Ch 8: Aqueous Solutions: Chemistry of the Hydrosphere

\[ \text{H}_2\text{S} + \text{Cu}^{2+} \rightarrow \text{CuS(s)} + 2\text{H}^+ \]

(Fe, Ni, Mn also)

\[ \text{HS}^- + 2\text{O}_2 \rightarrow \text{HSO}_4^- + \text{energy} \]

(supports life)
### TABLE 8.1 Average Concentrations of 11 Major Constituents of Seawater and Human Serum

<table>
<thead>
<tr>
<th>Constituent</th>
<th>SEAWATER</th>
<th>HUMAN SERUM</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>mmol/kg</td>
<td>mM</td>
</tr>
<tr>
<td>Na(^+)</td>
<td>468.96</td>
<td>480.57</td>
</tr>
<tr>
<td>K(^+)</td>
<td>10.21</td>
<td>10.46</td>
</tr>
<tr>
<td>Mg(^2+)</td>
<td>52.83</td>
<td>54.14</td>
</tr>
<tr>
<td>Ca(^2+)</td>
<td>10.28</td>
<td>10.53</td>
</tr>
<tr>
<td>Sr(^2+)</td>
<td>0.0906</td>
<td>0.0928</td>
</tr>
<tr>
<td>Cl(^-)</td>
<td>545.88</td>
<td>559.40</td>
</tr>
<tr>
<td>SO(_4^{2-})</td>
<td>28.23</td>
<td>28.93</td>
</tr>
<tr>
<td>HCO(_3^{-})</td>
<td>2.06</td>
<td>2.11</td>
</tr>
<tr>
<td>Br(^-)</td>
<td>0.844</td>
<td>0.865</td>
</tr>
<tr>
<td>B(OH)(_3)</td>
<td>0.416</td>
<td>0.426</td>
</tr>
<tr>
<td>F(^-)</td>
<td>0.068</td>
<td>0.070</td>
</tr>
</tbody>
</table>
Chapter Outline

- **8.1 Solutions and Their Concentrations**
- 8.2 Dilutions
- 8.3 Electrolytes and Nonelectrolytes
- 8.4 Acids, Bases, and Neutralization Reactions
- 8.5 Precipitation Reactions
- 8.6 Oxidation-Reduction Reactions
- 8.7 Titrations
- 8.8 Ion Exchange
Concentration of Solutions - Molarity

The *concentration* of a solution is the amount of solute present in a given quantity of solvent or solution.

Mass-to-mass ratios: \( \text{mg/kg solvent} \)

Mass-to-volume ratios: \( \text{mg/L solvent} \)

Parts per million: \[
\frac{1 \text{ g}}{10^6 \text{ g}} = \frac{\text{mg}}{10^3 \text{ g}} = \frac{\text{mg}}{\text{L}}
\]

\[d_{H_2O} = 1.00\]
\[\text{g/mL @ 25 }^\circ\text{C}\]

\[\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}\]

---

Sample Exercise 8.1: Converting Mass-per-Volume Concentrations into Molarity

Vinyl chloride (C₂H₃Cl, MW = 62.49) is one of the most widely used industrial chemicals. It is also one of the most toxic, and it’s a known carcinogen. The maximum concentration of vinyl chloride allowed in drinking water in the U.S. is 0.002 mg/L. What is that concentration in moles per liter?

**Plan:** \( \text{mg/L} \rightarrow \text{g/L} \rightarrow \text{mol/L} = \text{Molarity} \)

\[
\frac{0.002 \text{ mg}}{\text{L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol}}{62.49 \text{ g}} = 3.2 \times 10^{-8} \text{ mol/L}
\]
Sample Exercise 8.2: Converting Mass-per-Mass Concentrations into Molarity

A water sample from the Great Salt Lake in Utah contains 83.6 mg Na\(^+\) per gram of lake water (soln). What is the molar concentration of Na\(^+\) ions if the density of the lake water is 1.160 g/mL? AW Na = 22.99

**Plan:**

\[
\text{mg Na}^+ \rightarrow g \rightarrow \text{mol} \\
\frac{1 \text{ g soln}}{\text{mL soln}} \rightarrow \frac{\text{L soln}}{d_{\text{soln}} = 1.160 \text{ g/mL}} \\
\frac{\text{mol}}{\text{L}} = M
\]

\[
83.6 \text{ mg Na}^+ \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol Na}^+}{22.99 \text{ g}} = 3.636 \times 10^{-3} \text{ mol Na}^+
\]

\[
1 \text{ g soln} \times \frac{1 \text{ mL soln}}{1.160 \text{ g}} \times \frac{1 \text{ L soln}}{1000 \text{ mL}} = 8.6221 \times 10^{-4} \text{ L}
\]

\[
M = \frac{\text{mol}}{\text{L}} = \frac{3.636 \times 10^{-3} \text{ mol Na}^+}{8.6221 \times 10^{-4} \text{ L}} = 4.22 M
\]

Sample Exercise 8.3: Calculating the Quantity of Solute Needed to Prepare a Solution

An aqueous solution called phosphate-buffered saline (PBS) is used in biology research to wash and store living cells. It contains ionic solutes including 10.0 \(mM\) \(\text{Na}_2\text{HPO}_4\cdot2\text{H}_2\text{O}\) (MW = 178.0). How many grams of this solute would you need to prepare 10.0 L of PBS?

**Plan:**

\[
\text{mM} \rightarrow \text{M} \rightarrow \text{mol} \rightarrow g \\
\text{use the volume} \\
\text{Useful equation: } M \times V = \text{moles}
\]

\[
10.0 \text{ mmol} \times \frac{1 \text{ mol}}{1000 \text{ mmol}} \times 10.0 \times \frac{178.0 \text{ g}}{\text{mol}} = 17.8 \text{ g}
\]

\[
M \times V = \text{moles}
\]
Dilution is the procedure for preparing a less concentrated solution from a more concentrated solution.

\[
\text{Moles of solute before dilution (i)} = M_i V_i = \text{Moles of solute after dilution (f)} = M_f V_f
\]
**Dilution Equation:** \( M_i V_i = M_f V_f \)

The solution used in hospitals for intravenous infusion—called physiological or normal saline—is 0.155 M NaCl. It may be prepared by diluting a commercially available standard solution that is 1.76 M NaCl. What volume of standard solution \((V_i)\) is required to prepare 10.0 L of physiological saline?

**Collect data:**

\[
egin{align*}
M_i &= 1.76 \text{ M} & M_f &= 0.155 \text{ M} \\
V_i &= ? & V_f &= 10.0 \text{ L}
\end{align*}
\]

\[
M_i V_i = M_f V_f
\]

\[
V_i = \frac{M_i V_i}{M}
\]

\[
V_i = \left( \frac{0.155 \text{ M}}{1.76 \text{ M}} \right) (10.0 \text{ L}) = 0.881 \text{ L} = 881 \text{ mL}
\]

Add 881 mL of 1.76 M NaCl, and then “dilute to the mark” (10.0 L)
Chapter Outline

- 8.1 Solutions and Their Concentrations
- 8.2 Dilutions
- **8.3 Electrolytes and Nonelectrolytes**
- 8.4 Acids, Bases, and Neutralization Reactions
- 8.5 Precipitation Reactions
- 8.6 Oxidation-Reduction Reactions
- 8.7 Titrations
- 8.8 Ion Exchange

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### How Electrolytes Conduct

![Diagram of electrolyte conductance](attachment:image.png)

- To battery
- Positive electrode
- Negative electrode
Strong Electrolytes:

- Nearly 100% dissociated into ions
- Conduct current efficiently

Examples: ionic compounds such as NaCl, strong acids like HCl

- $\text{NaCl(s)} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$
- $\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-(aq)$

*Figure from John J. Fortman, J. Chem. Ed. Vol. 71, No. 1, 1994, p. 27-28*
• **Weak Electrolytes:**
  - Only partially dissociate into ions
  - Slightly conductive
  - Example: weak acids such as acetic acid

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

*Hydronium Ion: \( \text{H}_3\text{O}^+ \)*

The higher the concentration of \( \text{H}_3\text{O}^+ \) in solution, the stronger the acid.
Non-electrolytes

- Substances in which no ionization occurs. There is no flow of electrical current through the solution

- Examples: molecular or covalent compounds such as ethanol
Chapter Outline

- 8.1 Solutions and Their Concentrations
- 8.2 Dilutions
- 8.3 Electrolytes and Nonelectrolytes
- **8.4 Acids, Bases, and Neutralization Reactions**
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- 8.7 Titrations
- 8.8 Ion Exchange
Acids

Have a sour taste. Vinegar owes its taste to acetic acid. Citrus fruits contain citric acid.

Cause color changes in plant dyes (litmus paper).

React with certain metals to produce hydrogen gas.

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

React with carbonates and bicarbonates to produce carbon dioxide gas

\[
2\text{HCl \,(aq)} + \text{CaCO}_3 \,(s) \rightarrow \text{CaCl}_2 \,(aq) + \text{CO}_2 \,(g) + \text{H}_2\text{O \,(l)}
\]

Aqueous acid solutions conduct electricity.

Limestone dissolving in acid = cave

Most metals dissolve in acids to produce \( H_2 \)
Strong Acids and Bases

- **Strong Acids/Bases:**
  - Completely ionized in aqueous solution (i.e., strong electrolytes)
  - Memorize the 6 strong acids ➔
  - All other acids are weak

**TABLE 4.3 Strong Acids**

<table>
<thead>
<tr>
<th>Acid</th>
<th>Molecular Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
</tr>
<tr>
<td>Hydrobromic acid</td>
<td>HBr</td>
</tr>
<tr>
<td>Hydroiodic acid</td>
<td>HI</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>HNO₃</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>H₂SO₄</td>
</tr>
<tr>
<td>Perchloric acid</td>
<td>HClO₄</td>
</tr>
</tbody>
</table>

**KNOW!**

**TABLE 8.2 Strong Acids and the Gases That Form Them**

<table>
<thead>
<tr>
<th>Gases</th>
<th>Acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen chloride, HCl</td>
<td>Hydrochloric acid, HCl(aq)</td>
</tr>
<tr>
<td>Hydrogen bromide, HBr</td>
<td>Hydrobromic acid, HBr(aq)</td>
</tr>
<tr>
<td>Hydrogen iodide, HI</td>
<td>Hydroiodic acid, HI(aq)</td>
</tr>
<tr>
<td>Sulfur trioxide, SO₃</td>
<td>Sulfuric acid, H₂SO₄(ℓ)</td>
</tr>
<tr>
<td>Nitrogen dioxide, NO₂</td>
<td>Nitric acid, HNO₃(ℓ)</td>
</tr>
<tr>
<td></td>
<td>Perchloric acid, HClO₄(aq)</td>
</tr>
</tbody>
</table>

*Perchloric acid is prepared by reacting concentrated solutions of sodium perchlorate and hydrochloric acid: NaClO₄(aq) + HCl(aq) ➔ HClO₄(aq) + NaCl(s).*
Bases

Have a bitter taste.
Feel slippery. Many soaps contain bases.
Cause color changes in plant dyes (litmus paper).
Aqueous base solutions conduct electricity.

A *Brønsted acid* is a proton donor
A *Brønsted base* is a proton acceptor

For every *acid* there is a *conjugate base*, and for every *base* there is a *conjugate acid* (conj. acid-base pairs).
Water: Acid or Base?

- **Water as Base:**
  \[
  \text{HCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
  \]

- **Water as Acid:**
  \[
  \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)
  \]

- **Amphiprotic:** Acts as an **Acid** or **Base**.

Preface: Double-Displacement Reactions

Many of the reactions in this chapter follow this pattern:

\[
\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}
\]
Acid-Base (Neutralization) Reactions

\[
\text{acid + base } \rightarrow \text{ salt + water}
\]

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

\[
\text{H}_2\text{SO}_4(aq) + 2\text{KOH}(aq) \rightarrow 2\text{HOH}(aq) + \text{K}_2\text{SO}_4(aq)
\]

Net Ionic Equations

1. Write the balanced molecular equation.

2. Write the ionic equation showing the strong electrolytes completely dissociated into cations and anions.

3. Cancel the spectator ions on both sides of the ionic equation.

Tips:
- Do not break apart solids, liquids (such as H₂O), and gases (e.g. CO₂)
- Do not break apart weak electrolytes (such as weak acids, e.g. HC₂H₃O₂)
- Carbonic acid, H₂CO₃, decomposes into CO₂ and H₂O
- An acid such as H₂SO₄ breaks apart into 2H⁺(aq) + SO₄²⁻, not H₂²⁺ + SO₄²⁻.
1. Write the balanced molecular equation.

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

strong \hspace{1cm} strong \hspace{1cm} Not a strong electrolyte!! \hspace{1cm} strong

2. Write the ionic equation showing the strong electrolytes completely dissociated into cations and anions.

\[ 2 \text{H}^+ + \text{SO}_4^{2-} + 2 \text{K}^+ + 2\text{OH}^- \rightarrow 2 \text{H}_2\text{O} + 2 \text{K}^+ + \text{SO}_4^{2-} \]

3. Cancel the spectator ions on both sides of the ionic equation

\[ 2 \text{H}^+ + 2\text{OH}^- \rightarrow 2 \text{H}_2\text{O} \]

Net ionic equation:

\[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \]

---

Sample Exercise 8.5: Writing the Net Ionic Equation for a Neutralization Reaction

Write the net ionic equation describing the reaction that takes place when acid rain containing sulfuric acid reacts with a marble statue (calcium carbonate).

1. Write the balanced molecular equation.

\[ \text{H}_2\text{SO}_4(aq) + \text{CaCO}_3(s) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{CaSO}_4(aq) \]

strong \hspace{1cm} solid \hspace{1cm} Not strong electrolytes!! \hspace{1cm} strong

2. Write the ionic equation showing the strong electrolytes completely dissociated into cations and anions.

\[ 2\text{H}^+(aq) + \text{SO}_4^{2-}(aq) + \text{CaCO}_3(s) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) \]

*generally speaking, CaSO₄ is insoluble in water, but not in strong acid. More on this later when we study solubility rules.
Sample Exercise 8.5: Writing the Net Ionic Equation for a Neutralization Reaction

3. Cancel the spectator ions on both sides of the ionic equation

\[
2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) + \text{Ca}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})
\]

Net ionic equation:

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

Chapter Outline

- 8.1 Solutions and Their Concentrations
- 8.2 Dilutions
- 8.3 Electrolytes and Nonelectrolytes
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- 8.6 Oxidation-Reduction Reactions
- 8.7 Titrations
- 8.8 Ion Exchange
Double-Displacement Reactions Again

Precipitation reactions in this section follow the same pattern as acid-base reactions:

\[
AB + CD \rightarrow AD + CB
\]

Types of Solutions

- **Saturated solution:**
  - Contains the maximum amount of solute that can dissolve in a given volume at a given temperature.
  - **Solubility** = g solute/100 mL solvent.

- **Supersaturated solution:**
  - Contains more dissolved solute than predicted at a given temperature.
A Saturated Solution Example

\[
\frac{3.3 \text{ g NaCl}}{100 \text{ mL}} \quad \frac{35.9 \text{ g NaCl}}{100 \text{ mL}}
\]

(a) Seawater  (b) Saturated NaCl

Supersaturated Solution

Sodium acetate precipitates from a supersaturated solution.
Precipitation Reactions

- **Precipitate:**
  - **Solid product** formed from reactants in solution.
  - \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{NaNO}_3(aq) + \text{AgCl}(s) \)

- Can predict formation of precipitates based on solubility “rules.”

We learned in Ch 4, p. 134, that the electrostatic energy (strength) of an ionic bond depended on the **product of the charges**, divided by the **internuclear distance**. The solubility of ionic compounds generally follows the trend that **the greater the charge, and the shorter the internuclear distance, the more insoluble the compound is.**

\[
E \propto \frac{Q_+ \times Q_-}{d}
\]

\( \alpha \) (alpha) means “is proportional to”
\( Q^+ = \) charge on the cation
\( Q^- = \) charge on the anion
\( d = \) internuclear distance

<table>
<thead>
<tr>
<th>Cations</th>
<th>Compounds</th>
<th>Anions</th>
<th>Compounds</th>
<th>Solubility</th>
</tr>
</thead>
<tbody>
<tr>
<td>+1</td>
<td>alkali metals, ( \text{NH}_4^+ )</td>
<td>-1</td>
<td>( \text{NO}_3^- ), ( \text{CH}_3\text{COO}^- )</td>
<td>Always soluble</td>
</tr>
<tr>
<td>+2</td>
<td></td>
<td>-2</td>
<td>( \text{SO}_4^{2-} )</td>
<td>Soluble with exceptions</td>
</tr>
<tr>
<td></td>
<td></td>
<td>-2</td>
<td>( S^2-, O^2- )</td>
<td>Insoluble with exceptions</td>
</tr>
<tr>
<td>+3</td>
<td>( \text{Al}^{3+} )</td>
<td>-3</td>
<td>( \text{PO}_4^{3-} )</td>
<td>Insoluble with exceptions</td>
</tr>
</tbody>
</table>
### Table 8.3 Solubility Rules for Ionic Compounds

<table>
<thead>
<tr>
<th>Solubility Rules</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Always soluble</strong></td>
<td>Group I cations</td>
</tr>
<tr>
<td>• Cations: group 1 ions (alkali metals) and NH$_4^+$</td>
<td></td>
</tr>
<tr>
<td>• Anions: NO$_3^-$ and CH$_3$COO$^-$ (acetate)</td>
<td></td>
</tr>
<tr>
<td><strong>Soluble w/exceptions</strong></td>
<td>SO$_4^{2-}$</td>
</tr>
<tr>
<td>• Group 17 ions (halides), except the halides of Ag$^+$, Cu$^+$, Hg$_2^{2+}$, and Pb$^{2+}$</td>
<td></td>
</tr>
<tr>
<td>• SO$_4^{2-}$, except the sulfates of Ba$^{2+}$, Ca$^{2+}$, Hg$_2^{2+}$, Pb$^{2+}$, and Sr$^{2+}$ and Ag$^+$</td>
<td>Add to table</td>
</tr>
<tr>
<td><strong>Insoluble w/exceptions</strong></td>
<td></td>
</tr>
<tr>
<td>• All hydroxides except those of group 1 cations*</td>
<td></td>
</tr>
<tr>
<td>• All sulfides except those of group 1 cations and NH$_4^+$*</td>
<td></td>
</tr>
<tr>
<td>• All carbonates except those of group 1 cations and NH$_4^+$</td>
<td></td>
</tr>
<tr>
<td>• All phosphates except those of group 1 cations and NH$_4^+$</td>
<td></td>
</tr>
</tbody>
</table>

*The solubilities of the group 2 hydroxides and sulfides increase with increasing atomic number. MgS decomposes in water, forming H$_2$S and Mg(OH)$_2$. |

---

**Table 8.3 – Always Soluble Compounds**

- Any compound containing **Group I cations**
- Any compound containing **NH$_4^+$**
- Any compound with **NO$_3^-$ or CH$_3$COO$^-$ (C$_2$H$_3$O$_2$)$^-$**
Table 8.3 - **Soluble** Compounds *With Exceptions*

- Halides (Group 17)
  - Exceptions: Cu⁺, Hg₂²⁺, Ag⁺, Pb²⁺

- Sulfates (SO₄²⁻)
  - Exceptions: Hg₂²⁺, Ag⁺, Pb²⁺, Ca²⁺, Ba²⁺, Sr²⁺

Add to table

---

Table 8.3 - **Insoluble** Compounds

- All hydroxides (OH⁻)
- All sulfides (S²⁻)
- All carbonates (CO₃²⁻)
- All phosphates (PO₄³⁻)
- All chromates (CrO₄²⁻)

**Don’t Forget:**

except IA, NH₄⁺
Mnemonic device to help remember the exceptions to the soluble compounds -

Cute HAPpy halides

Cl-, Br-, I-

Hg₂²⁺, Ag⁺, Pb²⁺

HAPpy sulfates "2"

Ca²⁺, Sr²⁺, Ba²⁺

from Group 2

Mnemonic device to help remember the insoluble compounds -

Shop for a chrome car

uyh h a
ilor r b
frs o m
iop m o
dxh a n
eia t a
st et
ee s e
ss s s

S²⁻, CrO₄²⁻, CO₃²⁻, OH⁻, PO₄³⁻
Does a precipitate form when sodium chloride is mixed with silver nitrate? If so, write the net ionic equation for the formation of the precipitate:

\[
\text{NaCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{AgCl(s)} + \text{NaNO}_3(\text{aq})
\]

Applying the Solubility Rules – Will a Precipitate Form?

\[
\text{Na}^+(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{AgCl(s)} + \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq})
\]

\[
\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl(s)}
\]

Precipitation Reactions

Formation of an Insoluble Product

\[
\text{NaI}(\text{aq}) + \text{Pb(NO}_3)^2(\text{aq}) \rightarrow
\]

The soluble compounds break up into ions and recombine -

\[
\text{K}^+(\text{aq}) \quad \text{I}^-(\text{aq})
\]

\[
\text{Pb}^{2+}(\text{aq}) \quad \text{NO}_3^-(\text{aq})
\]
A precipitate forms when aqueous solutions of ammonium sulfate and barium chloride are mixed. Write the net ionic equation for the reaction.

\[
(NH_4)_2SO_4(aq) + BaCl_2(aq) \rightarrow 2 NH_4Cl(aq) + BaSO_4(s)
\]

[Sample Exercise 8.6]

\[
2NH_4^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2Cl^-(aq) \rightarrow BaSO_4(s) + 2NH_4^+(aq) + 2Cl^-(aq)
\]

\[
Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)
\]

---

**Sample Exercise 8.7 – Prediction the Mass of a Precipitate**

Barium sulfate is used to enhance X-ray imaging of the upper and lower GI tract. In upper GI imaging, patients drink a suspension of solid barium sulfate in H_2O. The compound is not toxic because of its limited solubility. To make pure barium sulfate, a precipitation reaction is employed: aqueous solutions of soluble barium nitrate and sodium sulfate are mixed together, and solid barium sulfate is separated by filtration. How many grams of barium sulfate (MW = 233.40) will be produced if exactly 1.00 L of 1.55 M barium nitrate is reacted with excess sodium sulfate?

\[
Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2 NaNO_3(aq)
\]

**Plan:**

\[
\text{M} \times V = \text{mol Ba(NO}_3\text{)_2} \rightarrow \text{ratio} \rightarrow \text{mol BaSO}_4 \rightarrow g \text{ BaSO}_4
\]

\[
g \text{ BaSO}_4 = \frac{1.55 \text{ mol}}{1.00 \text{ L}} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba(NO}_3\text{)_2}} \times 233.40 \text{ g BaSO}_4 = 362 \text{ g BaSO}_4
\]
Sample Exercise 8.8 – Calculating Solute Concentration from Mass of Precipitate

To determine the concentration of chloride ion in a 100.0 mL sample of groundwater, a chemist adds a large enough volume of a solution of AgNO₃ to precipitate all the Cl⁻ ions as silver chloride. The mass of the resulting precipitate is 71.7 mg. What is the Cl⁻ concentration in the sample in mg/L?

Plan:

\[
\text{mg AgCl} \rightarrow \text{g AgCl} \rightarrow \text{mol AgCl} \rightarrow \text{mol Cl}^- \rightarrow \text{g Cl}^- \rightarrow \text{mg/L Cl}^-
\]

Then \( \frac{\text{mg/L Cl}^-}{\text{g Cl}^-} \) = \( \frac{\text{g Cl}^-}{\text{mg Cl}^-} \) \( \times \) \( \frac{\text{mol Cl}^-}{\text{g Cl}^-} \) \( \times \) \( \frac{\text{mol AgCl}}{\text{mol Cl}^-} \) \( \times \) \( \frac{\text{g AgCl}}{\text{mol AgCl}} \) \( \times \) \( \frac{\text{mg AgCl}}{\text{g AgCl}} \)

Divide by liters of sample

Chapter Outline

- 8.1 Solutions and Their Concentrations
- 8.2 Dilutions
- 8.3 Electrolytes and Nonelectrolytes
- 8.4 Acids, Bases, and Neutralization Reactions
- 8.5 Precipitation Reactions
- 8.6 Oxidation-Reduction Reactions
- 8.7 Titrations
- 8.8 Ion Exchange
Section 8.6 - Oxidation-Reduction Reactions

Oxidation-Reduction Reactions (Redox):
Characterized by gain or loss of electrons by atoms involved in the reaction.

Oxidation:
Historical definition = gain of oxygen (sometimes with a simultaneous loss of hydrogen)
e.g. $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$
Modern definition = loss of electrons

Reduction:
Historical definition = loss of oxygen (sometimes with a simultaneous gain in hydrogen)
e.g. $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$
Modern definition = gain of electrons
Table 8.4 - Oxidation Number Rules

The charge the atom would have in a molecule (or an ionic compound) if electrons were completely transferred, i.e. applies to both ionic and covalent compounds.

1. The O.N. for pure elements = 0 (including diatomics).

\[ \text{H} \cdot\cdot\text{H} \quad \text{O} \cdot\cdot\cdot\text{O} \quad \text{Mg} \cdot \]

2. O.N = the charge on monovalent ions

\[ \text{Mg} \cdot \rightarrow \text{Mg}^{2+} + 2 e^{-} \quad \text{Cl} \cdot + e^{-} \rightarrow \left( \text{Cl} \cdot \right)^{-} \]

3. O.N. of fluorine = -1 for all of its compounds

4. O.N. of oxygen = -2 in nearly all of its compounds

\[ \text{O} \cdot\cdot\cdot\text{O}^{2-} + 4 e^{-} \rightarrow 2 \left( \text{O} \cdot\cdot\cdot\text{O} \right)^{2-} \]

Exceptions are peroxides (O.N. = -1) and superoxides (O.N. = -1/2), e.g. H₂O₂, KO₂

5. O.N. of hydrogen = +1 in nearly all of its compounds

\[ \text{H} \cdot\cdot\text{H} \rightarrow 2 \text{H}^{+} + 2 e^{-} \]

Exception = hydrides, e.g. CaH₂

6. O.N.'s of the atoms in a neutral molecule sum up to zero

\[ \text{CaCl}_2 \quad +2 + 2(\text{Cl}) = 0 \quad \text{so Cl} = -1 \]

7. O.N.'s of the atoms in a polyatomic ion sum up to the charge on the ion

\[ \text{PO}_4^{3-} \quad \text{P} + 4(-2) = -3 \quad \text{so P} = +5 \]
Sample Exercise 8.9

What are the oxidation states for sulfur in (a) $S_8$, (b) $SO_2$, (c) $Na_2S$, and (d) $CaSO_4$?

Electron Transfer in Redox Reactions

What follows is a method that keeps track of the flow of electrons within the atoms and molecules taking part in a chemical reaction.

\[ \Delta \text{O.N.} = \text{final O.N.} - \text{initial O.N.} \]

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>Lose $n$ Electrons</th>
<th>$\Delta$ O.N. = $+n$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reduction</td>
<td>Gain $n$ Electrons</td>
<td>$\Delta$ O.N. = $-n$</td>
</tr>
</tbody>
</table>

$Pb^0 \rightarrow Pb^{2+} + 2 \text{e}^- \quad \Delta\text{O.N.} = (2 - 0) = +2$

$Cu^{2+} + 2 \text{e}^- \rightarrow Cu^0 \quad \Delta\text{O.N.} = (0 - 2) = -2$
What is the net ionic equation for the reaction of zinc metal plus cupric sulfate?
Oxidizing Agent = oxidizes something else while being reduced (Cu^{2+} \rightarrow Cu^{0})

Reducing Agent = reduces something else while being oxidized (Zn \rightarrow Zn^{2+})

Sample Exercise 8.10:
Identifying Oxidizing and Reducing Agents and Determining Number of Electrons Transferred

Energy released by the reaction of hydrazine and dinitrogen tetroxide is used to orient and maneuver spacecraft and to propel rockets into space. Identify the elements that are oxidized and reduced, the oxidizing agent, and the reducing agent, and determine the number of electrons transferred in the balanced chemical equation.

\[
2 \text{N}_2\text{H}_4(g) + \text{N}_2\text{O}_4(g) \rightarrow 3 \text{N}_2(g) + 4 \text{H}_2\text{O}(g)
\]

\[
\begin{align*}
\text{O.N. rules} & \\
\text{H} = +1 & & \text{O} = -2
\end{align*}
\]

Element in its natural state
So N = -2

2N + 4(+1) = 0
So N = +4

2N + 4(-2) = 0

2 electrons gained

2 electrons lost

\[
\Delta \text{O.N.} = +2
\] (oxidation)

\[
\Delta \text{O.N.} = -2
\] (reduction)
\[ \Delta O.N. = 0 - (-2) = +2 \]

lost electrons
oxidation

\[ \Delta O.N. = (0 - 4) = -4 \]
gained electrons
reduction

\[ 2 \text{N}_2\text{H}_4(g) + \text{N}_2\text{O}_4(g) \rightarrow 3 \text{N}_2(g) + 4 \text{H}_2\text{O}(g) \]

Net number of electrons transferred during the reaction:

\[ \Delta O.N. \text{ from previous slide} \]

\[ 2 \times 2 \times (+2) = +8 \text{ (oxidation)} \]

\[ 2 \text{N}_2\text{H}_4(g) + \text{N}_2\text{O}_4(g) \rightarrow 3 \text{N}_2(g) + 4 \text{H}_2\text{O}(g) \]

\[ 1 \times 2 \times (-4) = -8 \text{ (reduction)} \]

\[ \Delta O.N. \text{ from previous slide} \]

So the net number of electrons flowing is 8
Balancing Redox Reactions in Acid and Base
(We’re going to use a simpler method than the book called the Method of Half Reactions)

Examples of Half Reactions:
copper wire immersed in a silver nitrate solution:

\[
\begin{align*}
\text{ox: } & \quad \text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2 \text{ e}^- \\
\text{red: } & \quad \text{Ag}^+(aq) + \text{ e}^- \rightarrow \text{Ag(s)}
\end{align*}
\]

\[\Delta \text{O.N. Cu} = +2 - (0) = +2 \quad \text{oxidation}\]
\[\Delta \text{O.N. Ag} = 0 - (1) = -1 \quad \text{reduction}\]

The half-reactions are:

\[\text{Cu(s)} + 2 \text{ AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + 2 \text{ Ag(s)}\]

The next topic is how to balance a redox reaction that occurs in acid or base.
Procedure for balancing a redox reaction in acid or base.

1. Determine the oxidation numbers for the reactants and compare to the products.
2. Write the oxidation and reduction half-reactions without electrons (yet).
4. Balance oxygen by adding water.
5. Balance hydrogen by adding (a) H⁺ in acidic solutions, (b) in basic solutions, continue as if in acidic solution, but at the end each H⁺ ion will be neutralized by adding OH⁻ ions.
6. Balance charge by adding electrons; for the oxidation half-reaction, the electrons will be on the right, for the reduction half-reaction, the electrons will appear on the left.
7. Make sure the number of electrons in each half-reaction are the same. Then add the half reactions together.
8. Make sure that the equation is balanced for mass and for charge.

NOTE: sometimes you have to cancel H₂O’s that are on each side, as well as H⁺ or OH⁻.

Sample Exercise 8.11
Balancing Redox Equations I: Acidic Solutions

Microorganisms called denitrifying bacteria that grow in waterlogged soil convert NO₃⁻ ions into N₂O gas as they feed on dead plant tissue (empirical formula = CH₂O), converting it into CO₂ and H₂O. Write a balanced net ionic equation describing this conversion of dissolved nitrates to N₂O gas. Assume the reaction occurs in slightly acidic water.

1. Determine the oxidation numbers for the reactants and compare to the products.

\[
\begin{align*}
\text{NO}_3^- (aq) & \rightarrow \text{N}_2\text{O}(g) + \text{CO}_2(g) + \text{H}_2\text{O}(l) \\
\text{Given:} & \text{ C = } +4 - (0) = +4 \quad \text{oxidation} \\
& \text{ N = } 1 - (5) = -4 \quad \text{reduction}
\end{align*}
\]
2. Write the oxidation and reduction half-reactions *without electrons (yet)*

   ox: \( \text{CH}_2\text{O(s)} \rightarrow \text{CO}_2\text{(g)} \)

   red: \( \text{NO}_3\text{-(aq)} \rightarrow \text{N}_2\text{O(g)} \)

3. Balance everything but oxygen and hydrogen

   ox: \( \text{CH}_2\text{O(s)} \rightarrow \text{CO}_2\text{(g)} \)

   red: \( 2 \text{NO}_3\text{-(aq)} \rightarrow \text{N}_2\text{O(g)} \)

4. Balance oxygen by adding water

   ox: \( \text{H}_2\text{O} + \text{CH}_2\text{O(s)} \rightarrow \text{CO}_2\text{(g)} \)

   red: \( 2 \text{NO}_3\text{-(aq)} \rightarrow \text{N}_2\text{O(g)} + 5 \text{H}_2\text{O} \)

5. Balance hydrogen by adding (a) \( \text{H}^+ \) in acidic solutions, (b) in basic solutions, continue as if in acidic solution, but at the end each \( \text{H}^+ \) ion will be neutralized by adding \( \text{OH}^- \) ions

   ox: \( \text{H}_2\text{O} + \text{CH}_2\text{O(s)} \rightarrow \text{CO}_2\text{(g)} + 4 \text{H}^+(aq) \)

   red: \( 10 \text{H}^+(aq) + 2 \text{NO}_3\text{-(aq)} \rightarrow \text{N}_2\text{O(g)} + 5 \text{H}_2\text{O} \)

6. Balance charge by adding electrons; for the oxidation half-reaction, the electrons will be on the right, for the reduction half-reaction, the electrons will appear on the left

   ox: \( \text{H}_2\text{O} + \text{CH}_2\text{O(s)} \rightarrow \text{CO}_2\text{(g)} + 4 \text{H}^+(aq) + 4 \text{e}^- \)

   red: \( 8 \text{e}^- + 10 \text{H}^+(aq) + 2 \text{NO}_3\text{-(aq)} \rightarrow \text{N}_2\text{O(g)} + 5 \text{H}_2\text{O} \)
7. Make sure the number of electrons in each half-reaction are the same. Then add the half reactions together

\[
\text{ox: } \left[ \text{H}_2\text{O} + \text{CH}_2\text{O(s)} \rightarrow \text{CO}_2(g) + 4 \text{ H}^+(aq) + 4 \text{ e}^- \right] \times 2 \\
\text{red: } 8 \text{ e}^- + 10 \text{ H}^+(aq) + 2 \text{ NO}_3^-(aq) \rightarrow \text{N}_2\text{O(g)} + 5 \text{ H}_2\text{O}
\]

\text{Net: } 2\text{H}_2\text{O} + 2\text{CH}_2\text{O(s)} + 10\text{H}^+(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{CO}_2(g) + 8\text{H}^+(aq) + \text{N}_2\text{O(g)} + 5\text{H}_2\text{O}

\textbf{NOTE:} sometimes you have to cancel H_2O's that are on each side, as well as H^+ or OH^-

\[
2\text{H}_2\text{O} + 2\text{CH}_2\text{O(s)} + 10\text{H}^+(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{CO}_2(g) + 8\text{H}^+(aq) + \text{N}_2\text{O(g)} + 5\text{H}_2\text{O} \\
2\text{CH}_2\text{O(s)} + 10\text{H}^+(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{CO}_2(g) + 8\text{H}^+(aq) + \text{N}_2\text{O(g)} + 3\text{H}_2\text{O} \\
2\text{CH}_2\text{O(s)} + 2\text{H}^+(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{CO}_2(g) + \text{N}_2\text{O(g)} + \text{3H}_2\text{O}
\]

8. Make sure that the equation is balanced for mass and for charge

\[
2\text{CH}_2\text{O(s)} + 2\text{H}^+(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{CO}_2(g) + \text{N}_2\text{O(g)} + \text{3H}_2\text{O}
\]

\begin{align*}
\text{Mass:} & \quad C = 2 & \quad C = 2 \\
& \quad H = 6 & \quad H = 6 \\
& \quad O = 8 & \quad O = 8 \\
& \quad N = 2 & \quad N = 2 \\
\text{Charge:} & \quad 2(+1) + 2(-1) = 0 & \quad = 0
\end{align*}
Permanganate ion is used in water purification to remove oxidizable substances. Complete and balance the reaction for the removal of sulfite ion.

1. Determine the oxidation numbers for the reactants and compare to the products.

\[
\begin{align*}
\text{MnO}_4^- (aq) & \quad \text{SO}_3^{2-} (aq) \quad \text{MnO}_2(s) + \text{SO}_4^{2-}(aq) \\
(+7) & \quad (-2) \quad (+4) \quad (-2) \quad (+6) \quad (-2)
\end{align*}
\]

\[\Delta \text{O.N. S} = +6 - (+4) = +2 \quad \text{oxidation}\]
\[\Delta \text{O.N. Mn} = +4 - (+7) = -3 \quad \text{reduction}\]

2. Write the oxidation and reduction half-reactions without electrons (yet)

\[\begin{align*}
\text{ox:} & \quad \text{SO}_3^{2-} (aq) \rightarrow \text{SO}_4^{2-} (aq) \\
\text{red:} & \quad \text{MnO}_4^- (aq) \rightarrow \text{MnO}_2(s)
\end{align*}\]

3. Balance everything but oxygen and hydrogen

\[\begin{align*}
\text{ox:} & \quad \text{SO}_3^{2-} (aq) \rightarrow \text{SO}_4^{2-} (aq) \\
\text{red:} & \quad \text{MnO}_4^- (aq) \rightarrow \text{MnO}_2(s) \\
\quad & \quad \text{S and Mn are already balanced}
\end{align*}\]

4. Balance oxygen by adding water

\[\begin{align*}
\text{ox:} & \quad \text{H}_2\text{O} + \text{SO}_3^{2-} (aq) \rightarrow \text{SO}_4^{2-} (aq) \\
\text{red:} & \quad \text{MnO}_4^- (aq) \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}
\end{align*}\]

5. Balance hydrogen by adding (a) H⁺ in acidic solutions, (b) in basic solutions, continue as if in acidic solution, but at the end each H⁺ ion will be neutralized by adding OH⁻ ions

\[\begin{align*}
\text{ox:} & \quad \text{H}_2\text{O} + \text{SO}_3^{2-} (aq) \rightarrow \text{SO}_4^{2-} (aq) + 2\text{H}^+ (aq) \\
\text{red:} & \quad 4\text{H}^+ (aq) + \text{MnO}_4^- (aq) \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}
\end{align*}\]
6. Balance charge by adding electrons; for the oxidation half-reaction, the electrons will be on the right, for the reduction half-reaction, the electrons will appear on the left.

ox: \[ \text{H}_2\text{O} + \text{SO}_3^{2-}(\text{aq}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2 \text{ e}^- \]

red: \[ 3 \text{ e}^- + 4\text{H}^+(\text{aq}) + \text{MnO}_4^-(\text{aq}) \rightarrow \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O} \]

7. Make sure the number of electrons in each half-reaction are the same. Then add the half reactions together.

ox: \[ \left[ \text{H}_2\text{O} + \text{SO}_3^{2-}(\text{aq}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2 \text{ e}^- \right] \times 3 \]

red: \[ 3 \text{ e}^- + 4\text{H}^+(\text{aq}) + \text{MnO}_4^-(\text{aq}) \rightarrow \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O} \] \times 2

Net: \[ 3\text{H}_2\text{O} + 3\text{SO}_3^{2-}(\text{aq}) + 8\text{H}^+(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) \rightarrow 3\text{SO}_4^{2-}(\text{aq}) + 6\text{H}^+(\text{aq}) + 2\text{MnO}_2(\text{s}) + 4\text{H}_2\text{O} \]

**NOTE:** sometimes you have to cancel H\text{H}_2\text{O}'s that are on each side, as well as H\text{H}^+ or OH^-.

\[ 3\text{SO}_3^{2-}(\text{aq}) + 8\text{H}^+(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) \rightarrow 3\text{SO}_4^{2-}(\text{aq}) + 6\text{H}^+(\text{aq}) + 2\text{MnO}_2(\text{s}) + 4\text{H}_2\text{O} \]

\[ 3\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{MnO}_4^-(\text{aq}) \rightarrow 3\text{SO}_4^{2-}(\text{aq}) + 2\text{MnO}_2(\text{s}) + 2\text{OH}^- \]
8. Make sure that the equation is balanced for mass and for charge

\[ 3\text{SO}_3^2-(aq) + \text{H}_2\text{O}(l) + 2\text{MnO}_4^-(aq) \rightarrow 3\text{SO}_4^{2-}(aq) + 2\text{MnO}_2(s) + 2\text{OH}^-(aq) \]

**Mass:**
- \( S = 3 \)
- \( O = 18 \)
- \( H = 2 \)
- \( \text{Mn} = 2 \)

**Charge:**
- \( 3(-2) + 2(-1) = -8 \)

✔️

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Titrations

In a titration a solution of accurately known concentration is added gradually to another solution of unknown concentration until the chemical reaction between the two solutions is complete.

**Equivalence point** – the point at which the reaction is complete

**Indicator** – substance that changes color at (or near) the equivalence point

Slowly add base to unknown acid UNTIL the indicator changes color

**Titration Example, p. 344-345**

\[
\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}
\]

(titrant)

(a) \[\text{H}_2\text{SO}_4 \quad \text{(unknown)}\]

(b) End Point
Suppose it takes 22.40 mL of 0.00100 M sodium hydroxide to react completely with the sulfuric acid in a 100.0 mL sample of acidic mine drainage. What was the molarity of the sulfuric acid?

\[
2\text{NaOH(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4
\]

\[
\begin{align*}
0.00100 \text{ M} & \quad 100.0 \text{ mL} \\
22.40 \text{ mL} & = 0.1000 \text{ L} \\
= 0.02240 \text{ L} & \quad M = \, ? \\
\hline
\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} & \quad \text{ Plan: } M \times V = \text{ mol NaOH} \rightarrow \text{ mol H}_2\text{SO}_4 \rightarrow M \\
\text{ mol H}_2\text{SO}_4 & = 0.00100 \text{ mol NaOH} \times 0.02240 \text{ L NaOH} \times 1 \text{ mol H}_2\text{SO}_4 \\
& \quad 2 \text{ mol NaOH} \\
& = 1.120 \times 10^{-5} \text{ mol} \\
M & = \frac{1.120 \times 10^{-5} \text{ mol}}{0.1000 \text{ L}} = 1.120 \times 10^{-4} \text{ M} \quad \checkmark
\]

**Sample Exercise 8.13:**

Apple cider vinegar is an aqueous solution of acetic acid (MW = 60.05) made from fermented apple juice. Commercial vinegar must contain at least 4 grams of acetic acid per 100 mL of vinegar. Suppose the titration of a 25.00 mL sample of vinegar requires 12.15 mL of 1.885 M sodium hydroxide. What is the molarity of acetic acid in the vinegar sample? Does it meet the 4 g/100 mL standard?

\[
\text{HC}_2\text{H}_3\text{O}_2\text{(aq)} + \text{NaOH(aq)} \rightarrow \text{H}_2\text{O} + \text{NaC}_2\text{H}_3\text{O}_2
\]

\[
\begin{align*}
25.00 \text{ mL} & \quad 12.15 \text{ mL} \\
M & = \, ? \\
= 0.01215 \text{ L} & \quad M = 1.885 \\
\text{mg/L} & = \, ?
\hline
\frac{1 \text{ mol HC}_2\text{H}_3\text{O}_2}{\text{mol NaOH}} & \quad \text{ Plan: } M \times V = \text{ mol NaOH} \rightarrow \text{ mol HC}_2\text{H}_3\text{O}_2 \rightarrow M \rightarrow \text{g/mL} \\
\text{ mol HC}_2\text{H}_3\text{O}_2 & = 1.885 \text{ mol NaOH} \times 0.01215 \text{ L NaOH} \times 1 \text{ mol HC}_2\text{H}_3\text{O}_2 \times \frac{1}{1 \text{ mol NaOH}} \\
& \quad \times \frac{1}{0.02500 \text{ L}} \\
& = 0.9161 \text{ M} \quad \checkmark
\]

\[
g = \frac{0.9161 \text{ mol}}{1 \text{ mol}} \times 60.05 \text{ g} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.05501 \text{ g/mL} = 5.501 \text{ g/100 mL} \quad \checkmark
\]