Chapter 4 Solutions Stoichiometry

Chapter 4: Types of Chemical reactions and Solution Stoichiometry

If you're not part of the solution, you're part of the precipitate

4.1 Water, The Common Solvent

- State why water acts as a common solvent.
- Draw the structure of water, including partial charge.
- Write equations for the dissociation of some ionic salts in water.

a. Water is not a linear molecule. It is bent at an angle of about 105°.
b. Electrons are not evenly distributed around the atoms in water. The molecule is polar because the charges are not distributed symmetrically.
c. Like dissolve like. The following classes of molecules, in general, are miscible:
   - polar and ionic
   - polar and polar
   - nonpolar and nonpolar

Ionic salts dissolve in water. Compounds that contain only carbon and hydrogen are nonpolar.

Polar Water Molecules Interact with the Positive and Negative Ions of a Salt Assisting in the Dissolving Process
Example 4.1A Will the Substances Mix?

Predict whether each pair of substances will mix. State why or why not.

a. NaNO₃ and H₂O
b. C₆H₁₄ and H₂O
c. I₂ and C₆H₁₄
d. I₂ and H₂O

The dissociation of simple salts in water is often written as shown in the following equations:

\[
\text{NaI(s)} \xrightarrow{\text{H}_2\text{O(l)}} \text{Na}^+(\text{aq}) + \text{I}^-\text{(aq)}
\]

\[
\text{K}_2\text{Cr}_2\text{O}_7\text{(s)} \xrightarrow{\text{H}_2\text{O(l)}} 2\text{K}^+(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}\text{(aq)}
\]

\[
\text{Ba(OH)}_2\text{(s)} \xrightarrow{\text{H}_2\text{O(l)}} \text{Ba}^{2+}\text{(aq)} + 2\text{OH}^-\text{(aq)}
\]

Example 4.1B Practice with equations

Complete each of the following dissociation equations:

a. \( \text{CaCl}_2\text{(s)} \xrightarrow{\text{H}_2\text{O(l)}} \)

b. \( \text{Fe(NO}_3\text{)}_2\text{(s)} \xrightarrow{\text{H}_2\text{O(l)}} \)

c. \( \text{KBr(s)} \xrightarrow{\text{H}_2\text{O(l)}} \)

d. \( (\text{NH}_4)\text{Cr}_2\text{O}_7 \xrightarrow{\text{H}_2\text{O(l)}} \)
### 4.2 The Nature of Aqueous Solutions: Strong and Weak Electrolytes

- Classify many substances as strong, weak or nonelectrolytes.

- **Solute**: 1. If it and the solvent are present in the *same phase*, it is the one in *lesser amount*.  
  2. If it and the solvent are present in *different phases*, it is the one that *changes phase*.  
  3. It is the one that *dissolves into the solvent*.

- **Solvent**: 1. If it and the solute are present in the *same phase*, it is the one in *greater amount*.  
  2. If it and the solvent are present in *different phases*, it is the one that *retains its phase*.  
  3. It is the one *into which the solute dissolves*.

An *aqueous solution* means that *water is the solvent*.
**Strong Acids**

HCl (\(\text{H}_2\text{O}(l)\)) \(\rightarrow\) \(\text{H}^+(aq) + \text{Cl}^-(aq)\)

HNO\(_3\) (\(\text{H}_2\text{O}(l)\)) \(\rightarrow\) \(\text{H}^+(aq) + \text{NO}_3^-(aq)\)

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

**Strong Bases**

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

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

**Strong acids, they are dissociated 100% in solution.**

- HCl hydrochloric acid
- HNO\(_3\) nitric acid
- H\(_2\text{SO}_4\) sulfuric acid
- HBr hydrobromic acid
- HI hydroiodic acid
- HClO\(_4\) perchloric acid

**Strong bases.**

- LiOH lithium hydroxide
- NaOH sodium hydroxide
- KOH potassium hydroxide
- RbOH rubidium hydroxide
- CsOH cesium hydroxide

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**Weak Base**

The Reaction of NH\(_3\) in Water

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

**Weak Acid**

Acetic Acid (HC\(_2\text{H}_3\text{O}_2\))

HC\(_2\text{H}_3\text{O}_2\)(aq) (\(\text{H}_2\text{O}(l)\)) \(\rightarrow\) \(\text{H}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq)\)
### Chapter 4 Solutions Stoichiometry

<table>
<thead>
<tr>
<th>Electrolyte</th>
<th>Conductivity</th>
<th>Degree of Dissociation</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Strong</td>
<td>high</td>
<td>total</td>
<td>Strong acids such as HCl; many salts such as NaCl and Sr(NO₃)₃; strong base such as NaOH, Ba(OH)₂ and Group I and II hydroxides</td>
</tr>
<tr>
<td>Weak</td>
<td>low to moderate</td>
<td>partial</td>
<td>Weak organic acids such as HCO₂H and HC₂H₃O₂; weak bases such as NH₃</td>
</tr>
<tr>
<td>non</td>
<td>non</td>
<td>close to zero</td>
<td>Sugar, AgCl, Fe₂O₃</td>
</tr>
</tbody>
</table>

#### Example 4.2 Strong, Weak, or Nonelectrolyte

List whether each of the following is a strong, weak or nonelectrolyte.

- a. HClO₄
- b. C₆H₁₂
- c. LiOH
- d. NH₃
- e. CaCl₂
- f. HC₂H₃O₂

#### 4.3 The Composition of Solutions

- Determine the molarity of a solution
- Calculate the molarity of each ion in a solution.
- Determine the mass and/or volume of reagents necessary to prepare a solution of a given molarity.
- Solve problems related to dilution.

**Molarity (M) is defined as moles of solute per liter of solution.**

$$M = \frac{\text{moles of solute}}{\text{liter of solution}}$$

Keep in mind that

$$\text{moles} \frac{\text{liter}}{\text{liter}} = \text{millimoles} \frac{\text{milliliter}}{\text{milliliter}} = \text{micromoles} \frac{\text{microliter}}{\text{microliter}}$$

but \( \text{moles} \frac{\text{liter}}{\text{liter}} \) DOES NOT EQUAL \( \text{millimoles} \frac{\text{liter}}{\text{milliliter}} \) or \( \text{moles} \frac{\text{liter}}{\text{microliter}} \)

Be very careful with your units!
Steps Involved in the Preparation of a Standard Aqueous Solution

Describe how you would prepare 100. mL of 1M ammonium sulfide solution.

Describe how you would prepare 50. mL of 2.5M potassium carbonate solution.

Example 4.3A Calculating Molarity

Calculate the molarity of a solution prepared by dissolving 11.85 g of solid KMnO₄ in enough water to make 750. mL of solution.

Example 4.3B Mass From Molarity

Calculate the mass of NaCl needed to prepare 175. mL of a 0.500 M of NaCl solution.

Example 4.3C Volume From Molarity

How many mL of solution are necessary if we are to have 2.48 M NaOH solution that contains 31.52 g of the dissolved solid?
When we calculate the molarity of solute we MUST take into account that **strong electrolytes completely dissociate.** It is generally acceptable to discuss the solution concentration as "molarity of NaCl" it is more correct to chemically to discuss "molarity of Na\(^+\) ions" and the "molarity of Cl\(^-\) ions"

A solution that is 0.85 M NaCl is really 0.85 M in Na\(^+\) ion and 0.85 M in Cl\(^-\) ion because NaCl completely dissociates, and the dissociation rate is 1 to 1 to 1 ratio; that is *one mole of NaCl dissociates into one mole of Na\(^+\) ion and one mole of Cl\(^-\) ion.*

**Example 4.3D  Molarity of Ions in Solution**

Calculate the molarity of all the ions in each of the following solutions.

- a. 0.25 M Ca(ClO\(_4\))\(_2\)
- b. 2 M CrCl\(_3\)

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**Example 4.3E  Molarity of Ions in Solution**

Determine the molarity of Cl\(^-\) ion in a solution prepared by dissolving 9.82 g of CuCl\(_2\) in enough water to make 600 mL of solution.

**Strategy**

1. Calculate the solute concentration (molarity of the solute)
2. Determine the ion-to-solute mole ratio by writing the dissociation equation.
3. Use the mole ratio, along with solute concentration, to calculate the ion concentration.

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**Example 4.3F  Practice with Ions Concentration**

Determine the molarity of Fe\(^{3+}\) ions and SO\(_4^{2-}\) ions in a solution prepared by dissolving 48.05 g of Fe\(_2\)(SO\(_4\))\(_3\) in enough water to make 800 mL of solution.

- If the compound is a hydrate the molar mass used MUST include the water molecules
An important part of your chemistry experience is to be able to prepare dilute solutions from more concentrated ("stock") solutions. The most important idea in diluting solutions is that

\[ \text{moles of solute after dilution} = \text{moles of solute before dilution} \]

If the moles of solute remains identical before and after dilution (only the amount of water changes), then

\[ M_iV_i = M_fV_f \]

**Example 4.3G Preparation of a Dilute Solution**

What volume of 12 M hydrochloric acid must be used to prepare 600 mL of a 0.30 M HCl solution?

**Example 4.3H More Practice Preparing Dilute Solutions**

What volume of 1.0 M sodium hydroxide must be used to prepare 1.2 L of a 0.9 M NaOH solution?

### 4.7 Stoichiometry of Precipitation Reactions

- Solve a variety of problems involving the formation of precipitates

Solving problems involving precipitates from solution makes use of molarity, solubility rules, balancing equations, and limiting reactant calculations.

**SIX STEPS to solving solution problems**

1. Identify the species present in the combined solution, and determine what reaction occurs.
2. Write the balanced net ionic equation for the reaction.
3. Calculate the moles of reactants.
4. Determine which reactant is limiting.
5. Calculate the moles of product or products, as required.
6. Convert to grams or other units, as required.
**Example 4.7A An Introduction to Problems Based on Precipitation Reactions**

Calculate the mass of Ag₂S produced when 125. mL of 0.200 M AgNO₃ is added to excess Na₂S solution.

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**Example 4.7B Practice With Precipitation Problems**

What mass of Fe(OH)₃ is produced when 35. mL of 0.250 M Fe(NO₃)₃ solution is mixed with 55. mL of 0.180 M KOH solution?

Calculate the concentration of each ion remaining in solution after precipitation is complete.

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### 4.8 Acid-Base Reactions

- Solve a variety of problems related to acid-base neutralization.

Acid-base reaction problem solving require the same strategy as the other types of problems in this chapter. **Writing a balanced chemical equation is always your first, and most important step.**

- *AN ACID IS A PROTON DONOR*
- *A BASE IS A PROTON ACCEPTOR*

**Steps for Acid-Base Calculations**

1. List the species present in solution before reaction.
2. Write the balanced net ionic equation.
3. Calculate the moles of reactants.
4. Determine the limiting reactant where appropriate.
5. Calculate the moles of required reactant or product.
6. Convert to grams or volume (of solution), as required.

**Neutralization Reaction** is an acid-base reaction when "just enough" base is added to react with the acid in a solution, we say the acid has been neutralized.
Example 4.8A Neutralization of a Strong Acid

How many mL of a 0.800 M NaOH solution is needed to just neutralize 40.00 mL of a 0.600 M HCl solution?

Potassium hydroxide base solution (900.0 mL, 1.00 M) is added to 600.0 mL of 1.5 M HCl solution. What is the concentration of the excess H⁺ or OH⁻ ions are left in this solution after the neutralization is complete?

4.9 Oxidation-Reduction Reactions

- Assign oxidation states to atoms in a compound.
- Define oxidation and reduction

Oxidation-Reduction or Redox Reactions - are reactions in which one or more electrons are transferred.

<table>
<thead>
<tr>
<th>The Oxidation State of . . .</th>
<th>Summary</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>An atom in an element</td>
<td>Element: 0</td>
<td>Na(s), O₂(g), O₃(g), Hg(l)</td>
</tr>
<tr>
<td>A monatomic ion is the same as its charge</td>
<td>Monatomic ion: charge of ion</td>
<td>Na⁺, Cl⁻</td>
</tr>
<tr>
<td>Fluorine is −1 in its compounds</td>
<td>Fluorine: −1</td>
<td>HF, PF₃</td>
</tr>
<tr>
<td>Oxygen is usually −2 in its compounds</td>
<td>Oxygen: −2</td>
<td>H₂O, CO₂</td>
</tr>
<tr>
<td>Exception: Peroxides (containing O₂²⁻), in which oxygen is −1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen is +1 in its covalent compounds</td>
<td>Hydrogen: +1</td>
<td>H₂O, HCl, NH₃</td>
</tr>
</tbody>
</table>
Example 4.9A Assigning Oxidation States

Assign oxidation states to each of the atoms in the following compounds

a. CaF₂  
b. C₂H₆  
c. H₂O  
d. ICl₅  
e. KMnO₄  
f. SO₄²⁻

Oxidation - is an increase in oxidation state (lose of electrons)

Reduction - is a decrease in oxidation state (a gain of electrons)

• Losing Electrons is Oxidation
• Gaining Electrons is Reduction

\[
\text{CO}_2(g) + \text{H}_2(g) \rightleftharpoons \text{CO}(g) + \text{H}_2\text{O}(g) \\
+4 \quad 0 \quad +2 \quad +1
\]

Carbon is called the oxidizing agent (electron acceptor)

Hydrogen is called the reducing agent (electron donor)
Chapter 4 Solutions Stoichiometry

**Example 4.9B Practice With Oxidation States**
Assign the oxidation states to each atom in the equation

\[ \text{Fe}_2\text{O}_3 + 2\text{Al} \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe} \]

**Example 4.9C Which Atoms Undergo Redox?**
For each reaction, identify the atoms that undergo reduction or oxidation. Also list the oxidizing and reducing agents.

a. \(2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})\)

b. \(\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})\)

c. \(2\text{AgCl}(\text{s}) + \text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{Ag}(\text{s}) + 2\text{Cl}^-(\text{aq})\)

d. \(2\text{MnO}_3(\text{aq}) + 16\text{H}^+(\text{aq}) + 5\text{C}_2\text{O}_4^{2-}(\text{aq}) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 10\text{CO}_2(\text{g}) + 8\text{H}_2\text{O}(\text{l})\)
4.10 Balancing Oxidation-Reduction Equations

- Balance redox equations using the oxidation states method.
- Write the unbalanced equation.
- Determine the oxidation states of all atoms in the reactants and products.
- Show electrons gained and lost using “tie lines.”
- Use coefficients to equalize the electrons gained and lost.
- Balance the rest of the equation by inspection.
- Add appropriate states.

Example:

- Balance the reaction between solid zinc and aqueous hydrochloric acid to produce aqueous zinc(II) chloride and hydrogen gas.

1. What is the unbalanced equation?

   • \( \text{Zn(s) + HCl(aq)} \rightarrow \text{Zn}^{2+}(aq) + \text{Cl}^-(aq) + \text{H}_2(g) \)

2. What are the oxidation states for each atom?

   • \( \begin{align*}
   \text{Zn(s) + HCl(aq)} & \quad \text{Zn}^{2+}(aq) + \text{Cl}^-(aq) + \text{H}_2(g) \\
   0 & +1 & -1 & +2 & -1 & 0
   \end{align*} \)

3. How are electrons gained and lost?

   - 1 e\(^-\) gained (each atom)

   \( \begin{align*}
   \text{Zn(s) + HCl(aq)} & \rightarrow \text{Zn}^{2+}(aq) + \text{Cl}^-(aq) + \text{H}_2(g) \\
   0 & +1 & -1 & +2 & -1 & 0
   \end{align*} \)

   - 2 e\(^-\) lost

   - The oxidation state of chlorine remains unchanged.

4. What coefficients are needed to equalize the electrons gained and lost?

   - \( 1 \text{ e}^- \text{ gained (each atom)} \times 2 \)

   \( \begin{align*}
   \text{Zn(s) + HCl(aq)} & \rightarrow \text{Zn}^{2+}(aq) + \text{Cl}^-(aq) + \text{H}_2(g) \\
   0 & +1 & -1 & +2 & -1 & 0
   \end{align*} \)

   - 2 e\(^-\) lost

   \( \text{Zn(s) + 2HCl(aq)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{Cl}^-(aq) + \text{H}_2(g) \)

5. What coefficients are needed to balance the remaining elements?

   • \( \text{Zn(s) + 2HCl(aq)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{Cl}^-(aq) + \text{H}_2(g) \)