to Calculating pH of an Aqueous Solution

When expressing log values, the number of decimal places in the pH is equal to the number of significant figures in the coefficient of $[\text{H}_3\text{O}^+]$.

$[\text{H}_3\text{O}^+] = 1 \times 10^{-4}$  \hspace{1cm} pH = 4.0

$[\text{H}_3\text{O}^+] = 8.0 \times 10^{-6}$  \hspace{1cm} pH = 5.10

$[\text{H}_3\text{O}^+] = 2.4 \times 10^{-8}$  \hspace{1cm} pH = 7.62

Calculating pH

Find the pH of a solution with a $[\text{H}_3\text{O}^+]$ of $1.0 \times 10^{-3}$:

**STEP 1:** Enter $[\text{H}_3\text{O}^+]$

Enter $1 \times 10^{-3}$ by pressing 1 (EE) 3
The EE key gives an exponent of 10 and change sign (+/- key or – key)

**STEP 2:** Press log key and change sign

$- \log (1 \times 10^{-3}) = -(-3)$

**STEP 3:** Adjust figures after decimal point to equal the significant figures in the coefficient.

3  $ightarrow$ 3.00  Two significant figures in $1.0 \times 10^{-3}$

Learning Check

What is the pH of coffee if the $[\text{H}_3\text{O}^+]$ is $1 \times 10^{-5}$ M?

1)  pH = 9.0  
2)  pH = 7.0  
3)  pH = 5.0

What is the pH of the solution?

1)  4.0  
2)  3.7  
3)  10.3

What is the pH of the solution?

1)  3.00  
2)  11.00  
3)  -11.00

Solution

What is the pH of coffee if the $[\text{H}_3\text{O}^+]$ is $1 \times 10^{-5}$ M?

3)  pH = 5.0

$pH = -\log [1 \times 10^{-5}] = -(-5.0) = 5.0$
Calculating $[H_3O^+]$ from pH

The $[H_3O^+]$ can be expressed by using the pH as the negative power of 10.

$$[H_3O^+] = 1 \times 10^{-\text{pH}}$$

For pH = 3.0, the $[H_3O^+] = 1 \times 10^{-3}$ M

On a calculator:
1. Enter the pH value
2. Change sign
3. Use the inverse log key (or $10^x$) to obtain the $[H_3O^+]$.

Learning Check

A. What is the $[H_3O^+]$ of a solution with a pH of 10.0?
1. $1 \times 10^{-4}$ M
2. $1 \times 10^{-10}$ M
3. $1 \times 10^{-3}$ M

B. What is the $[OH^-]$ of a solution with a pH of 2.00?
1. $1.0 \times 10^{-2}$ M
2. $1.0 \times 10^{-12}$ M
3. $2.0$ M
Acids and Metals

Acids react with metals
- such as K, Na, Ca, Mg, Al, Zn, Fe, and Sn.
- to produce hydrogen gas and the salt of the metal.

Molecular equations:

- $2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$
- $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Acids and Carbonates

Acids react
- with carbonates and hydrogen carbonates.
- to produce carbon dioxide gas, a salt, and water.

Molecular equations:

- $2HCl(aq) + CaCO_3(s) \rightarrow CO_2(g) + CaCl_2(aq) + H_2O(l)$
- $HCl(aq) + NaHCO_3(s) \rightarrow CO_2(g) + NaCl(aq) + H_2O(l)$

Learning Check

Write the products of the following reactions of acids.

A. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

B. $MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$

Solution

Write the products of the following reactions of acids.

A. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

B. $MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$

Neutralization Reactions

In a neutralization reaction
- an acid such as HCl reacts with a base such as NaOH.

- $HCl + H_2O \rightarrow H_3O^+ + Cl^−$
- $NaOH \rightarrow Na^+ + OH^−$
- the $H_3O^+$ from the acid and the $OH^−$ from the base form water.

$H_3O^+ + OH^− \rightarrow 2H_2O$

Neutralization Equations

In the equation for neutralization, an acid and a base produce a salt and water.

- $HCl + NaOH \rightarrow NaCl + H_2O$
- $2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$
Guide to Balancing an Equation for Neutralization

1. Write the reactants and products.
2. Balance the H\(^+\) in the acid with the OH\(^-\) in the base.
3. Balance the H\(^+\)O\(^-\) with the OH\(^-\) and the H\(^+\).
4. Write the salt from the remaining ions.

Write the balanced equation for the neutralization of magnesium hydroxide and nitric acid.

**STEP 1:** Write the acid and base.

\[
\text{Mg(OH)}_2 + \text{HNO}_3
\]

**STEP 2:** Balance H\(^+\) in acid with OH\(^-\) in base.

\[
\text{Mg(OH)}_2 + 2\text{HNO}_3
\]

**STEP 3:** Balance with H\(_2\)O.

\[
\text{Mg(OH)}_2 + 2\text{HNO}_3 \rightarrow \text{salt} + 2\text{H}_2\text{O}
\]

**STEP 4:** Write the salt from remaining ions.

\[
\text{Mg(NO}_3\text{)}_2 + 2\text{H}_2\text{O}
\]

Learning Check

Select the correct group of coefficients for each of the following neutralization equations.

A. HCl(aq) + Al(OH)\(_3\)(aq) \rightarrow AlCl\(_3\)(aq) + H\(_2\)O(l)
   1) 1, 3, 3, 1 2) 3, 1, 1, 1 3) 3, 1, 3

B. Ba(OH)\(_2\)(aq) + H\(_3\)PO\(_4\)(aq) \rightarrow Ba\(_3\)(PO\(_4\))\(_2\)(s) + H\(_2\)O(l)
   1) 3, 2, 2, 2 2) 3, 2, 1, 6 3) 2, 3, 1, 6

Solution

A. 3) 3, 1, 1, 3
   \[
   3\text{HCl(aq)} + \text{Al(OH)}_3\text{(aq)} \rightarrow \text{AlCl}_3\text{(aq)} + 3\text{H}_2\text{O(l)}
   \]

B. 2) 3, 2, 1, 6
   \[
   3\text{Ba(OH)}_2\text{(aq)} + 2\text{H}_3\text{PO}_4\text{(aq)} \rightarrow \text{Ba}_3\text{(PO}_4\text{)}_2\text{(s)} + 6\text{H}_2\text{O(l)}
   \]

Basic Compounds in Some Antacids

Antacids are used to neutralize stomach acid (HCl).

Learning Check

Write the neutralization reactions for stomach acid HCl and Mylanta™.
Solution

Write the neutralization reactions for stomach acid HCl and Mylanta™.

Mylanta: Al(OH)₃ and Mg(OH)₂

\[ 3\text{HCl}(aq) + \text{Al(OH)₃}(aq) \rightarrow \text{AlCl₃}(aq) + 3\text{H₂O}(l) \]

\[ 2\text{HCl}(aq) + \text{Mg(OH)₂}(aq) \rightarrow \text{MgCl₂}(aq) + 2\text{H₂O}(l) \]

Acid-Base Titration

Titration
• is a laboratory procedure used to determine the molarity of an acid.
• uses a base such as NaOH to neutralize a measured volume of an acid.

Indicator

An indicator
• is added to the acid in the flask.
• causes the solution to change color when the acid is neutralized.

End Point of Titration

At the end point,
• the indicator gives the solution a permanent pink color.
• the volume of the base used to reach the end point is measured.
• the molarity of the acid is calculated using the neutralization equation for the reaction.

Calculating Molarity

What is the molarity of an HCl solution if 18.5 mL of 0.225 M NaOH are required to neutralize 10.0 mL of HCl?

\[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H₂O}(l) \]

STEP 1: Given: 18.5 mL of 0.225 M NaOH; 10.0 mL of HCl
Need: Molarity of HCl

STEP 2: \[ 18.5 \text{ mL} \times 0.225 \text{ mole NaOH/l L} = 4.125 \text{ moles NaOH} \]
\[ 4.125 \text{ moles NaOH} \times 1 \text{ mol HCl/1 mol NaOH} = 4.125 \text{ mol HCl} \]
\[ \frac{4.125 \text{ mol HCl}}{0.0100 \text{ L HCl}} = 412.5 \text{ M HCl} \]

Calculating Molarity (continued)

STEP 4: Calculate the molarity of HCl.

\[ \frac{18.5 \text{ mL NaOH} \times 0.225 \text{ mole NaOH}}{1000 \text{ mL NaOH}} \times \frac{1 \text{ mole HCl}}{1 \text{ mole NaOH}} \]

\[ = 0.0404 \text{ mole HCl} \]

\[ M_{\text{HCl}} = \frac{0.0404 \text{ mole HCl}}{0.0100 \text{ L HCl}} = 4.04 \text{ M HCl} \]
Calculate the mL of 2.00 M H₂SO₄ required to neutralize 50.0 mL of 1.00 M KOH.

$$\text{H}_2\text{SO}_4(aq) + 2\text{KOH}(aq) \rightarrow \text{K}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)$$

1) 12.5 mL
2) 50.0 mL
3) 200. mL

**Learning Check**

A 25.0 mL sample of phosphoric acid is neutralized by 42.6 mL of 1.45 M NaOH. What is the molarity of the phosphoric acid solution?

$$3\text{NaOH}(aq) + \text{H}_3\text{PO}_4 (aq) \rightarrow \text{Na}_3\text{PO}_4(aq) + 3\text{H}_2\text{O}(l)$$

1) 0.620 M
2) 0.824 M
3) 0.185 M

**Solution**

1) 12.5 mL

\[
\frac{0.0500 \, \text{L KOH} \times 1.00 \, \text{mole KOH}}{2 \, \text{mole KOH}} \times \frac{1 \, \text{mole H}_2\text{SO}_4}{1 \, \text{L KOH}} = \frac{1 \, \text{L H}_2\text{SO}_4}{2.00 \, \text{mole H}_2\text{SO}_4} \times 1000 \, \text{mL} = 12.5 \, \text{mL}
\]

2) 0.824 M

\[
0.0426 \, \text{L} \times 1.45 \, \text{mole NaOH} \times \frac{1 \, \text{mole H}_3\text{PO}_4}{3 \, \text{mole NaOH}} = 0.0206 \, \text{mole H}_3\text{PO}_4
\]

\[
\frac{0.0206 \, \text{mole H}_3\text{PO}_4}{0.0250 \, \text{L}} = 0.824 \, \text{mole/L} = 0.824 \, \text{M}
\]

**Buffer**

When an acid or base is added

- to a buffer solution, the pH is maintained; pH does not change
Buffers

Buffers
• resist changes in pH from the addition of acid or base.
• in the body, absorb $H_3O^+$ or $OH^-$ from foods and cellular processes to maintain pH.
• are important in the proper functioning of cells and blood.
• in blood maintain a pH close to 7.4. A change in the pH of the blood affects the uptake of oxygen and cellular processes.

Buffers

Components of a Buffer

A buffer solution
• contains a combination of acid-base conjugate pairs.
• may contain a weak acid and a salt of its conjugate base.
• typically has equal concentrations of a weak acid and its salt.
• may also contain a weak base and a salt of the conjugate acid.

Buffer Action

In the acetic acid/acetate buffer with acetic acid ($HC_2H_3O_2$) and sodium acetate ($NaC_2H_3O_2$)
• the salt produces acetate ions and sodium ions.
  $NaC_2H_3O_2(aq) \rightarrow C_2H_3O_2^-(aq) + Na^+(aq)$
• the salt is added to provide a higher concentration of the conjugate base $C_2H_3O_2^-$ than the weak acid alone.
  $HC_2H_3O_2(aq) + H_2O(l) \rightarrow C_2H_3O_2^-(aq) + H_3O^+(aq)$

Function of the Weak Acid in a Buffer

The function of the weak acid in a buffer is to neutralize a base. The acetate ion produced adds to the available acetate.

Function of the Conjugate Base

The function of the acetate ion $C_2H_3O_2^-$ is to neutralize $H_3O^+$ from acids. The acetic acid produced contributes to the available weak acid.

Summary of Buffer Action

Buffer action occurs as
• the weak acid in a buffer neutralizes base.
• the conjugate base in the buffer neutralizes acid.
• the pH of the solution is maintained.
Learning Check

Which combination(s) make a buffer solution?

A. HCl and KCl
B. H\textsubscript{2}CO\textsubscript{3} and NaHCO\textsubscript{3}
C. H\textsubscript{3}PO\textsubscript{4} and NaCl
D. HC\textsubscript{2}H\textsubscript{3}O\textsubscript{2} and KC\textsubscript{2}H\textsubscript{3}O\textsubscript{2}

Solution

B. H\textsubscript{2}CO\textsubscript{3} + NaHCO\textsubscript{3}  A weak acid and its salt
D. HC\textsubscript{2}H\textsubscript{3}O\textsubscript{2} + KC\textsubscript{2}H\textsubscript{3}O\textsubscript{2}  A weak acid and its salt.