Solutions

Chapter 7: Solutions
Educational Goals

1. Compare and contrast **mixtures** and **pure substances**.

2. Understand, compare, and contrast the terms **homogeneous mixture** and **heterogeneous mixture**. For a homogeneous mixture, explain the difference between **solute(s)** and **solvent**.

3. Predict the effect of temperature and pressure on the **solubility** of gases in water and the effect of temperature on the solubility of solids in water.

4. Be able to use the **Solubility Rules Table** to determine if an ionic compound will significantly dissolve in water.

5. Predict whether a **precipitation reaction** will occur when two specified aqueous solutions of ionic compounds are mixed; if a precipitation reaction will occur, write a balanced chemical equation for the reaction.

6. Compare the relative solubilities of organic molecules based on the functional groups or the relative sizes of the hydrocarbon (nonpolar) regions.
Educational Goals

7. Explain, compare, and contrast the terms hydrophilic, hydrophobic, and amphipathic, and give examples of compounds that belong to each category.

8. Be able to calculate the concentration of a solution using various concentration units of measurements. (%, parts per thousand, ppm, ppb, molarity, molality, osmolality, osmolarity, and Eq/L)

9. Given the concentration, be able to convert from the volume of solution to the amount of solute (and vice versa).

10. Given a solution’s initial concentration, be able to use the dilution equation to determine the concentration of the solution after dilution.

11. Compare and contrast solutions, suspensions, and colloids.

12. Describe the processes of diffusion and osmosis. Define osmotic pressure and predict the effect of solute concentration on the osmotic pressure.
Matter can be classified into one of two groups:

1) Mixtures
2) Pure Substance
Pure Substances

A pure substance can be either an element (just one type of atom) or a compound.

**elements**
- mercury
- silver
- oxygen

**compounds**
- water
- NaCl
- sucrose
Mixtures

• Most objects around us are *mixtures*. 
  – Example: most foods and beverages, air, the chair you are sitting in.

• The components are physically mixed together and can be separated by physical methods.

• The properties of mixtures depend on the *identity* and *amount* of each component in the mixture 
  – Therefore to fully describe a mixture, the composition of the mixture must be stated.
    • Example: 5% Starch solution, 75:25% Zinc/Copper Alloy
Mixtures

- Mixtures can be either uniform in composition (homogeneous) or the composition can vary throughout the sample (heterogeneous).
• **Homogeneous mixtures** = uniform throughout, appears to be one thing
  – Homogenous mixtures are called **solutions**.
  – Solutions can be liquid, gas, or solid
  – Examples: Sugar dissolved in water (liquid)
    Air (gas)
    14-karat gold (solid mixture of gold, silver, and copper)

• **Heterogeneous mixtures** = non-uniform, contains regions with different properties than other regions
  – Examples: chocolate chip cookies
    sand in water
Types of Mixtures

- *solution*—a homogenous mixture

- *suspension*—a solution that appears uniform while being stirred but separates quickly after stirring

- *colloid*—a mixture in which small particles are suspended throughout

We will discuss *solutions* first, later we will come back to briefly describe suspensions and colloids.
Solutions

• The primary ingredient in a solution is called the **solvent**.
• The other ingredients are the **solute**s and are said to be **dissolved** in the solvent.

• Water is the most common solvent.
• Water is a unique solvent because so many substances can dissolve in it.
• Solutions in which water is the solvent are called **aqueous** solutions.
Solutions

- Example: Solution of ethanol in water
Electrolyte Solutions

- **Electrolytes** are compounds that dissolve in water to form ions.

  - The consequences of having too much or too little of a particular electrolyte can have profound effects on your health.

  - Ions important to your health:
    - \( \text{Na}^+ \), \( \text{K}^+ \) = transmission of nerve impulses
    - \( \text{Ca}^{2+} \) = muscle contraction and blood clotting
Electrical conductivity of aqueous solutions

(a) Pure water *does not conduct* an electric current.
(b) When an ionic compound is dissolved in water, current flows and the lamp lights.
Dissolving: Dissolution of CuCl$_2$

CuCl$_2$: Strong Electrol yte
Dissolving: Dissolution of NaCl

NaCl:
Strong Electrolyte
Aqueous ethanol solution

Ethanol molecules are **solvated**, but do not ionize!!!

- Ethanol is not an ionic compound
- Non electrolyte
Aqueous sugar solution

Sugar molecules are solvated, but do not ionize!!!

Non Electrolyte
Solutions: Solubility

The formation of liquid solutions, such as Kool Aid or dissolved NaCl, requires that solute particles be able to interact with the solvent molecules through noncovalent interactions.

“like dissolves like”
Solubility

• **Solubility** refers to the amount of *solute* that will dissolve in a given amount of solvent at a given temperature.

• **Temperature** can affect solubility.
  – For all gaseous solutes (O$_2$, CO$_2$, etc.), an increase in temperature leads to a decrease in solubility.
  – Most liquid and solid solutes become more soluble in water as temperature increases.
Solubility

Some ionic compounds are *highly* soluble in water solutions.

Other ionic compounds are *insoluble* or only *slightly* soluble.
With *ionic compounds*, it is not so easy to predict if they will dissolve in water!

- We need to look it up in a solubility table

<table>
<thead>
<tr>
<th>Water Soluble</th>
<th>Compound</th>
<th>Example</th>
<th>Exceptions</th>
<th>Exception Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrates</td>
<td>NaNO₃</td>
<td>None</td>
<td>None</td>
<td>None</td>
</tr>
<tr>
<td>Chlorides, Bromides, and Iodides</td>
<td>NaCl</td>
<td>Compounds containing Ag⁺, Pb²⁺, or Hg⁺, and HgI₂</td>
<td>AgCl</td>
<td></td>
</tr>
<tr>
<td>Sulfates</td>
<td>K₂SO₄</td>
<td>Compounds containing Pb²⁺, Sr²⁺, Ba²⁺, or Hg⁺</td>
<td>PbSO₄</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Water Insoluble</th>
<th>Compound</th>
<th>Example</th>
<th>Exceptions</th>
<th>Exception Example(s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydroxides</td>
<td>Mg(OH)₂</td>
<td>Compounds containing alkali (Group I) metals or Ca²⁺, Sr²⁺, Ba²⁺, NH₄⁺</td>
<td>NaOH</td>
<td></td>
</tr>
<tr>
<td>Phosphates, Carbonates, and Chromates</td>
<td>FePO₄</td>
<td>Compounds containing alkali (Group I) metals or NH₄⁺</td>
<td>K₂CO₃, Li₃PO₄, Na₂CrO₄</td>
<td></td>
</tr>
</tbody>
</table>
For homework problem 7.7:

- **CuS** Copper (II) sulfide *is insoluble*
Group Work

Indicate whether each ionic compound is soluble or insoluble in water.

a. $\text{KNO}_3$

b. $\text{BaSO}_4$

c. $\text{LiI}$

**Strategy:** Refer to Table
Saturated Solution

At some point, even with highly soluble compounds, a solution will reach its capacity to dissolve.

If more of solute is added to a saturated solution it will not dissolve.

• At that point, solute particles are re-crystallizing at the same rate that they are dissolving.

• equilibrium
Example:

a. If 45.0 g of NaCl are added to 100g of water at 20 °C, will all of the NaCl dissolve?

b. Is the resulting mixture homogeneous or heterogeneous?
The Solubility of Gases in Water
The solubility of gases in solution depend on:

1) **Temperature**
2) **Pressure**
Henry’s Law

The solubility of a gas in a liquid is proportional to the pressure of the gas over the liquid.

The higher the pressure of a gas above a liquid, the greater its solubility.
Any Fishermen in the Class?

On a hot day, it is often easier to catch fish by casting your line into a deep, cool part of a lake than into a shallow, warm spot. One reason that fish gather in cool water may be related to the levels of oxygen dissolved in the water. Explain.
In a double replacement reaction, two substances “switch partners.”

The general form of a double replacement reaction, where compounds $AX$ and $BY$ switch partners, is:

$$AX + BY \rightarrow AY + BX$$
Reactions of Ions in Aqueous Solutions

Double Replacement Reactions:
  1) Precipitation reactions
  2) Gas Producing Reactions
1) Precipitation Reactions

*Precipitation reactions* may occur when *two* solutions that contain *dissolved ions* are mixed.

In a precipitation reaction, two compounds in aqueous solution appear to exchange *ions*.

\[ \text{A}^+ \text{X}^- + \text{B}^+ \text{Y}^- \rightarrow \text{A}^+ \text{Y}^- + \text{B}^+ \text{X}^- \]

If one of the new pairs formed is *insoluble*, a new substance (solid/precipitate) is formed.
For a precipitation reaction to occur, at least one of the products formed is **insoluble** in water.

Therefore, a **solid** is *always* formed in a precipitation reaction.

Often, many *tiny* crystals are formed and this gives the mixture a cloudy appearance. The cloudy appearance may be white, black, or some other color, depending on the identity of the particular solid that is formed.

We say the solid “*precipitated*” from the solution.

The appearance of the solid precipitate indicates the formation of **new ionic bonds** and that a **reaction** has occurred.
The *educational goals* for *precipitation reactions* are:

- Predict if a precipitation reaction will occur when two aqueous ionic compounds are combined.
- Write the balanced chemical equation for the reaction.
Method for Predicting if a Precipitation Reaction will Occur and Writing the Balanced Chemical Equation for Precipitation Reactions

Example: The reaction of lead(II) nitrate and potassium chromate.

Step 1: Write reactants’ names and arrow for the chemical equation using word form (not the chemical formula).

\[ \text{lead(II) nitrate} + \text{potassium chromate} \rightarrow \]

Step 2: Add the “possible” products to the word equation by switching anions.

\[ \text{lead(II) nitrate} + \text{potassium chromate} \rightarrow \text{lead(II) chromate} + \text{potassium nitrate} \]
Method for Predicting if a Precipitation Reaction will Occur and Writing the Balanced Chemical Equation for Precipitation Reactions

Step 3: Convert the word equation to a formula equation.

\[ \text{lead(II) nitrate} + \text{potassium chromate} \rightarrow \text{lead(II) chromate} + \text{potassium nitrate} \]

\[ \text{Pb(NO}_3\text{)}_2 + \text{K}_2\text{CrO}_4 \rightarrow \text{PbCrO}_4 + \text{KNO}_3 \]

Note: Students often need to review the section in chapter 3 that discusses naming ionic compounds in order to perform Step 3.

Step 4: Balance the equation:

\[ \text{Pb(NO}_3\text{)}_2 + \text{K}_2\text{CrO}_4 \rightarrow \text{PbCrO}_4 + 2 \text{KNO}_3 \]

Step 5: Add the phase of each of the reactants and “possible” products to the chemical equation.

\[ \text{Pb(NO}_3\text{)}_2(aq) + \text{K}_2\text{CrO}_4(aq) \rightarrow \text{PbCrO}_4(s) + 2 \text{KNO}_3(aq) \]
Method for Predicting if a Precipitation Reaction will Occur and Writing the Balanced Chemical Equation for Precipitation Reactions

Step 5: Add the **phase** of each of the reactants and “possible” products to the chemical equation.

- In all precipitation reactions, **the reactants are always aqueous**.
- Use the Solubility Rules Table to determine the phase of the “possible” products.
- If a compound is water **soluble**, it remains dissolved and we write “(aq).”
- If a compound is water **insoluble**, it precipitates as a solid and we write “(s).”

\[
Pb(NO_3)_2 (aq) + K_2CrO_4 (aq) \rightarrow PbCrO_4 (s) + 2 KNO_3 (aq)
\]
Method for Predicting if a Precipitation Reaction will Occur and Writing the Balanced Chemical Equation for Precipitation Reactions

Example: The reaction of sodium chloride and silver nitrate.

**Step 1:** Write *reactants’ names* and arrow for the chemical equation using *word form* (not the chemical formula).

\[
\text{sodium chloride} + \text{silver nitrate} \rightarrow 
\]

**Step 2:** Add the “possible” *products* to the word equation by switching *anions*.

\[
\text{sodium chloride} + \text{silver nitrate} \rightarrow \text{sodium nitrate} + \text{silver chloride} 
\]
Method for Predicting if a Precipitation Reaction will Occur and Writing the Balanced Chemical Equation for Precipitation Reactions

Step 3: Convert the word equation to a formula equation.

\[
\text{sodium chloride} + \text{silver nitrate} \rightarrow \text{sodium nitrate} + \text{silver chloride} \\
\text{NaCl} + \text{AgNO}_3 \rightarrow \text{NaNO}_3 + \text{AgCl}
\]

Step 4: Balance the equation:

- In this example, the equation is already balanced; each of the coefficients is “1.”

\[
\text{NaCl} + \text{AgNO}_3 \rightarrow \text{NaNO}_3 + \text{AgCl}
\]

Step 5: Add the phase of each of the reactants and “possible” products to the chemical equation.

\[
\text{NaCl} (aq) + \text{AgNO}_3 (aq) \rightarrow \text{NaNO}_3 (aq) + \text{AgCl} (s)
\]
Example: Determine if a precipitation reaction would occur when a *sodium chloride solution* is mixed with a *potassium nitrate solution*.

**Step 1:** Write *reactants’ names* and arrow for the chemical equation using *word form* (not the chemical formula).

\[
\text{sodium chloride} + \text{potassium nitrate} \rightarrow \]

**Step 2:** Add the “possible” *products* to the word equation by switching *anions*.

\[
\text{sodium chloride} + \text{potassium nitrate} \rightarrow \text{sodium nitrate} + \text{potassium chloride} \]
Step 3: Convert the **word** equation to a **formula** equation.

sodium chloride  +  potassium nitrate  →  sodium nitrate  +  potassium chloride

\[
\text{NaCl} + \text{KNO}_3 \rightarrow \text{NaNO}_3 + \text{KCl}
\]

Step 4: **Balance** the equation:

In this example, the equation is already balanced; each of the coefficients is “1.”

\[
\text{NaCl} + \text{KNO}_3 \rightarrow \text{NaNO}_3 + \text{KCl}
\]
Step 5: Add the \textit{phase} of each of the reactants and “possible” products to the chemical equation.

\[
\text{NaCl} \ (aq) \ + \ \text{KNO}_3 \ (aq) \ \rightarrow \ \text{NaNO}_3 \ (aq) \ + \ \text{KCl} \ (aq)
\]

**NO REACTION**

**IMPORTANT:** If both of the “possible” products are \textit{water soluble}, then no reaction occurred.

There were solvated cations and anions in each of the two solutions before mixing, then the solutions were mixed and the cations and anions remained solvated in the mixture.

No new chemical bonds were made, therefore no chemical reaction occurred.

When no reaction occurs in precipitation reaction problems such as this example, you can write “No Reaction” instead of the “possible” products.
You try one:

Determine if a precipitation reaction would occur when a silver nitrate solution is mixed with a barium chloride solution and, if a reaction does occur, write the balanced chemical equation.
2) Gas Producing Double Replacement Reactions

A gas producing double replacement reaction is a special type of double replacement in which a gas is produced.

Example: The reaction of aqueous hydrogen monochloride (HCl, also know as hydrochloric acid) and aqueous sodium bicarbonate (NaHCO$_3$).

\[ \text{HCl} + \text{NaHCO}_3 \rightarrow \text{HHCO}_3 + \text{NaCl} \]

In this reaction, the bicarbonate and chloride anions switch partners to form aqueous carbonic acid (HHCO$_3$) and sodium chloride.
In the chemical equation on the previous slide, I wrote the formula of carbonic acid as $\text{HHCO}_3$ in order to help you see how $\text{Cl}^-$ and $\text{HCO}_3^-$ “switched partners”; however the correct way to write the formula for carbonic acid is $\text{H}_2\text{CO}_3$, as described below.

$$\text{HCl} \text{ (aq)} + \text{NaHCO}_3 \text{ (aq)} \rightarrow \text{HHCO}_3 \text{ (aq)} + \text{NaCl} \text{ (aq)}$$

Carbonic acid decomposes to $\text{H}_2\text{O} \text{ (l)}$ and $\text{CO}_2 \text{ (g)}$

$$\text{HCl} \text{ (aq)} + \text{NaHCO}_3 \text{ (aq)} \rightarrow \text{H}_2\text{CO}_3 \text{ (aq)} + \text{NaCl} \text{ (aq)}$$

$$\text{HCl} \text{ (aq)} + \text{NaHCO}_3 \text{ (aq)} \rightarrow \text{H}_2\text{O} \text{ (l)} + \text{CO}_2 \text{ (g)} + \text{NaCl} \text{ (aq)}$$

Gas Producing Double Replacement Reaction
Organic Compounds in Solutions

Note: a solution can be made in other solvents than water
When will solutions be formed?

“Like dissolves like”

Solutions form when these three kinds of forces are similar.
Solubility and London forces.

(a) In pure pentane \((\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3)\) and pure hexane \((\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3)\), the molecules are held to one another by London forces.

(b) When the two hydrocarbons are mixed, London forces are still at work.
Solubility and hydrogen bonding.

(a) In pure ethanol (CH$_3$CH$_2$OH) and pure water, the molecules are held to one another primarily by hydrogen bonds. (b) When the two compounds are mixed, hydrogen bonds form between alcohol and water molecules.
• Which pairs will form solutions?
  – CCl₄ and water
  – Gasoline and MgSO₄
  – Hexane (C₆H₁₄) and heptane (C₇H₁₄)
  – Ethyl alcohol (C₂H₅OH) and heptanol (C₇H₁₅OH)
The Solubility of Organic Compounds in Water

- Although hydrocarbons are not soluble in water, many organic compounds are.
  - Hydrogen bonding with water helps.
  - Polar molecules can be soluble in water.
Example: Alcohols

- The larger the hydrocarbon in an alcohol, the less soluble it is in water.

\[ \text{CH}_3\text{CH}_2\text{OH} \quad \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{OH} \]

Soluble \quad \text{Slightly Soluble}

Ethanol \quad \text{1-Octanol}
Example: Carboxylic acids

• As with alcohols, the larger the hydrocarbon that a carboxylic acid has, the less soluble it is in water.

Soluble

Acetic acid

Octanoic acid

Slightly Soluble
Example: Esters

- A similar pattern is seen with esters

\[
\begin{align*}
\text{Ethyl methanoate} & : & \text{Butyl heptanoate} \\
H-C-OCH_2CH_3 & & CH_3CH_2CH_2CH_2CH_2CH_2C-OCH_2CH_2CH_2CH_2CH_3
\end{align*}
\]

Slightly Soluble \hspace{2cm} \text{Not Soluble}
Biochemical Compounds
The compounds found in living things can be placed into one of three solubility classes.

1. **hydrophilic** - those that are soluble in water

2. **hydrophobic** - those that are insoluble in water

3. **amphipathic** - have both hydrophilic and hydrophobic parts
Hydrophilic Compounds

Since \textit{like dissolves like}, to be hydrophilic, a molecule must resemble \textit{water}:

–polar groups \textit{and/or}

–those able to form hydrogen bonds
Examples of **Hydrophilic Compounds**

**Carbohydrates**

- Glucose
- Fructose
Examples of Hydrophilic Compounds

- Glycine
- Alanine
- Phenylalanine
- Valine

Amino Acids
Examples of Hydrophobic Compounds

Cholesterol is insoluble in water... therefore hydrophobic.
Examples of **Hydrophobic** Compounds

Triglycerides are insoluble in water....therefore hydrophobic.

```
CH₃(CH₂)₅CH=CH(CH₂)₇C−OCH₂

CH₃(CH₂)₄(CH=CHCH₂)₂(CH₂)₆C−OCH

CH₃(CH₂)₁₄C−OCH₂
```
Examples of Hydrophobic Compounds

Fatty acids (long hydrocarbon chain carboxylic acids) are insoluble in water.

- **Stearic acid**
  \[
  \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3\text{COOH}
  \]

- **Oleic acid**
  \[
  \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3\text{COOH}
  \]

- **Linoleic acid**
  \[
  \text{CH}_3\text{CH}_2\text{CH}=\text{CHCH}_2\text{CH}=\text{CHCH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3\text{COOH}
  \]
Example of Amphipathic Compounds

Amphipathic compounds have one part of the molecule that is hydrophilic and another part that is hydrophobic.

Very often the polar end of amphipathic compounds will have a formal charge.
Example of Amphipathic Compounds

\[ \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{C}^\text{O} \text{O}^\text{-} \]

Nonpolar “Tail”            Polar “Head”

Amphipathic compounds are often illustrated using the ball and stick model.

- Ball = Polar Head
- Stick = Nonpolar Tail
Amphipathic Compounds in Water

When added to water, amphipathic compounds can form a:

• (a) monolayer
• (b) micelle
Emulsification by Amphiphatic Compounds: Soaps

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Problem:
Based on its structure, do you expect vitamin C to be hydrophobic or hydrophilic? Explain.
Sodium lauryl sulfate is a detergent often present in shampoos. Identify this molecule as hydrophilic, hydrophobic, or amphipathic.

CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂O−SO−O−Na⁺
Concentration:

Describing the amount of solute in a solution
The term **concentration** is used to refer to the amount of solute that is dissolved in a solvent.

– A **saturated solution** holds the maximum amount of solute that can be dissolved at a particular temperature.

– An **unsaturated solution** holds less than a saturating amount of solute.
Concentration by Percent

In general, percentage can be defined as:

\[ \% = \left( \frac{\text{Part}}{\text{Whole}} \right) \times 100 \]

In chemistry, we are usually interested in percent by mass:

\[ \% \text{ Mass} = \left( \frac{\text{Mass of Part}}{\text{Total Mass}} \right) \times 100 \]
Concentration by Percent

• Percent means the same thing as “parts per hundred”, so when percent is used as a concentration unit, the number of parts of solute present in every 100 parts of solution is being specified.

• There are three commonly used percent measurements for concentration:

1. weight/weight  \%(w/w) 
2. volume/volume  \%(v/v) 
3. weight/volume  \%(w/v)
\[
\% \text{ (Weight/Weight)} = \left( \frac{\text{grams of Solute}}{\text{grams of Solution}} \right) \times 100\%
\]

\[
\% \text{ (Vol/Vol)} = \left( \frac{\text{mL of Solute}}{\text{mL of Solution}} \right) \times 100\%
\]

\[
\% \text{ (Weight/Vol)} = \left( \frac{\text{grams of Solute}}{\text{mL of Solution}} \right) \times 100\%
\]
Example

Potassium iodide (KI) is used to treat iodine deficiencies. What is the %(w/v) of a 75 mL solution containing 2.0g of KI?

\[
\frac{\text{grams of Solute}}{\text{mL of Solution}} \times 100\% 
\]

\[
\frac{2.0 \text{ g of KI}}{75 \text{ mL of Solution}} \times 100\% 
\]

\[= 2.7\% (w/v)\]
Parts per thousand, parts per million, and parts per billion

These units are mostly used for very dilute solutions.

If 30 mg of sucrose is dissolved in a swimming pool, the sucrose concentration is 0.001 ppm and 1 ppb.
Parts per thousand, parts per million, and parts per billion

Recall that % is parts per hundred:

\[
\% \ (w/v) = \left( \frac{\text{grams of Solute}}{\text{mL of Solution}} \right) \times 100\% 
\]

We can easily calculate parts per thousand:

\[
\text{parts per thousand} = \left( \frac{\text{grams of Solute}}{\text{mL of Solution}} \right) \times 1000 
\]
Parts per thousand, parts per million, and parts per billion

**parts per million:**

$$\text{ppm (w/v)} = \left( \frac{\text{grams of Solute}}{\text{mL of Solution}} \right) \times 10^6$$

**parts per billion:**

$$\text{ppb (w/v)} = \left( \frac{\text{grams of Solute}}{\text{mL of Solution}} \right) \times 10^9$$
Many cities add sodium fluoride to their drinking water to help reduce dental cavities. If 25 L of city water contains 0.018g of sodium fluoride, what is the concentration in parts per billion?

Example:

\[
\text{ppb (w/v)} = \left( \frac{\text{grams of Solute}}{\text{mL of Solution}} \right) \times 10^9
\]

\[
\text{ppb (w/v)} = \left( \frac{0.018\text{g of NaF}}{25000 \text{ mL of Solution}} \right) \times 10^9
\]

\= 720 \text{ ppb}

Chemists usually use \textit{molarity} to describe the concentration of a solution.

**Molarity**

Molarity (M) is defined as the number of moles of solute present in each liter of solution.

\[
\text{Molarity (M)} = \left( \frac{\text{moles of Solute}}{\text{liters of Solution}} \right)
\]

Note: The unit M is the same as mole/L.
A solution is prepared by dissolving 0.10 moles of the amino acid alanine in enough water to give a final volume of 75 mL. What is the molarity of the solution?

Molarity (M) = \( \frac{\text{moles of Solute}}{\text{liters of Solution}} \)

Molarity (M) = \( \frac{0.10 \text{ moles}}{0.075 \text{ L}} \)

= 1.3 M or 1.3 moles/L
You try it

A solution is prepared by dissolving 0.57 moles of the adrenaline in enough water to give a final volume of 0.018 L. What is the molarity of the solution?

\[ = 32 \text{ M or } 32 \text{ moles/L} \]
Concentration as a Conversion Factor

- Knowing the concentration allows us to convert between moles of solute and volume of solution.
Molarity as a Conversion Factor

Since Molarity (M) is the *relationship* between moles of solute and volume for a solution, it can be used as a conversion factor.

- convert between *# of moles* present and *volume*!

\[
\text{Molarity (M)} = \left( \frac{\text{# moles}}{\text{Volume (L)}} \right)
\]

\[
\left( \frac{\text{# moles}}{\text{Volume (L)}} \right) \rightarrow \text{Conversion Factors} \rightarrow \left( \frac{\text{Volume (L)}}{\text{# moles}} \right)
\]
Molarity as a Conversion Factor

Example : volume to moles

A glucose solution has a concentration of 0.0500 moles/L. How many moles of glucose are in 0.0250 ml of this solution?

Strategy:

The concentration of 0.0500 moles/L can be used as a conversion factor.

\[
\frac{0.025 \text{ L glucose}}{1 \text{ L}} \times 0.0500 \text{ moles glucose} = 1.25 \times 10^{-3} \text{ moles glucose}
\]
Concentration as a Conversion Factor

Example: volume to mass

A sample of blood serum is tested and found to contain testosterone at a concentration of 0.0075 moles/L. What volume (L) of blood would contain 1.2 moles of testosterone?

160 L
% Concentration as a Conversion Factor

Since % is the relationship between solute and amount of a solution, it can be used as a conversion factor.

- Example % (weight to volume, w/v)
- convert between # of grams present and volume!

\[
\% \text{ (w/v)} = \left( \frac{\# \text{ grams}}{100. \text{ ml solution}} \right)
\]
% Concentration as a Conversion Factor
Example: volume to grams

If low fat milk is 2.00% (w/v) fat, how many grams of fat are contained in 275 mL of low fat milk?

\[
\frac{2.00 \text{ g fat}}{100. \text{ mL milk}} \rightarrow \text{Conversion Factors} \rightarrow \frac{100. \text{ mL milk}}{2.00 \text{ g fat}}
\]

\[
\frac{275 \text{ mL milk}}{100. \text{ mL milk}} \cdot \frac{2.00 \text{ g fat}}{100. \text{ mL milk}} = 5.50 \text{ g fat}
\]
Other Concentration Units
Osmolarity is the number of osmoles per liter of solution.

• An osmole is the number of moles of dissolved particles that are contained in a solution.

\[
\text{Osmolarity} = \frac{\# \text{ osmoles}}{\text{L of solution}}
\]
Concentration in Osmolarity

Example:

One mole of NaCl is equal to 2 osmole. (NaCl ionizes)

\[ \text{NaCl}(s) \rightarrow \text{Na}^+ (aq) + \text{Cl}^- (aq) \]

One mole of sucrose is equal to 1 osmole. (non-ionizing)

\[ \text{Sucrose} (s) \rightarrow \text{Sucrose} (aq) \]

How many osmole(s) in 1 mole of CuCl\(_2\)?

\[ \text{CuCl}_2 \rightarrow \text{Cu}^{2+} + 2 \text{Cl}^- \]
Concentration in Osmolarity

Example: 4.35 moles of CuCl₂ is dissolved in enough water to make 5.80 L of solution. What is the osmolarity of the solution?
Molality $(m)$ is another way to express moles of solute in a solution.

Molality is the number of moles of solute per kg of solvent (moles/kg).

\[
\text{Molality (} m \text{) } = \left( \frac{\# \text{ moles}}{\text{kg of solvent}} \right)
\]
Concentration in Osmolality

- Osmolality is the number of osmoles per kg of solvent.

\[
\text{Osmolality} = \left( \frac{\text{# osmoles}}{\text{kg solvent}} \right)
\]
In the case of dilute aqueous solutions, scientists often use the *approximation*:

\[
\text{Molarity (mol/L)} = \text{Molality (mole/kg solvent)}
\]

This is a very close approximation for dilute solutions because:

1) 1 kg of water = 1 L

2) volume of solvent >> volume of solute therefore:
   - volume of solution ≈ volume of solvent
   - mass of solution ≈ mass of solvent
Concentration in equivalents (Eq/L)

An equivalent (Eq) is the number of moles of charge that a solute contributes to a solution.

\[
(\text{Eq/L}) = \left( \frac{\# \text{ equivalents}}{\text{L of solution}} \right)
\]
Concentration in Equivalents

Example:

One mole of NaCl is equal to 2 equivalents.

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

One mole of sucrose has no equivalents.(non-ionizing)

Sucrose (s) $\rightarrow$ Sucrose (aq)

How many equivalents are in 1 mole of dissolved CuCl$_2$?

4 \[ \text{CuCl}_2 \rightarrow \text{Cu}^{2+} + 2\text{Cl}^- \]

• 2 from Cu$^{2+}$ and 1 from each Cl$^-$
Concentration in Eq/L

Example: 4.35 moles of CuCl$_2$ is dissolved in enough water to make 5.80 L of solution.

• What is the concentration in Eq/L of the solution?

$$(\text{Eq/L}) = \left( \frac{\text{# equivalents}}{\text{L of solution}} \right)$$

4.35 moles CuCl$_2$ | 4 Eq | = 17.4 Eq
------------------- | ---- | = 17.4 Eq
1 mole CuCl$_2$     |      |

$$(\text{Eq/L}) = \left( \frac{17.4 \text{ Eq}}{5.80 \text{ L of solution}} \right) = 3.00 \text{ Eq/L}$$
Milliequivalents (or mEq) per liter is another way to express moles of solute in a volume of solution.

An equivalent (Eq) is the number of moles of charges that one mole of a solute contributes to solution.

Example:

One mole of Ca\(^{2+}\) is equal to 2 Eq of Ca\(^{2+}\)

One mole of Na\(^+\) is equal to 1 Eq of Na\(^+\)
The serum concentration of Ca\(^{2+}\), is sometimes reported in either mmol/L or mEq/L. The normal range of Ca\(^{2+}\) in blood serum is 1.15-1.30 mmol/L. Convert this range into mEq/L.

<table>
<thead>
<tr>
<th>1.15 mmole Ca(^{2+})</th>
<th>2 mEq Ca(^{2+})</th>
<th>=</th>
<th>2.30 mEq Ca(^{2+})</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 L</td>
<td>1 mmole Ca(^{2+})</td>
<td></td>
<td>L</td>
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</tbody>
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<tr>
<th>1.30 mmole Ca(^{2+})</th>
<th>2 mEq Ca(^{2+})</th>
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<td>1 L</td>
<td>1 mmole Ca(^{2+})</td>
<td></td>
<td>L</td>
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</tbody>
</table>
Dilution

• When more solvent is added to a solution, it has been *diluted*.

• **Dilution** is an easy way to reduce the concentration of a solution.

• Dilution is also important in health care because some drugs must be diluted to the proper concentration before being administered.
**Dilution Equation**

Original mass or moles of solute

\[ V_{\text{original}} \times C_{\text{original}} = V_{\text{final}} \times C_{\text{final}} \]

Final mass or moles of solute

\[ V = \text{volume} \]

\[ C = \text{concentration} \]
Example

You begin with 25 mL of a 1.8 M aqueous LiCl solution and add enough water to give a final volume of 35 mL. What is the new concentration?

**Strategy:**

\[ M_1 \times V_1 = M_2 \times V_2 \]
Solution

\[ \frac{M_1V_1}{V_2} = \frac{M_2V_2}{V_2} \]

\[ M_1 = 1.8 \text{ M} \]
\[ V_1 = 25 \text{ mL} \]
\[ V_2 = 35 \text{ mL} \]

\[ M_2 = \frac{M_1V_1}{V_2} = \frac{(1.8 \text{ M})(25 \text{ mL})}{35 \text{ mL}} = 1.3 \text{ M} \]
Group Work

25 mL of a 1.8 M aqueous LiCl solution is diluted to a final volume of 450 mL. What is the new concentration?
Colloids and Suspensions
• Colloids and suspensions are mixtures that resemble solutions, but have some important differences:

• A suspension contains large particles suspended in a liquid.
  • The particles will eventually, slowly settle to the bottom in the absence of agitation

• In a colloid, the particles are larger than the typical solutes in a solution, but smaller than the particles that make up a suspension
  • Therefore they will not settle to the bottom of the container.
Suspensions: Muddy Water

Before settling

Suspension seems homogeneous

Water molecule
Clay particle

After settling

Suspension particles settle out

Water molecule
Clay particle
Suspensions: Orange Juice

- Water molecule
- Orange pulp
Examples of Colloids: Milk

- Milk appears to be homogeneous. But under a microscope you see that milk contains globules of fat and small lumps of the protein casein dispersed in a liquid called whey. Milk is colloid because the particles of casein do not settle out after standing.

Other Examples of Suspensions

Pepto Bismal is a suspension of bismuth subsalicylate particles.

Milk of magnesia is a suspension of magnesium hydroxide particles.
Suspensions scatter light.
Colloids often scatter light
Common Examples of Colloids

Mayonnaise is a colloidal suspension of water and vinegar droplets in edible oil, emulsified by egg yolk.

Other Examples:
fog, liquid foam, whipped cream, hand cream, paint, pigmented ink, solid foam, gelatin, jelly, cheese
## Summary of Solutions, Colloids, and Suspensions

<table>
<thead>
<tr>
<th>Mixture</th>
<th>Particle Size</th>
<th>Appearance</th>
<th>Particle Settling?</th>
<th>Separation of Particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solution</td>
<td>Small (&lt;1 nm)</td>
<td>Clear</td>
<td>No</td>
<td>Particles cannot be separated out by filtration or centrifugation</td>
</tr>
<tr>
<td>Colloid</td>
<td>Intermediate (1 nm–1 μm)</td>
<td>Usually cloudy</td>
<td>No</td>
<td>Particles can usually be separated out by special filtration or centrifugation techniques</td>
</tr>
<tr>
<td>Suspension</td>
<td>Large (&gt;1 μm)</td>
<td>Cloudy</td>
<td>Yes</td>
<td>Particles can be separated out by filtration or centrifugation</td>
</tr>
</tbody>
</table>
Diffusion and Osmosis
Just like gases, liquids can diffuse.

Liquid molecules travel randomly in all directions and mix quickly with other liquid molecules in a process called **diffusion**.

**Diffusion** is the movement of one substance within another substance until it is evenly distributed.
Substances move from areas of higher concentration to those of lower concentration by diffusion.
Certain membranes, called *semipermeable membranes*, are barriers to diffusion because they allow solvents, but not all solutes to pass through.

- The semipermeable membrane in the image allows water to pass but not the Na\(^+\) or Cl\(^-\) ions.
Although water moves back and forth through most semipermeable membranes, it flows more forcefully from the side of the membrane with lesser solute concentration to the side with greater solute concentration.
The net movement of water across a membrane from a solution of lesser solute concentration to one of greater solute concentration is called osmosis.
The pressure associated with the transport of water in the osmosis process is called: **osmotic pressure**.

The greater the difference in solute concentrations across the semipermeable membrane, the greater the osmotic pressure.
**Initial State:**
Equal amounts of liquid are placed on opposite sides of a membrane. Saltwater one side and pure water on the other side.

**Final State:**
Water molecules moved from the chamber with lesser solute concentration (right side) to greater solute concentration (left side).

The water levels change **until** the pressure caused from the difference in water column heights **equalizes the transport of water molecules between each side of the membrane**: This pressure is called osmotic pressure.

- **H₂O** molecule
- **Na⁺** ion
- **Cl⁻** ion
Cell membranes are semipermeable and are, therefore, affected by osmosis.

- In an **isotonic solution**, the solute concentration is the same on each side the cell membrane and the cells retain their normal size and shape.
• In a **hypertonic solution**, the solute concentration is greater outside the cell membrane and the cell shrinks because osmosis pulls water out.
• In a **hypotonic solution**, the solute concentration is greater *inside* the cell membrane and the cell swells because osmosis brings water in.