Lemons and limes are examples of foods that contain acidic solutions.

Chapter Outline

- 15.1 Acids and Bases
- 15.2 Reactions of Acids and Bases
- 15.3 Salts
- 15.4 Electrolytes and Nonelectrolytes
- 15.5 Introduction to pH
- 15.6 Neutralization
- 15.7 Writing Net Ionic Equations
- 15.8 Acid Rain

Arrhenius Acids

Arrhenius Acid: An acid solution contains an excess of H⁺ ions.

Common Properties of Acids

1. Sour taste
2. Turns litmus paper pink
3. Reacts with:
   - Metals to produce H₂ gas
   - Bases to yield water and a salt
   - Carbonates to give carbon dioxide
Arrhenius Bases

**Arrhenius Bases:** A basic solution contains an excess of \( \text{OH}^- \) ions.

**Common Properties of Bases**

1. Bitter/ caustic taste
2. Turns litmus paper blue
3. Slippery, soapy texture
4. Neutralizes acids

Brønsted-Lowry Acids and Bases

A Brønsted-Lowry acid is a proton (H\(^+\)) donor. A Brønsted-Lowry base is a proton (H\(^+\)) acceptor.

\[
\text{HCl (g)} + \text{H}_2\text{O (l)} \rightarrow 2 \text{H}_3\text{O}^+ (aq) + \text{Cl}^- (aq)
\]

**Conjugate Acid/ Base Pairs:** Two species that differ from each other by the presence of one hydrogen ion.

**Conjugate pair 1**

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

**Conjugate pair 2**

\[
\text{Acid} \quad \text{Base} \quad \text{Conjugate Acid} \quad \text{Conjugate Base}
\]

Identify the base in the following reaction and the compound’s conjugate acid.

\[
\text{H}_2\text{PO}_4^- (aq) + \text{H}_2\text{O (l)} \leftrightarrow \text{HPO}_4^{2-} (aq) + \text{H}_3\text{O}^+ (aq)
\]

**Base**

a. \( \text{H}_2\text{PO}_4^- (aq) \)

b. \( \text{H}_2\text{O (l)} \)

c. \( \text{HPO}_4^{2-} (aq) \)

d. \( \text{H}_3\text{O}^+ (aq) \)

**Conjugate Acid**

a. \( \text{H}_2\text{PO}_4^- (aq) \)

b. \( \text{H}_2\text{O (l)} \)

c. \( \text{HPO}_4^{2-} (aq) \)

d. \( \text{H}_3\text{O}^+ (aq) \)
Lewis Acid: electron pair acceptor.
Lewis Base: electron pair donor.

Slide 7

Summary of the Acid/Base Theories

The theory that best explains the reaction of interest is used.

<table>
<thead>
<tr>
<th>Theory</th>
<th>Acid</th>
<th>Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>Arrhenius</td>
<td>A hydrogen-containing substance that produces hydrogen ions in aqueous solution</td>
<td>A hydroxide-containing substance that produces hydroxide ions in aqueous solution</td>
</tr>
<tr>
<td>Brønsted-Lowry</td>
<td>A proton (H⁺) donor</td>
<td>A proton (H⁺) acceptor</td>
</tr>
<tr>
<td>Lewis</td>
<td>Any species that will bond to an unshared pair of electrons (electron pair acceptor)</td>
<td>Any species that has an unshared pair of electrons (electron pair donor)</td>
</tr>
</tbody>
</table>

Reactions of Acids

1. Reactivity with Metals
   Acids react with any metals above hydrogen in the activity series.
   \[ 2 \text{HCl (aq)} + \text{Mg (s)} \rightarrow \text{MgCl}_2 (s) + \text{H}_2 (g) \]
   In general: \ acid + metal \rightarrow \ salt + hydrogen

2. Reactivity with Bases
   Also called a neutralization reaction.
   \[ 2 \text{HCl (aq)} + \text{Ca(OH)}_2 (aq) \rightarrow \text{CaCl}_2 (aq) + 2 \text{H}_2\text{O (l)} \]
   In general: \ acid + base \rightarrow \ salt + water

© 2014 John Wiley & Sons, Inc. All rights reserved.
Reactions of Acids

3. Reactivity with Metal Oxides

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

In general:
\[ \text{acid} + \text{metal oxide} \rightarrow \text{salt} + \text{water} \]
(base)

4. Reactivity with Metal Carbonates

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

In general:
\[ \text{acid} + \text{carbonate} \rightarrow \text{salt} + \text{water} + \text{carbon dioxide} \]
(base)

Base Reactions

Bases can be amphoteric
(act as either Bröнстed acids or bases)

As a base:
\[ \text{Zn(OH)}_2 \text{ (aq)} + 2 \text{HBr (aq)} \rightarrow \text{ZnBr}_2 \text{ (aq)} + 2 \text{H}_2\text{O (l)} \]

As an acid:
\[ \text{Zn(OH)}_2 \text{ (aq)} + 2 \text{NaOH (aq)} \rightarrow \text{Na}_2\text{Zn(OH)}_4 \text{ (aq)} \]

NaOH and KOH can also react with metals.

\[ 2 \text{NaOH (aq)} + 2 \text{Al (s)} + 6 \text{H}_2\text{O (l)} \rightarrow 2 \text{NaAl(OH)}_4 \text{ (aq)} + 3 \text{H}_2 \text{(g)} \]

In general:
\[ \text{base} + \text{metal} + \text{water} \rightarrow \text{salt} + \text{hydrogen} \]

Salts

Salts: products from acid-base reactions.

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

Sodium chloride (table salt)

Salts are ionic compounds.

Salts contain a cation (a metal or ammonium ion) derived from the base and an anion (excluding oxide or hydroxide ions) derived from the acid.

Salts are generally crystalline compounds with high melting and boiling points.
Which salt forms in the reaction of aluminum oxide and hydrobromic acid?

\[ 3 \text{Al}_2\text{O}_3 (s) + 18 \text{HBr} (aq) \rightarrow 6 \text{AlBr}_3 (aq) + 9 \text{H}_2\text{O} (l) \]

- a. AlBr
- b. AlBr₃
- c. Al₂Br
- d. Al₂Br₃

Electrolytes: compounds that conduct electricity when dissolved in water. Nonelectrolytes: substances that do not conduct electricity when dissolved in water.

Comparing Solution Conductivity

<table>
<thead>
<tr>
<th>Electrolytes</th>
<th>Nonelectrolytes</th>
</tr>
</thead>
<tbody>
<tr>
<td>(Distilled water)</td>
<td>(Sugar solution)</td>
</tr>
<tr>
<td>(Salt solution)</td>
<td></td>
</tr>
</tbody>
</table>

Ion movement causes conduction of electricity in water.

3 classes of compounds, acids, bases, and salts are electrolytes because they produce ions in water when they dissolve.
Electrolytes and Nonelectrolytes Practice

Which compound will not dissociate in water?

a. HCl  
a. HCl (acid, dissociates)
b. KBr  
b. KBr (salt, dissociates)
c. NaOH  
c. NaOH (base, dissociates)
d. CH₃OH  
d. CH₃OH (organic, does not dissociate)

Dissociation of Electrolytes

Salts dissociate into their respective cations and anions when dissolved in water.

NaCl (s)  →  Na⁺ (aq) + Cl⁻ (aq)

Hydrated sodium (purple) and chloride (green) ions

The negative end of the water dipole is attracted to the positive Na⁺ ion.

When NaCl dissolves in water, each ion is surrounded by several water molecules. The permanent dipoles in the water molecules cause specific alignment around the ions.

Electrolyte Ionization

Ionization: process of ion formation in solution. Ionization results from the chemical reaction between a compound and water.

Acids ionize in water, producing the hydronium ion (H₃O⁺) and a counter anion.

HCl (g) + H₂O (l) → H₃O⁺ (aq) + Cl⁻ (aq)

H₃PO₄ (aq) + H₂O (l) ⇌ H₂PO₄⁻ (aq) + H₃O⁺ (aq)

Bases ionize in water, producing the hydroxide ion (OH⁻) and a counter cation.

NH₃ (aq) + H₂O (l) ⇌ OH⁻ (aq) + NH₄⁺ (aq)
**Strong and Weak Electrolytes**

**Strong electrolytes**: undergo complete ionization in water.
Example: HCl (strong acid)

**Weak electrolytes**: undergo incomplete ionization in water.
Example: CH$_3$COOH (weak acid)

HCl (left) is 100% ionized.
CH$_3$COOH exists primarily in the unionized form.

---

**Strong and Weak Electrolytes**

<table>
<thead>
<tr>
<th>Strong electrolytes</th>
<th>Weak electrolytes</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>H$_3$CO$_3$</td>
</tr>
<tr>
<td>H$_2$SO$_4$</td>
<td>H$_2$CO$_3$</td>
</tr>
<tr>
<td>HNO$_3$</td>
<td>HNO$_2$</td>
</tr>
<tr>
<td>HCl</td>
<td>CH$_3$COOH</td>
</tr>
<tr>
<td>HF</td>
<td>H$_2$S</td>
</tr>
</tbody>
</table>

Double arrows indicate incomplete ionization (typically weak electrolytes).

HF (aq) + H$_2$O (l) $\rightleftharpoons$ F$^-$ (aq) + H$_3$O$^+$ (aq)

---

**Salts**

Salts can dissociate into more than 2 ions, depending upon the compound.

A 1 M solution of NaCl produces a total of 2 M of ions.

NaCl (s) $\rightarrow$ Na$^+$ (aq) + Cl$^-$ (aq)

1M 1M 1M

A 1 M solution of CaCl$_2$ produces a total of 3 M of ions.

CaCl$_2$ (s) $\rightarrow$ Ca$^{2+}$ (aq) + 2 Cl$^-$ (aq)

1M 1M 2M
What is the concentration of bromide ion in a 1.5 M solution of magnesium bromide?

- a. 0.75 M
- b. 1.5 M
- c. 3.0 M
- d. 4.5 M

For every one mole of MgBr₂, 2 moles of Br⁻ ionize.

\[
\text{MgBr}_2 (s) \rightarrow \text{Mg}^{2+} (aq) + 2 \text{Br}^- (aq)
\]

Colligative Properties of Electrolyte Solutions

**Colligative properties:** depend only on the number of moles of dissolved particles present. This must be taken into consideration when calculating freezing point depression or boiling point elevation due to the presence of solute particles.

**Example:**

What is the boiling point elevation of a 1.5 m aqueous solution of CaCl₂? (\(K_b\) for water is 0.512 °C/m).

Because CaCl₂ is a strong electrolyte, 3 mol of ions (1 mol Ca²⁺ and 2 mol Cl⁻ ions) will be present in the solution.

\[
\Delta T_b = 1.5 \text{ m CaCl}_2 \times \frac{3 \text{ mol ions}}{1 \text{ mol CaCl}_2} \times \frac{0.512 \text{ °C}}{\text{m}} = 2.3 \text{ °C}
\]

Colligative Properties of Electrolytes Practice

What is the boiling point of a 2.0 m aqueous solution of NaCl? (\(K_b\) for water = 0.512 °C/m)

- a. 101.02 °C
- b. 1.02 °C
- c. 2.05 °C
- d. 102.05 °C

Because NaCl is a strong electrolyte, 2 mol of ions (1 mol Na⁺ and 1 mol Cl⁻ ions) will be present in the solution.

\[
\Delta T_b = 2.0 \text{ m NaCl} \times \frac{2 \text{ mol ions}}{1 \text{ mol NaCl}} \times \frac{0.512 \text{ °C}}{\text{m}} = 2.05 \text{ °C}
\]

\[
T_b = 100.0 \text{ °C} + 2.05 \text{ °C} = 102.05 \text{ °C}
\]
Autoionization of Water

Pure water auto(self) ionizes according to the reaction:

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

Based on the reaction stoichiometry:

Concentration \( \text{H}_3\text{O}^+ \) = Concentration \( \text{OH}^- \) = \( 1 \times 10^{-7} \) M

\[ [\text{H}_3\text{O}^+] \times [\text{OH}^-] = (1 \times 10^{-7})^2 = 1 \times 10^{-14} \]

When acid or base is present in water, \( [\text{H}_3\text{O}^+] \) and \( [\text{OH}^-] \) change.

In acidic solutions, \( [\text{H}_3\text{O}^+] > [\text{OH}^-] \).
In basic solutions, \( [\text{H}_3\text{O}^+] < [\text{OH}^-] \).

Introduction to pH

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

In pure water, \( [\text{H}_3\text{O}^+] = 1 \times 10^{-7} \) M, so

\[ \text{pH} = -\log(1 \times 10^{-7}) = 7 \]

The pH scale

<table>
<thead>
<tr>
<th>pH</th>
<th>Increasing acidity</th>
<th>Increasing basicity</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>Low HB^-</td>
<td>Neutral</td>
</tr>
<tr>
<td>1</td>
<td>Increasing HB^-</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>Neutral</td>
<td></td>
</tr>
<tr>
<td>8</td>
<td>Increasing HB^-</td>
<td></td>
</tr>
<tr>
<td>9</td>
<td></td>
<td></td>
</tr>
<tr>
<td>10</td>
<td></td>
<td></td>
</tr>
<tr>
<td>11</td>
<td></td>
<td></td>
</tr>
<tr>
<td>12</td>
<td></td>
<td></td>
</tr>
<tr>
<td>13</td>
<td></td>
<td></td>
</tr>
<tr>
<td>14</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Neutral

High HB^- Low OH-

Increasing acidity Increasing basicity

<table>
<thead>
<tr>
<th>Solution</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chicha juice</td>
<td>1.0</td>
</tr>
<tr>
<td>1 M HCl</td>
<td>1.0</td>
</tr>
<tr>
<td>Lemon juice</td>
<td>2.2</td>
</tr>
<tr>
<td>Vinegar</td>
<td>2.8</td>
</tr>
<tr>
<td>0.1 M HCl, H_2O_2</td>
<td>2.9</td>
</tr>
<tr>
<td>Orange juice</td>
<td>3.7</td>
</tr>
<tr>
<td>Tomato juice</td>
<td>4.1</td>
</tr>
<tr>
<td>Caffeine</td>
<td>5.0</td>
</tr>
<tr>
<td>Tea</td>
<td>6.0</td>
</tr>
<tr>
<td>Milk</td>
<td>6.6</td>
</tr>
<tr>
<td>Pure water (25°C)</td>
<td>7.0</td>
</tr>
<tr>
<td>Blood</td>
<td>7.4</td>
</tr>
<tr>
<td>Household ammonia</td>
<td>11.0</td>
</tr>
<tr>
<td>1 M NaOH</td>
<td>14.0</td>
</tr>
</tbody>
</table>
By what factor is a solution of pH = 3 more acidic than a solution with a pH = 5?

a. 2 

b. 20 

c. 200 

d. 100 

Factor = \frac{1 \times 10^{-3} \text{M}}{1 \times 10^{-5} \text{M}} = 100

pH = -\log[H_3O^+] 

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

Generalizations

[\text{H}_3\text{O}^+] = 1 \times 10^{-\text{pH}} \quad \text{Exponent} = \text{pH} 

[\text{H}_3\text{O}^+] = 2 \times 10^{-\text{pH}} \quad \text{Exponent} = \text{pH} 

If exactly 1

If a number between 1 and 10

The pH is between the exponent and the next lowest whole number

Calculate the pH of a 0.015 M [H_3O^+] solution.

pH = -\log(0.015) = 1.8

Note: the digits to the left of the decimal place in the pH reflect the power of ten from the [H_3O^+].

The number of decimal places for the mantissa must equal the number of significant figures in the original number.
What is the pH of a 0.020 M HCl solution?

\[
pH = -\log(0.020) = 1.7
\]

- a. 0.020
- b. -2.0
- c. 1.70
- d. 1.7

What is the pH of a NaOH solution with a \([\text{OH}^-] = 2.5 \times 10^{-11} \text{ M}\)?

\[
[H_3O^+] \times [\text{OH}^-] = 1 \times 10^{-14}
\]

\[
[H_3O^+] = \frac{1 \times 10^{-14}}{2.5 \times 10^{-11}} = 4.0 \times 10^{-4} \text{ M}
\]

\[
pH = -\log(4.0 \times 10^{-4}) = 3.4
\]

General Reaction

acid + base \rightarrow salt + water

Example

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

Overall Ionic Equation:

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

All species are included; soluble compounds shown as ions.

Net Ionic Equation:

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

Spectator ions (green) are removed from both sides.
**Titration**

**Titration:** experiment where the volume of one reagent (titrant) required to react with a measured amount of another reagent is measured. Titrations allow the amount of an acid or base present in a sample to be determined. **Indicators** are used to signal the endpoint of a titration, the point when enough titrant is added to react with the acid/base present. **Burets** deliver measured amounts of the titrant into a solution of the unknown reagent.

**Slide 34**

---

**Slide 35**

If 22.59 mL of 0.1096 M HCl is used to titrate 25.00 mL of NaOH, what is the molarity of the base?

**Reaction**

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

**Knowns**

- 22.59 mL of 0.1096 M HCl
- 25.00 mL NaOH

**Solving for**

- Molarity of base

**Calculate**

\[
\begin{align*}
22.59 \text{ mL} \times & \frac{0.1096 \text{ mol HCl}}{1000 \text{ mL soln}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} = 0.002476 \text{ mol NaOH} \\
\text{Molarity NaOH} = & \frac{0.002476 \text{ mol NaOH}}{0.02500 \text{ L soln}} = 0.09903 \text{ M NaOH}
\end{align*}
\]

**Slide 36**

What is the molarity of a NaOH solution if 21.93 mL of base is required to titrate 0.243 g of oxalic acid (H$_2$C$_2$O$_4$)?

**Reaction**

\[ \text{H}_2\text{C}_2\text{O}_4 \text{(aq)} + 2 \text{NaOH (aq)} \rightarrow \text{NaC}_2\text{O}_4 \text{(aq)} + 2 \text{H}_2\text{O (l)} \]

**Knowns**

- 0.243 g H$_2$C$_2$O$_4$
- 21.93 mL NaOH

**Solving for**

- Molarity of base

**Calculate**

\[
\begin{align*}
0.243 \text{ g H}_2\text{C}_2\text{O}_4 & \times \frac{1 \text{ mol H}_2\text{C}_2\text{O}_4}{90.04 \text{ g H}_2\text{C}_2\text{O}_4} \times \frac{2 \text{ mol NaOH}}{1 \text{ mol H}_2\text{C}_2\text{O}_4} = 0.005398 \text{ mol NaOH} \\
\text{Molarity NaOH} = & \frac{0.005398 \text{ mol NaOH}}{0.02193 \text{ L soln}} = 0.246 \text{ M NaOH}
\end{align*}
\]

*Always check reaction stoichiometry!*
What is the concentration of a nitric acid solution if 10.0 mL of the solution is neutralized by 3.6 mL of 0.20 M NaOH?

- **a. 0.072 M**
- **b. 53.6 M**
- **c. 0.56 M**
- **d. 5.6 M**

**Reaction**

\[ \text{HNO}_3(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) \]

**Knowns**

- 3.60 mL, 0.20 M NaOH
- 10.00 mL HNO₃

**Solving for Molarity of acid**

\[ \text{Molarity HNO}_3 = \frac{0.00072 \text{ mol HNO}_3}{0.01000 \text{ L soln}} = 0.072 \text{ M HNO}_3 \]

---

**Net Ionic Equations**

**Rules for Writing Net Ionic Equations**

1. **Strong** electrolytes are written as the corresponding ions. Example: NaOH (aq) is written as Na⁺(aq) + OH⁻(aq)
2. **Weak** electrolytes and nonelectrolytes are written as molecules. Example: CH₃OH(aq), CH₃COOH(aq), etc.
3. Solids and gases are written as their molecular forms.
4. The net ionic equation does not include spectator ions.
5. The net ionic equation must balance atoms and charge.

---

**Net Ionic Equations Practice**

Write the net ionic equation (NIE) for the following reaction:

\[ \text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{aq}) + 2 \text{NaCl}(\text{aq}) \]

**Total Ionic Equation**

\[ \text{Ba}^{2+}(\text{aq}) + 2 \text{Cl}^- (\text{aq}) + 2 \text{Na}^+(\text{aq}) + \text{SO}_4^{2-} (\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2 \text{Na}^+(\text{aq}) + 2 \text{Cl}^- (\text{aq}) \]

**Spectator Ions**

**Net Ionic Equation**

\[ \text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-} (\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) \]
Write the net ionic equation (NIE) for the following reaction:

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

**Total Ionic Equation**

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

**Spectator Ions**

\[ 2 \text{H}^+ (aq) + \text{CO}_3^{2-} (aq) \rightarrow \text{CO}_2 (g) + \text{H}_2\text{O} (l) \]

What is the net ionic equation when hydrobromic acid reacts with potassium hydroxide?

- a. \( \text{H}^+ (aq) + \text{Br}^- (aq) \rightarrow \text{HBr} (l) \)
- b. \( \text{K}^+ (aq) + \text{OH}^- (aq) \rightarrow \text{KOH} (s) \)
- c. \( \text{K}^+ (aq) + \text{H}^+ (aq) \rightarrow \text{KH} (s) \)
- d. \( \text{H}^+ (aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O} (l) \)

Total Ionic Equation

\[ \text{HBr} (aq) + \text{KOH} (aq) \rightarrow \text{KBr} (aq) + \text{H}_2\text{O} (l) \]

Acid Rain

**Acid rain**: atmospheric precipitation more acidic than typical.

**General Process for Acid Rain Formation**:

1. Emission of nitrogen or sulfur oxides.
2. Transportation of these chemicals throughout the atmosphere.
3. Chemical reaction of the oxides with water. This forms sulfuric and nitric acids.
4. Precipitation carries the acids to the ground.
15.1 Acids and Bases
Compare the definitions of acids and bases, including Arrhenius, Brønsted-Lowry, and Lewis acids/bases.

15.2 Reactions of Acids and Bases
Describe the general reactions of acids and bases.

15.3 Salts
Explain how a salt is formed and predict the formula of a salt given an acid and base precursor.

15.4 Electrolytes and Nonelectrolytes
Describe properties, ionization, dissociation, and strength of electrolytes and compare them to nonelectrolytes.

15.5 Introduction to pH
Calculate the pH of a solution from the hydrogen ion concentration.

15.6 Neutralization
Describe a neutralization reaction and do calculations involving titrations.

15.7 Writing Net Ionic Equations
Write net ionic equations using the stated rules.

15.8 Acid Rain
Describe how acid rain forms and the effects on society.