Reactions in Aqueous Solutions
Chapter 4

- General Properties of Aqueous Solutions (4.1)
- Precipitation Reactions (4.2)
- Acid-Base Reactions (4.3)
- Oxidation-Reduction Reactions (4.4)
- Concentration of Solutions (4.5)
- Solution Stoichiometry (4.6)
Chapter 4

• Definition of solution (4.1)
  ◦ Solvent, solute, concentration
  ◦ Electrolyte vs. nonelectrolyte

• Reactions
  ◦ Precipitation reactions (4.2)
  ◦ Acid / Base reactions (4.3)
  ◦ Redox reactions (4.4)

• Calculations of solutions (4.5)
  ◦ Determination of concentration, mass, volume
  ◦ Dilutions

• Stoichiometry with solutions (4.6)
4.1 General Properties of Aqueous Solutions

- **Key Definitions:**
  - **Solution** *(review)*:
    - Homogeneous mixture of two or more pure substances
  - **Solvent**:
    - Substance present in larger amounts
  - **Solute**:
    - Substance present in smaller amounts
4.1 General Properties of Aqueous Solutions

• Key Definitions:
  ◦ Solvation:
    • Process of forming a solution through the interaction of the solvent with the solute
  ◦ Hydration:
    • Process of forming a solution through the interaction of \textit{water} with the solute
    • Process in which an ion is surrounded by water molecules arranged in a specific manner
4.1 General Properties of Aqueous Solutions

- **Review:** What are the three ways to represent matter?
  - Macroscopic
  - Symbolic
  - Particle

- **Review:** What are the three representations for water as a liquid?

- **Review:** How is the reaction of hydrogen and oxygen to make water represented?
4.1 General Properties of Aqueous Solutions

- Draw one molecule of water
  - Identify any uneven charge distribution on the molecule showing this with a “+” and “−” sign
  - Add another water molecule and show how these would interact

- Review: In an ionic compound, what kind of charge is on a cation and an anion?
Combining the water molecules you have drawn:
- to what part of a water molecule would a **cation** be attracted?
- to what part of a water molecule would an **anion** be attracted?

What does this look like for sodium chloride? How is this consistent with the definition that sodium chloride is **soluble**?
4.1 General Properties of Aqueous Solutions

Figure 4.2, p. 99
4.1 General Properties of Aqueous Solutions

- What is the balanced equation for sodium chloride in water?
- What does this look like on the macroscopic level?
- If silver chloride is insoluble in water, how are these three representations different?
- Will methanol dissolve in water? How is this different from sodium chloride or silver chloride?
4.1 General Properties of Aqueous Solutions

- **Key Definitions:**
  - **Electrolyte:**
    - Substance that, when dissolved in water, results in a solution that can conduct electricity
  - **Non-electrolyte:**
    - Substance that, when dissolved in water, results in a solution that does not conduct electricity

Figure 4.1, p. 98
4.1 General Properties of Aqueous Solutions

<table>
<thead>
<tr>
<th>Strong Electrolyte</th>
<th>Weak Electrolyte</th>
<th>Nonelectrolyte</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>CH₃COOH</td>
<td>(NH₂)₂CO (urea)</td>
</tr>
<tr>
<td>HNO₃</td>
<td>HF</td>
<td>CH₃OH (methanol)</td>
</tr>
<tr>
<td>HClO₄</td>
<td>HNO₂</td>
<td>C₂H₅OH (ethanol)</td>
</tr>
<tr>
<td>H₂SO₄*</td>
<td>NH₃</td>
<td>C₆H₁₂O₆ (glucose)</td>
</tr>
<tr>
<td>NaOH</td>
<td>H₂O†</td>
<td>C₁₂H₂₂O₁₁ (sucrose)</td>
</tr>
<tr>
<td>Ba(OH)₂</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*H₂SO₄ has two ionizable H⁺ ions.
†Pure water is an extremely weak electrolyte.

Table 4.1, p. 99
4.2 Precipitation Reactions

- **Review: Key Definitions:**
  - **Soluble:**
    - Dissolves in solution
    - Separates into ions in solution
  - **Insoluble**
    - Does not dissolve in solution
    - Does not separate into ions in solution
  - **Precipitate**
    - An insoluble solid that separates from the solution
# TABLE 4.2: Solubility Rules for Common Ionic Compounds in Water at 25°C

<table>
<thead>
<tr>
<th>Soluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing alkali metal ions (Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺) and the ammonium ion (NH₄⁺)</td>
<td>Halides of Ag⁺, Hg₂²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>Nitrates (NO₃⁻), bicarbonates (HCO₃⁻), and chlorates (ClO₃⁻)</td>
<td>Sulfates of Ag⁺, Ca²⁺, Sr²⁺, Ba²⁺, Hg²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>Halides (Cl⁻, Br⁻, I⁻)</td>
<td></td>
</tr>
<tr>
<td>Sulfates (SO₄²⁻)</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Insoluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbonates (CO₃²⁻), phosphates (PO₄³⁻), chromates (CrO₄²⁻), and sulfides (S²⁻)</td>
<td>Compounds containing alkali metal ions and the ammonium ion</td>
</tr>
<tr>
<td>Hydroxides (OH⁻)</td>
<td>Compounds containing alkali metal ions and the Ba²⁺ ion</td>
</tr>
</tbody>
</table>

Table 4.2, p. 101
4.2 Precipitation Reactions

• What is the solubility of:
  ◦ Na$_2$CO$_3$
  ◦ PbCl$_2$
  ◦ Al(OH)$_3$
  ◦ Ca(HCO$_3$)$_2$
  ◦ Fe(NO$_3$)$_3$
  ◦ SrSO$_4$

• Which macroscopic representation is best for each compound above?
4.2 Precipitation Reactions

- How are soluble ionic compounds represented on the particle level? 
  Fe(NO₃)₃(aq) and Na₂CO₃(aq)
- When these compounds are combined, what new compounds can form?
  ◦ What is the solubility of these compounds?
- Show this reaction on the three levels.
4.2 Precipitation Reactions

- What happens when two ionic compounds which are soluble are combined?
  - Fe(NO$_3$)$_3$(aq) and Na$_2$CO$_3$(aq)

- What is the molecular equation?
- What is the total ionic equation?
- What are the spectator ions?
- What is the net ionic equation?
4.2 Precipitation Reactions

Figure 4.3, p. 101
4.2 Precipitation Reactions
More practice:

- What happens $\text{AgNO}_3$ and $\text{KOH}$ are combined in water?
- What happens $\text{MgBr}_2$ and $\text{Fe}_2(\text{SO}_4)_3$ are combined in water?
- Does a reaction occur?
  - If so, what are the molecular equation, total ionic equation, spectator ions, net ionic equation
  - If not, write no reaction
More practice:

- What would precipitate only lead in a solution of Fe$^{3+}$, Cu$^{+}$, Pb$^{2+}$ and Ba$^{2+}$?
- What is the molecular, total ionic and net ionic equation for the combination of sodium hydroxide with:
  - Silver nitrate
  - Iron(III) sulfate
  - Barium chloride
Chapter 4 Practice

An partial illustration showing the hydration of a soluble ionic compound is shown in the figure. What is the designation of the ion in the figure?

A. = the anion
B. = the cation
C. = the anion
D. = the cation

What are the spectator ions for the precipitation reaction shown?

\[ 2\text{Fe(NO}_3\text{)}_3\text{ (aq) + 3K}_2\text{CO}_3\text{ (aq) → Fe}_2\text{(CO}_3\text{)}_3\text{ (s) + 6KNO}_3\text{ (aq)} \]

A. \( \text{K}^+ \) and \( \text{NO}_3^- \)  B. \( \text{K}^+ \) and \( \text{CO}_3^{2-} \)  C. \( \text{Fe}^{3+} \) and \( \text{NO}_3^- \)  D. \( \text{Fe}^{3+} \) and \( \text{CO}_3^{2-} \)
Chapter 4 Practice

In lab, you selectively precipitated out nine cations and three anions in order to positively identify each. What could you add to a solution of $\text{Ba}^{2+}$, $\text{Ca}^{2+}$, $\text{Fe}^{3+}$, $\text{Mn}^{2+}$, and $\text{Al}^{3+}$ to separate $\text{Ba}^{2+}$ and $\text{Ca}^{2+}$ from $\text{Fe}^{3+}$, $\text{Mn}^{2+}$ and $\text{Al}^{3+}$?

A. $\text{Na}_2\text{CO}_3$  
B. $\text{NaNO}_3$  
C. $\text{NaOH}$  
D. $\text{Na}_2\text{SO}_4$
4.3 Acid-Base Reactions

- **Key Properties:**
  - **Acid:**
    - sour taste, change litmus from blue to red, react with certain metals, strong or weak electrolytes
  - **Base:**
    - bitter taste, feel slippery, change litmus from red to blue, strong or weak electrolytes
4.3 Acid-Base Reactions

- **Key Definitions:**
  - **Acid:**
    - *Arrhenius*: a substance that increases the $H^+$ concentration in water
    - *Brønsted-Lowry*: a substance that is a $H^+$ donor
  - **Base**
    - *Arrhenius*: a substance that increases the $OH^-$ concentration in water
    - *Brønsted-Lowry*: a substance that is a $H^+$ acceptor
4.3 Acid-Base Reactions

\[
\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-
\]

\[
\text{NH}_3 + \text{H}_2\text{O} \leftrightarrow \text{NH}_4^+ + \text{OH}^-
\]
4.3 Acid-Base Reactions

• Do hydrogen ions exist in water?

Margin Figure, p. 106
4.3 Acid-Base Reactions

- What is an acid?
- What is a base?
- What is strong?
- What is weak?
4.3 Acid-Base Reactions

- Weak Acid
- Weak base
4.3 Acid-Base Reactions

- What happens when an acid and base combine?

![Diagram showing acid-base reactions with ions and molecules]
4.3 Acid-Base Reactions

- What happens when a strong acid and a strong base combine?

- What is the net ionic equation and the spectator ions?
4.3 Acid-Base Reactions

- What happens when a weak acid and a strong base combine?

- What is the net ionic equation and the spectator ion?
What happens when a **strong** acid and a **weak** base combine?

What is the net ionic equation and the spectator ion?
Do these reactions always result in all species in solution?

- Sulfuric acid and barium hydroxide
- A precipitate can form between the “spectator ions” – \( \text{BaSO}_4 \)

Figure 4.5, p. 103
Do these reactions always result in all species in solution?

- Sodium bicarbonate and acetic acid
  - A gas can evolve – CO$_2$
4.3 Acid-Base Reactions

• Practice:
  ◦ What is the molecular, total ionic and net ionic equation for the combination of sodium hydroxide (a strong base) with:
    • Perchloric acid (a strong acid)
    • Hypochlorous acid (a weak acid)
Chapter 4 Practice

What is the net ionic equation for the reaction of a hydrocyanic acid (HCN, a weak acid) and potassium hydroxide (KOH, a strong base)?

A. HCN(aq) + KOH(aq) → H₂O(l) + KCN(aq)
B. HCN(aq) + OH⁻(aq) → H₂O(l) + CN⁻(aq)
C. CN⁻(aq) + K⁺(aq) → KCN(aq)
D. H⁺(aq) + OH⁻(aq) → H₂O(l)
4.4 Oxidation-Reduction Reactions

• Key Definitions:
  ◦ Oxidation-reduction reactions (redox):
    • Electron-transfer reactions
  ◦ Oxidation:
    • Process of losing electrons
    • Oxidation reaction: half reaction that involves loss of electrons
  ◦ Reduction:
    • Process of gaining electrons
    • Reduction reaction: half reaction that involves gain of electrons
4.4 Oxidation-Reduction Reactions

- Reactions of electron transfer
- Electrons are transferred between
  - Species which loses electrons
    - Species is oxidized
  - Species which gains electrons
    - Species is reduced
Think about the formation of NaCl

- Na atoms become sodium ions
  Loses 1 electron
  \[
  \text{Na} \rightarrow \text{Na}^+ + 1\text{e}^-
  \]

- Cl atoms become chlorine ions
  Gains 1 electron
  \[
  \text{Cl} + 1\text{e}^- \rightarrow \text{Cl}^-
  \]
Think about the formation of NaCl

- Na atoms become sodium ions
  - Loses 1 electron
  - \( \text{Na} \rightarrow \text{Na}^+ + 1\text{e}^- \)

- Cl atoms become chlorine ions
  - Gains 1 electron
  - \( \text{Cl} + 1\text{e}^- \rightarrow \text{Cl}^- \)
4.4 Oxidation-Reduction Reactions

- Think about the formation of NaCl
  - Na atoms become sodium ions
    - Loses 1 electron
    - \( \text{Na} \rightarrow \text{Na}^+ + 1\text{e}^- \)
  - Cl molecules become chlorine ions
    - Gains 2 electron
    - \( \text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^- \)
4.4 Oxidation-Reduction Reactions

- Overall:
  \[2 \, (\text{Na} \rightarrow \text{Na}^+ + 1\, \text{e}^-) \text{ oxidation rxn} \]
  \[\text{Cl}_2 + 2\, \text{e}^- \rightarrow 2\, \text{Cl}^- \text{ reduction rxn}\]

\[2\, \text{Na} + \text{Cl}_2 \rightarrow 2\, \text{NaCl}\]

oxidized \quad reduced
4.4 Oxidation-Reduction Reactions

Figure 4.9, p. 111
4.4 Oxidation-Reduction Reactions

- Determining oxidation numbers:
  - For monoatomic species:
    - The oxidation state is equal to the charge of the species
    - Na$^+$ (ions)
    - Al$^{3+}$
    - Cl$_2$ (free elements)
    - Fe
4.4 Oxidation-Reduction Reactions

- Determining oxidation numbers:
  - For polyatomic species:
    - Assume complete ionic character
      - Oxygen will usually have a -2 ox state.
      - Halides often have a -1 ox state with fluorine having a -1 ox state. When halides are combined with oxygen – oxygen will have a -2 ox state and the halide will have a positive ox state.
      - Hydrogen will have +1 ox state except with metals (metal hydrides) where it will have -1 ox state.
4.4 Oxidation-Reduction Reactions

- Determining oxidation numbers:
  - Overall:
  - The sum of the oxidation state for all species in the compound must sum to zero for a neutral compound or molecule or to the net charge on a polyatomic ion
  - Oxidation states can be simple fractions
Figure 4.10, p. 114
4.4 Oxidation-Reduction Reactions

- Determining oxidation numbers:
  - Practice:
    - KNO$_3$
    - H$_2$SO$_4$
    - FeCr$_2$O$_7$
    - AlPO$_4$
    - H$_2$
    - H$_2$O
4.4 Oxidation-Reduction Reactions

- Combination reactions
  - Formation of NaCl
  - Formation of CO$_2$

Margin figure, p. 114
4.4 Oxidation-Reduction Reactions

- Decomposition reactions
  - Reaction $\text{KClO}_3$ and heat
    $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$

Margin figure, p. 114
4.4 Oxidation-Reduction Reactions

- Combustion reactions
  - Methane
    \[ CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \]
4.4 Oxidation-Reduction Reactions

- Displacement Reactions
  - Hydrogen displacement
  - Metal displacement
  - Halogen displacement

May be familiar as “single displacement reactions” (where you used the activity series)
4.4 Oxidation-Reduction Reactions

- Hydrogen displacement reactions
- Reaction of metals with water

2Na(s) + 2H₂O(l) → 2NaOH(aq) + H₂(g)

Figure 4.11, p. 115
4.4 Oxidation-Reduction Reactions

- Hydrogen displacement reactions
  - Reaction of metals with acid

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

Figure 4.12, p. 116
4.4 Oxidation-Reduction Reactions

- Hydrogen displacement reactions
  - Will any metal which reacts with water also react with acid?
  - Will any metal which reacts with acid also react with water?

- p. 118, Figure 4.14
<table>
<thead>
<tr>
<th>Element</th>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li → Li⁺ + e⁻</td>
<td>React with cold water to produce H₂</td>
</tr>
<tr>
<td>K → K⁺ + e⁻</td>
<td></td>
</tr>
<tr>
<td>Ba → Ba²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Ca → Ca²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Na → Na⁺ + e⁻</td>
<td></td>
</tr>
<tr>
<td>Mg → Mg²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Al → Al³⁺ + 3e⁻</td>
<td></td>
</tr>
<tr>
<td>Zn → Zn²⁺ + 2e⁻</td>
<td>React with steam to produce H₂</td>
</tr>
<tr>
<td>Cr → Cr³⁺ + 3e⁻</td>
<td></td>
</tr>
<tr>
<td>Fe → Fe²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Cd → Cd²⁺ + 2e⁻</td>
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</tr>
<tr>
<td>Co → Co²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Ni → Ni²⁺ + 2e⁻</td>
<td>React with acids to produce H₂</td>
</tr>
<tr>
<td>Sn → Sn²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Pb → Pb²⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>H₂ → 2H⁺ + 2e⁻</td>
<td></td>
</tr>
<tr>
<td>Cu → Cu²⁺ + 2e⁻</td>
<td>Do not react with water or acids to produce H₂</td>
</tr>
<tr>
<td>Ag → Ag⁺ + e⁻</td>
<td></td>
</tr>
<tr>
<td>Hg → Hg²⁺ + 2e⁻</td>
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</tr>
<tr>
<td>Pt → Pt²⁺ + 2e⁻</td>
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</tr>
<tr>
<td>Au → Au³⁺ + 3e⁻</td>
<td></td>
</tr>
</tbody>
</table>

Fig 4.14, p. 117
4.4 Oxidation-Reduction Reactions

- Metal displacement reactions

\[ \text{CuSO}_4(\text{aq}) + \text{Zn(s)} \rightarrow \text{Cu(s)} + \text{ZnSO}_4(\text{aq}) \]

What is the net ionic equation and what are the spectator ion(s)?

Figure 4.13, p. 116
4.4 Oxidation-Reduction Reactions

- Metal displacement reactions
  
  \[ \text{CuSO}_4(\text{aq}) + \text{Zn}(s) \rightarrow \text{Cu}(s) + \text{ZnSO}_4(\text{aq}) \]

  Will Cu(s) reduce Zn\(^{2+}\)(aq)?
<table>
<thead>
<tr>
<th>Element</th>
<th>Reaction</th>
<th>Activity Series</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li → Li⁺ + e⁻</td>
<td>React with cold water to produce H₂</td>
<td>Activity Series</td>
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<td>K → K⁺ + e⁻</td>
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</table>

Increasing activity – more reactive
As the strength of the reducing agent increases, it is more likely to participate in a redox reaction by undergoing oxidation (when coupled with a metal which will reduce).

Mg(s) will react with Cd\(^{2+}\)

Al(s) will react with Pb\(^{2+}\)

But:

Cu(s) will not react with Ni\(^{2+}\)

Ni(s) will not react with Na\(^{+}\)

You will be given this on an exam.
4.4 Oxidation-Reduction Reactions

- Halogen displacement reactions
  
  \[ 2\text{NaI}(aq) + \text{Br}_2(l) \rightarrow 2\text{NaBr}(aq) + \text{I}_2(s) \]
4.4 Oxidation-Reduction Reactions

- Halogen displacement reactions

Will any halogen oxidize iodide?

$\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$
4.4 Oxidation-Reduction Reactions

- Practice

  - Identify whether the reaction is a redox reaction. If it is, determine which species is oxidized and which is reduced.

1. \(2\text{Fe}^3+(aq) + 3\text{S}^2-(aq) \rightarrow \text{Fe}_2\text{S}_3(s)\)
2. \(2\text{FeCl}_3(aq) + 3\text{Co}(s) \rightarrow 2\text{Fe}(s) + 3\text{CoCl}_2(aq)\)
3. \(2\text{NH}_4\text{Cl}(aq) + \text{F}_2(g) \rightarrow 2\text{NH}_4\text{F}(aq) + \text{Cl}_2(aq)\)
4. \(\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)\)
5. \(2\text{HNO}_3(aq) + \text{Zn}(s) \rightarrow \text{Zn(NO}_3)_2(aq) + \text{H}_2(g)\)
6. \(2\text{H}_2\text{O}(l) + 2\text{K}(s) \rightarrow 2\text{KOH}(aq) + \text{H}_2(g)\)
Aluminum ions will be reduced by sodium metal (reaction 1) but not by chromium metal (reaction 2). Which substance is the strongest oxidizing agent? $\text{Al}^{3+}, \text{Na}^+, \text{Cr}^{3+}, \text{Na(s)}$

*Water molecules and spectator ions are not shown for clarity.*
Chapter 4 Practice

What is the oxidation number of phosphorus in sodium phosphate, \( \text{Na}_3\text{PO}_4 \)?

A. +8  
B. +5  
C. +3  
D. –3

Will silver ions react with copper metal? Why or why not?

A. Yes, because copper will reduce silver ions.
B. Yes, because copper will oxidize silver ions.
C. No, because copper will not reduce silver ions.
D. No, because copper will not oxidize silver ions.
4.5 Concentration of Solutions

- Key Definitions:
  - **Concentration:**
    - The amount of solute present in a given amount of solvent or a given amount of solution
  - **Molarity:**
    - Number of moles of solute per liter of solution
4.5 Concentration of Solutions

- Preparing a Solution

Figure 4.15, p. 120

Marker showing known volume of solution

Meniscus
4.5 Concentration of Solutions

Key Definitions:

- **Concentration:**
  - The amount of solute present in a given amount of solvent or a given amount of solution

- **Molarity:**
  - Number of moles of solute per liter of solution

- **Dilution:**
  - The procedure for preparing a less concentrated solution from a more concentrated solution
What is the final concentration of the chloride ion when 45 mL of a 0.250 M solution of aluminum chloride is diluted with 55 mL of water?
4.6 Solution Stoichiometry

- Key Definitions:
  - Gravimetric analysis:
    - An analytical technique based on the measurement of mass

Figure 4.17, p. 124
What is the molar concentration of 42.0 mL chromium (III) sulfate needed to react completely with 95.0 mL of a 1.087 M solution of sodium carbonate.

What mass of solid product will be produced?
4.6 Solution Stoichiometry

- **Key Definitions:**
  - **Titration:**
    - A technique in which a standard solution is added slowly to another solution of unknown concentration
  - **Equivalence point:**
    - The point at which the acid has completely reacted with or been neutralized by the base
  - **Indicator:**
    - Substances that have distinctly different colors in acidic or basic media

Figure 4.18, p. 125
4.6 Solution Stoichiometry

A standardized solution of sodium hydroxide has a concentration of 0.1055 M.

A 25.00 mL solution of carbonic acid reacts with 39.5 mL of the sodium hydroxide solution.

What is the molar concentration of the carbonic acid?
When the two solutions shown (which are aluminum chloride and barium chloride) are combined to a final volume of 105 mL (by adding water), what is the concentration of all ions?

Each symbol represents 0.01 mol of an ion. Water molecules are not shown for clarity.

Which ion is represented by which symbol?
Chapter 4 Practice

What is the molar concentration of the chloride ion in 125 mL of a 0.0500 M FeCl₃?

A. 0.00625 M  B. 0.0167 M  C. 0.0500 M  D. 0.150 M

What volume (in mL) of water is added to 25.00 mL of a 0.750 M sodium chloride solution to yield a final concentration of 0.350 M? Assume the volume are additive.

A. 11.7 mL  B. 13.3 mL  C. 28.6 mL  D. 53.6 mL