CHAPTER 9: Solutions

Mixtures and Solutions

REVIEW: A pure substance: a single chemical consisting of only one kind of matter

Two types of pure substances:

- **element:**
  - consists of only one type of atom
  - e.g. carbon (C), hydrogen (H₂), sulfur (S₈), copper wire (Cu)

- **compound:**
  - consists of more than one type of atom and *has a specific formula*
  - e.g., water (H₂O), carbon dioxide (CO₂), and ethanol (C₂H₅OH)

Two or more pure substances can combine to form a **mixture**. Water forms three main types of mixtures when it is mixed with a solid. These are solutions, colloids, and suspensions.

- **solution:** a completely uniform (homogeneous) mixture of two or more substances present as single atoms, ions, or molecules.

- **colloid:** a mostly uniform (homogeneous) mixture made of clusters of atoms, molecules, or ions that have not completely dissolved; these solid particles simply float in the water ... cluster sizes are between 1 nm and 1000 nm.

- **suspension:** like a colloid, except that the clusters of particles are so large that they eventually settle at the bottom of the water (non-uniform) ... cluster sizes are greater than 1000 nm (> 1 µm).

**Solution terminology:**

The term “solution” usually refers to a uniform mixture of a solid in water (like salt in water) or a uniform mixture of a liquid in water (like alcohol in water). Technically, however, a solution is any homogeneous mixture of two or more pure substances, so air (gas-gas) and alloys (metal-metal) are also solutions.

- **solute:** component present in smaller amount
- **solvent:** component present in greater amount

Example: Identify the solute and solvent in each of the following:

a. salt water: solute= _______________ solvent= _______________

b. carbonated water (water + CO₂ gas): solute= _______________ solvent= _______________

Note: Unless otherwise stated, the solvent for most solutions in this class will be water.
The Solution Process

What happens when a solute mixes into a liquid (solvent)? At a molecular level, three things happen:

1. First, atoms or molecules in the solute must move apart so they can mix with the solvent. This means overcoming the intermolecular attractions between solute particles.
2. Second, the solvent particles move apart as well, to make space for the new solute particles. This means also overcoming intermolecular attractions in the solvent.
3. Third, the solute and solvent mix, and form new intermolecular attractions between solvent and solute.

1. **IONIC SOLIDS** in water

As the solute crystal is dropped into a solution, the water molecules begin to attack and pull apart the ions in the ionic lattice. Each ion is surrounded by water molecules (they are now hydrated):

![Ionic SOLIDS diagram](image)

A. Ionic bonds break  
B. Water molecules move apart to accommodate the new Na\(^+\) and Cl\(^-\) ions.  
C. New attractions form between Cl\(^-\) and water molecules, and between Na\(^+\) and water molecules.

2. **COVALENT MOLECULES** in water

When covalent solids, liquids, or gases dissolve with water, the water molecules surround each molecule, BUT ... covalent molecules do not break apart in water. Each covalent molecule stays intact. Water molecules are attracted to any charged (polar) areas on the covalent molecule.

In general, covalent bonds are NEVER BROKEN when substances dissolve. (There is one exception, which we will discuss later: covalent acids.)

![Covalent MOLECULES diagram](image)

Water molecules surround the large molecule sucrose. The negative dipoles on each water molecule attract the positive dipoles on sucrose, and vice versa.
**Solubility**

Don’t confuse how FAST something dissolves (the rate of dissolving) with the total AMOUNT that can be dissolved (the solubility).

1. **solubility**: the amount (grams) of solute that will dissolve in a solvent at a given temperature
2. **rate of dissolving**: how quickly a solute will dissolve in a solvent

**RATE OF DISSOLVING** can be increased by:

1. Heating the solution:
   - solvent molecules move faster → attack solute more frequently
   → break apart solute more quickly

2. Stirring the solution:
   - solvent encounters solute molecules more often
   → solvent breaks apart solute more quickly

3. Grinding the solute into smaller particles:
   - increases the surface area of the solute
   → more solute surface for solvent to attack
   → solvent breaks apart solute more quickly

**Predicting Solubility**

Solubility varies widely among compounds. In general, however, our understanding of intermolecular forces helps us predict what will dissolve in water (a polar solvent) vs. oil (a non-polar solvent).

"Like dissolves like"

→ polar solids dissolve in polar liquids
→ nonpolar solids dissolve in nonpolar liquids
→ but polar and non-polar substances do not dissolve very well in each other

**Polar substances** → ionic compounds, water, ammonia (NH₃), alcohol (ethanol, methanol), CH₃Cl
Non-Polar substances → oil, methane, plastic, tar, and other hydrocarbons (waxes, CH₄), CCl₄

IMPORTANT EXCEPTION: Very small non-polar molecules (O₂, N₂) will dissolve to some extent in water. In addition, some gases actually react with water, increasing their solubility. (This occurs for CO₂.)

**Solubility terms differ for solid solutes vs. liquid solutes:**

miscible: two liquids dissolve together
immiscible: two liquids don’t dissolve together
soluble: solid will dissolve in liquid
insoluble: solid won’t dissolve in liquid
<table>
<thead>
<tr>
<th>polar liquid/solvent</th>
<th>nonpolar liquid/solvent</th>
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<tbody>
<tr>
<td>polar liquid/solvent</td>
<td>Miscible</td>
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<tr>
<td>nonpolar liquid/solvent</td>
<td>Immiscible</td>
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<td>polar solid/solute</td>
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<tr>
<td>nonpolar solid/solute</td>
<td>Insoluble</td>
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<tr>
<td>ionic solid/solute</td>
<td>Check Solubility Rules!</td>
</tr>
</tbody>
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Ex. 1: Indicate if the following are soluble/insoluble in or miscible/immiscible with water:

- vegetable oil (nonpolar) _________________
- vinegar: _________________
- NaCl (table salt): _________________
- HBr (l): _________________
- solid grease: _________________
- AgI: _________________
- C_5H_{12} (l): _________________
- I_2 (s): _________________
- Br_2 (l): _________________
- Mg(OH)_2: _________________

Ex. 2 Circle all of the following that would be soluble in or miscible with vegetable oil:

I_2 (s) KCl P_4 (s) diamond water C_2H_5OH (l)

Hydrophilic versus Hydrophobic

**hydrophilic** ("water loving"):  
- polar molecules or polar end of a molecule that dissolves in water and other polar solvents

**hydrophobic** ("water hating"):  
- nonpolar molecules or nonpolar end of a molecule that does not dissolve in water or other polar solvents

Many larger molecules have sections that are hydrophilic AND other sections that are hydrophobic.
Soaps
This is especially true of soap. Soap consists of a long nonpolar chain of hydrocarbons and a polar end. Explain which end allows soap to dissolve grease and which end allows soap to be water soluble and which end is hydrophilic and which end is hydrophobic.

Factors Affecting Solubility

TEMPERATURE

Rising temperature increases the solubility of most solids in water, but it decreases the solubility of gases in water.

SOLIDS $\Rightarrow$ At higher T, solids are more likely to break apart and be surrounded by water molecules

GASES $\Rightarrow$ At higher T, gas molecules inside the liquid are moving more quickly → they have a higher tendency to find the surface (between liquid and air) → they escape more quickly, leading to fewer gas molecules in the liquid!

Example: We keep soda cold because warm soda holds less dissolve carbon dioxide (less fizz).
PRESSURE (Henry’s Law, Section 9.6)

Rising pressure does not generally affect the solubility of solids in liquid, but it can have a large effect on gases:

Solids → Rising pressure has no effect on solubility
Gases → Rising pressure increases gas solubility

Henry's Law: Solubility of gas is proportional to partial P of gas above liquid
– Higher partial pressures correspond to more gas particles above the liquid
  – More gas particles → Higher solubility of the gas in the liquid

Henry's Law (solubility of gas A in liquid B)

Henry’s Law states that the amount of gas dissolved in a solution is directly proportional to the partial pressure of that gas above the solution.

Increase the pressure of gas A above a solution with solvent B, and more of the gas dissolves.

Mathematically:

\[ P = kc, \]

where \( P = \) partial pressure of the gas above the solution
\( k = \) Henry’s Law constant, a unique, experimentally determined constant for every gas
\( c = \) concentration of the gas in the solution itself

Precipitation vs. Crystallization (vs. solubility)

**Saturated Solution:** When a liquid has dissolved the maximum amount of solid (i.e., the solubility), it is said to be saturated.

**Crystallization:** A previously dissolved solid appears because a (saturated) solution cooled and the solid’s solubility decreased

**Precipitation:** A newly made solid is created by a chemical reaction in a solution, and the new solid happens to not be soluble in water.

When does crystallization occur?
1. If a solution that is saturated with dissolved solid is disturbed (crystal or dust particle added)
2. If a solution is cooled so its solubility is lower than the concentration of previously-dissolved solute.

For some substances, the recrystallization process is **exothermic**

→ The solution releases heat when it recrystallizes
  o hot packs used to warm hands and feet in winter
For some substances, the recrystallization process is endothermic → The solution absorbs heat when it recrystallizes and becomes colder  
  ○ cold packs used for sports injuries

Supersaturation: An unstable, temporary state that sometimes occurs when a liquid temporarily keeps solids dissolved above the solubility limit. It can happen when saturated solutions cool and the solute fails to crystallize.

A supersaturated NaC₂H₃O₂ solution recrystallizing after addition of more solute:

![Image of supersaturated solution recrystallizing after addition of more solute]

### Units of Concentration

\[
\text{concentration} = \frac{\text{amount of solute}}{\text{amount of solution}}
\]

The amount of solute can be expressed in mass or moles, and the amount of solute can be expressed in mass or volume. Either way, the more solute in a given amount of solution → the more concentrated the solution

A. Molarity  
B. Mass/volume or weight/volume percent  
C. Volume/Volume percent  
D. Parts per million

#### A. MOLARITY or MOLAR CONCENTRATION

\[
\text{Molarity} = \frac{\text{moles of solute}}{\text{liters (L) of solution}} \quad \text{(in units of M=molar)}
\]

Note: The molarity of a substance is indicated when it’s in square brackets:

\( [\text{KCl}] = \text{“the molarity or molar concentration of KCl”} \)
Solving Molar Concentration Problems

Ex. 1  Find the molarity of a solution prepared by dissolving 0.250 moles of NaOH in 100.0 mL of solution:

Ex. 2  Find the molarity of a solution prepared by dissolving 12.0 g of KCl in 150.0 mL of solution.

Ex. 3  Find the molarity of a solution prepared by dissolving 25.0 g of NaNO$_3$ in 250.0 mL of solution.

USING MOLAR CONCENTRATION IN CONVERSIONS

Write 2 unit factors for each of the following:

1.  6.00M HCl solution  
2.  0.125M NaCl solution
Keep in mind that if molarity and volume are both given, you can calculate # of moles since:

\[
\text{volume} \times \text{molarity} = \frac{\text{moles of solute}}{\text{liters (L) of solution}}
\]

so volume units will cancels \(\rightarrow\) # of moles

Thus, when given volume and molarity for a solution, multiply them to get # of moles!

**Ex. 1** Calculate the number of moles of \(\text{H}_2\text{SO}_4\) present in 50.0 mL of 6.00 M \(\text{H}_2\text{SO}_4\).

**Ex. 2** Calculate the mass of \(\text{CuCl}_2\) in 25.0 mL of 0.500 M \(\text{CuCl}_2\).

**Ex. 3** Calculate the mass of \(\text{Mg(NO}_3)_2\) in 75.0 mL of 1.50 M \(\text{Mg(NO}_3)_2\).

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**B. MASS/VOLUME PERCENT (% m/v) (also called WEIGHT/VOLUME PERCENT (% w/v))**

\[
\% \text{ m/v} = \frac{\text{mass of solute}}{\text{mL of solution}} \times 100\%
\]
Solving Mass/Volume Percent Concentration Problems

Ex. 1 Calculate the mass/volume percent for a solution that is prepared by dissolving 5.0 g of KI in enough water to give 250 mL of solution.

Ex. 2 If a physiological saline solution is 0.900% (w/v) NaCl. What mass of NaCl is present in 100.0 mL of the solution?

Ex. 3 A topical antibiotic is 1.0% (m/v) Clindamycin. What mass of Clindamycin is present in 60.0 mL of the 1.0% (w/v) Clindamycin solution?

Ex. 4 Intravenous injections of glucose are sometimes administered to patients with low blood sugar. If a normal glucose solution is 5.00% (m/v), what is the volume of solution that contains 1.15 g of glucose?
Ex. 5  Intravenous saline injections are sometimes administered to restore electrolyte balance in trauma patients. What is the volume of solution is required to dissolve 0.350 g of NaCl for a 0.90% m/v NaCl solution? (you may ignore the volume of the NaCl)

Ex. 6  What mass of solute is required to prepare 150.0 mL of a 2.50% (m/v) KCl solution?

Ex. 7  What volume of water is required to dissolve 5.00 g of NaOH to prepare a 10.0% (m/v) NaOH solution?

Ex. 8  A person accused of a DWI violation submitted a 6.00 mL sample of blood for alcohol content analysis. The analysis determined the presence of $5.73 \times 10^{-3}$ g of alcohol in the blood. If a person with a blood alcohol content (mass/volume percent of alcohol in the blood) of 0.08% is considered legally impaired, was this person driving while impaired—i.e. was the blood alcohol content greater than or equal to 0.08%?
Volume/Volume Percent Concentration (% v/v)

% v/v = \frac{\text{volume solute (mL)}}{\text{volume solution (mL)}}

Ex. 1  A solution is prepared by dissolving 5.0 mL of ethanol in 95.0 mL of water.

a. The solute is ________________, and the solvent is ________________.

b. Calculate the % (v/v) concentration.

Ex. 2  How many mL of isopropyl alcohol are needed to prepare 55.0 mL of a 10.0 % (v/v) solution?

Parts per Million Calculations (ppm)

All concentration percentages are simply “parts per hundred”, because the amount of solute has been divided by 100 equivalents of the solvent. In some cases, very small percent concentrations (such as 0.000033%) are more conveniently expressed as parts per million.

\[
\text{ppm} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \quad \text{OR} \quad \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}}
\]

Ex 1:  The maximum allowed concentration of lead in drinking water is 15 ppb (parts per billion). If a glass contains 250 grams of water, how many grams of lead does it contain if its concentration is 15 ppb?
Dilution

**Dilution:** The act of adding water to a solution in order to decrease its concentration.

It is often necessary to dilute a solution, in order to make a less concentrated solution. It is also necessary, at times, to calculate the new molarity when an unintended dilution has occurred.

In all cases, the key to dilution calculations is recognizing that the amount of solute has not changed; only the volume of solution:

\[(\text{amount of solute before}) = (\text{amount of solute after})\]

If the units of concentration are molarity, the amount of solute is given by M x V (molarity times volume).

Thus, \(M_1V_1 = M_2V_2\).

**Ex 1:** Calculate the new concentration when 25 mL of 1.2 M HCl is diluted to 77 mL.

**Ex 2:** How many mL of water are needed to dilute 10.0 mL of 1.0 M solution to 0.77 M?

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**Electrolytes: Ions in solution**

**Electrolyte:** A solid or liquid that breaks apart into ions upon dissolving in solution.

**Non-Electrolyte:** A solid or liquid that does not dissolve to produce ions in solution.

Examples of electrolytes: All soluble ionic solids, acids and bases

Examples of non-electrolytes: Soluble covalent compounds, such as CH\(_3\)Cl, sugar (C\(_{12}\)H\(_{22}\)O\(_{11}\)), alcohol (CH\(_3\)CH\(_2\)OH), etc.

Because electrolytes produced charged particles in solution, such solutions are able to conduct electricity (which is the flow of charged particles). Hence the name: electrolytes.
**Properties of Solutions**

**Colligative Properties:** solution properties that depend on the concentration of dissolved particles

Note that the identity of the dissolved particles does not matter: it is only their concentration. Please know that dissolved particles affect solutions in four key ways:

1. Vapor pressure lowering
2. Raise boiling point
3. Lower freezing point
4. Cause osmosis to occur when a semi-permeable membrane is present

**Membranes, Osmosis, and Dialysis**

**Semi-permeable membrane:** A special membrane that lets solvent through, but not solute

**osmosis:** Net movement of solvent molecules through a semi-permeable membrane from a low concentration to a high concentration

Osmosis occurs whenever two solutions, connected by a semi-permeable membrane, have different concentrations.

Another way to think of osmosis is to realize that when the concentration of **solute** is low, the concentration of **solvent** is high. Then we can think of osmosis as the flow of solvent from an area of high concentration to low concentration. (from an area with lots of solvent → to area with less solvent)

![Osmosis Diagram]

At the molecular level:

Note that **only the solvent molecules** can move through the semi-permeable membrane.
The different in height between the two liquids is a measure of the opposing pressure generated by the concentration differences.

→ This pressure is called the osmotic pressure and given the symbol $\Pi$ (this is different from the irrational number “$\Pi$”, which is also confusingly given the same Greek symbol, but in lowercase: $\pi$)

**Osmosis Terminology:**

- **isotonic:** when two solutions have equal osmotic pressures
- **hypertonic:** the more concentrated solution of two solutions with different osmotic pressures
- **hypotonic:** the less concentrated solution of two solutions with different osmotic pressures

**osmolarity:** osmolarity = $i \cdot M$, where $i$ is the number of particles the solute breaks into when it dissolves.

Examples:
- NaCl breaks into $Na^+$ and $Cl^-$, so $i = 2$ and a 1.0 M solution is $2 \cdot 1.0 \text{ M} = 2.0$ osmolar.
- CaCl$_2$ breaks into $Ca^{2+}$ and 2 $Cl^-$, so $i = 3$ and a 1.0 M solution is $3 \cdot 1.0 \text{ M} = 3.0$ osmolar.
- Sucrose does not break apart upon dissolving, so $i = 1$ and a 1.0 M solution is 1.0 osmolar.

**osmotic pressure ($\Pi$):** pressure required to stop osmosis

$$\Pi = iMRT$$

where $i$=# of moles of particles per mole of solute, equal to 1 for non-electrolytes like sugar,

$M$=molarity  $R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$  $T$=temp in K

Ex. 1: A 1.25M sample of seawater was taken at 25˚C. Calculate the osmotic pressure of the seawater sample. Let $i = 1$. 

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CHEM 121  page 15 of 17
Ex. 2: A solution is prepared by dissolving 35.0 g of hemoglobin (Hb) in enough water to make up 1.00 L in volume. If the osmotic pressure of the solution is found to be 10.00 mmHg at 25˚C, calculate the molar mass of hemoglobin. (Let i=1.)

If we make the external pressure large enough (i.e. for \( P > \Pi \)), we can make water molecules go from a solution to pure water.

This process is called reverse osmosis and is often used to obtain drinking water.

Dialysis

dialyzing membrane: similar to an osmotic membrane, but with larger pores to allow both solvent molecules and some small molecules through. It blocks the passage of larger molecules and colloidal particles.

dialysis: net movement of solute molecules through a semi-permeable membrane, from a high concentration to a low concentration of solute.
Dialysis for Kidney Malfunction
For people suffering from kidney malfunction, a similar technique is used to clean the blood.

- Blood is pumped through a tube made of a dialyzing membrane.
- The tubing passes through a bath to collect impurities, leaving blood proteins and large molecules in the blood.