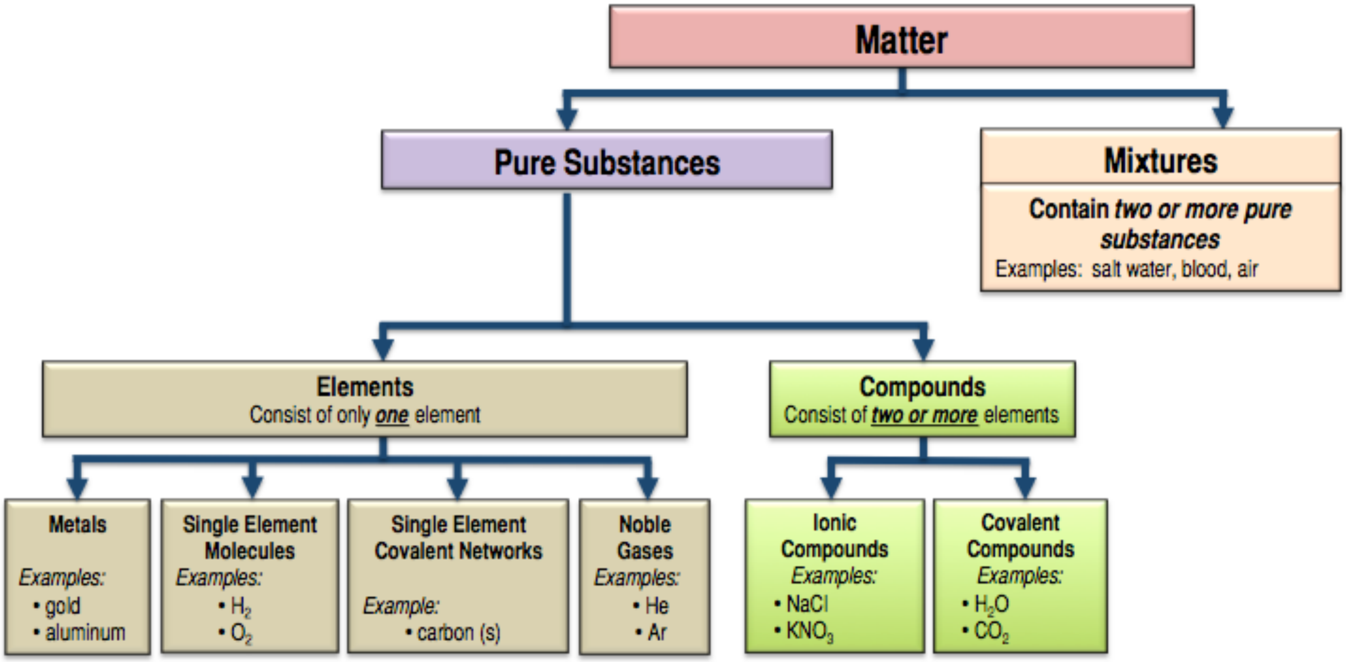
Chapter 7 Lecture Notes: Solutions, Colloids, and Suspensions

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.
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.



Most matter on earth is **not** composed of just one pure substance.

When ***two or more pure substances*** are combined, we refer to the combination as a ________________.

In this chapter, you will learn about **three** types of ***mixtures***:

1) **solutions**
2) **colloids**
3) **suspensions**

Macro-Scale Classification of Mixtures: Homogeneity

One way in which mixtures are classified is by their macro-scale, ***visually*** observed **homogeneity**.

A _____________________ ***mixture*** appears to be the same throughout the entire sample/object.

- Example: Sugar dissolved in water

A _____________________ ***mixture*** has visible regions of varying composition.

a. Example: A chocolate chip cookie

Introduction to Solutions

_____________ are **mixtures of pure substances** in which the pure substance particles (molecules, ions, or noble gas atoms) are __________ distributed throughout the entire volume of the mixture.

The pure substance that is in the **greatest abundance** is referred to as the _____________.

- Typically, especially in biological systems, ***the solvent is water***.

The other pure substance components of a solution are called _____________.

With very few exceptions, the solution takes the same *physical phase* (gas, liquid, or solid) as the *solvent*.

The ***solutes*** are said to be "_____________" in the ***solvent***.

Solutions can be **gas-phase**, **liquid-phase**, or **solid-phase**.

An example of a **gas-phase solution** is air.

- Air is a mixture of several gases, mostly nitrogen, oxygen, and carbon dioxide.

An example of a **solid-phase solution** is brass.

- Brass is a mixture of copper and zinc. When one solid (such as zinc) is evenly dispersed in another solid (such as copper), the solid solution is called an ____________.

We will focus on **liquid-phase solutions** for the remainder of this course.

- There are 3 types of *liquid-phase solutions:*
  - ***Gas*** *dissolved in a* ***liquid*** solvent
    - An example of a ***gas dissolved in a liquid solvent*** is CO_2 dissolved in water; this is how beverages are carbonated. Another example of a gaseous solute dissolved in liquid is O_2 dissolved in water; fish extract the O_2 from water using their gills. Whenever a gas is present above a liquid, some of the gas will dissolve in the liquid.
  - ***Liquid*** *dissolved in a* ***liquid*** solvent
    - An example of a ***liquid dissolved in a liquid solvent*** is ethyl alcohol dissolved in water; this is the basis of adult beverages.
  - ***Solid*** *dissolved in a* ***liquid*** solvent
    - An example of a ***solid dissolved in a liquid solvent*** is table salt (NaCl) or table sugar (sucrose molecules) in water.

Almost all of the solutions used in biological applications, such as in biomedical analytical labs and research labs, are *liquid-phase solutions*.

For the remainder of this course, I will use the term **solution** to mean **liquid-phase solution.**

Understanding Check

If the following pairs of pure substances are mixed in the ratios given to form solutions, identify each substance as either **solvent *or* solute**:

a) 5 grams of sodium chloride and 100 grams of water

b) 10 mL of ethyl alcohol and 250 mL of water

c) 100 mL of acetone and 10 mL of water

The Dissolution Process: Solvation

Another term for "*dissolving*" is ___________________.

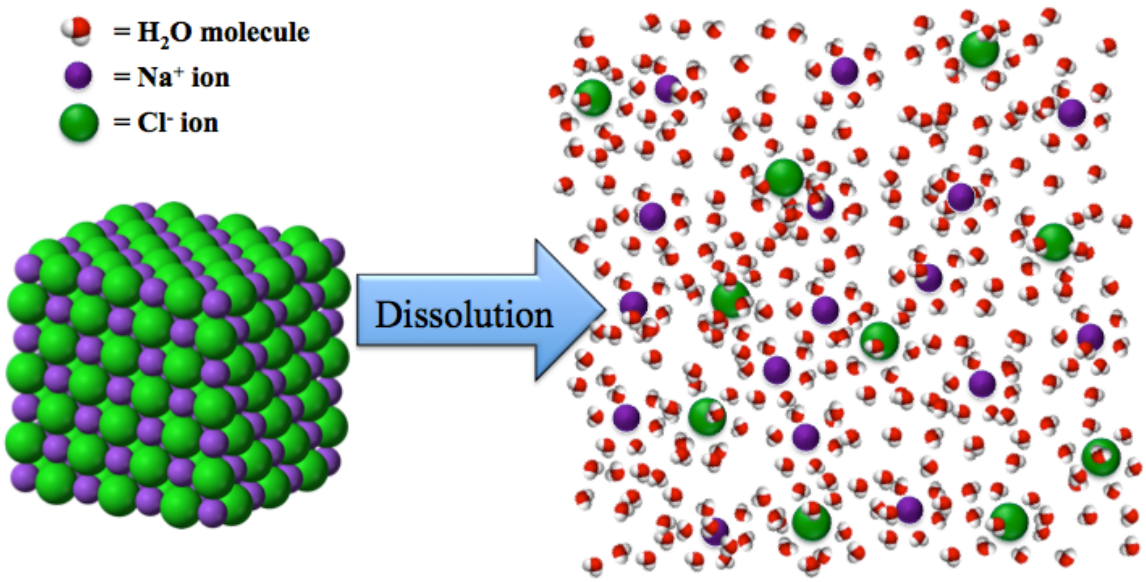
In liquid and solid phase solutions, the solute and solvent particles are ______________ to each other by one or more of the five types of ***noncovalent interactions***.

The stronger the **solute-solvent** interactions, the more *solute* that can be dissolved.

The Dissolution of Ionic Compounds

Example: The dissolution of sodium chloride in water.

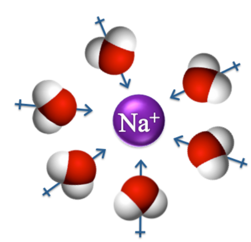
When ionic compounds dissolve, the ions __________________ (become separated from each other) and enter the liquid phase solution.



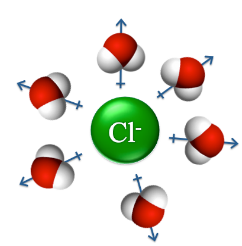
Sodium Chloride Crystal **Dissolved Sodium Chloride, NaCl (*aq*)**

The water molecules and ions are attracted to each other through ***ion-dipole forces***.

The ***partially negatively-charged ends*** of several water molecule dipoles are attracted to the ***positive charge*** of each sodium.



Likewise, the ***partially positively-charged ends*** of several water molecule dipoles are attracted to the ***negative charge*** of each chloride ion.



We use the term ______________ to describe a *solute* particle becoming surrounded by *solvent* molecules.

Solvation causes dissolution of the solute.

The Dissolution of Molecules

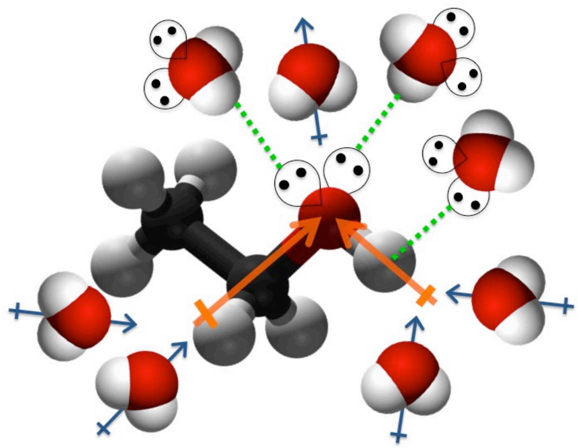
Molecules will dissolve in a particular solvent when the **solute-solvent** noncovalent interactions are strong enough to overcome the solute-solute interactions.

Unlike ionic compounds that ***dissociate***, when molecules dissolve, the ____________ _____________ become ***solvated***.

Example: The dissolution of ethyl alcohol in water.

Ethyl alcohol dissolves in water because it has significantly strong noncovalent interactions with water.

These **solute-solvent** intermolecular forces enable the solvation of ethyl alcohol molecules by water molecules as illustrated in below.



All *three* of the intermolecular forces (hydrogen bonding, dipole-dipole forces, and London forces) occur between ethyl alcohol and water.

Since water molecules are relatively small, the London forces between water and ethyl alcohol are not very strong; if they were the only intermolecular forces present, ethyl alcohol would not dissolve in water.

Since ethyl alcohol *and* water are *polar molecules*, they can also interact through dipole-dipole forces.

- The dipoles of ethyl alcohol's **highly-polar bonds** are indicated by arrows on top of the ethanol structure, and water's molecular dipoles are indicated by dipole arrows behind the water molecules.

Because of the relatively small size of solvated molecules and ions, the forces imparted upon them from *collisions* with solvent molecules are much greater than the force of gravity.

It is for this reason that solvated ions and molecules do not settle to the bottom of a mixture, but instead, move in random directions in-between collisions.

This type of random, chaotic movement is called **Brownian motion** and results in the solute being ***evenly dispersed*** within the solvent.

Since ions and molecules are **evenly dispersed** within the solvent, ***solutions are*** _________________.

S*olutes* that are polar are capable of dissolving in polar *solvents*.

S*olutes* that are nonpolar are capable of dissolving in nonpolar *solvents*.

This phenomenon is summarized in the easily-remembered phrase,
"*like dissolves like*."

Understanding Check

If you dissolved ***one mole*** of the following substances in water, how many moles of solvated ions would be present (include both cations and anions in the number of moles of solvated ions)?

a. NaCl

b. $BaCl_2$

Understanding Check

List ***<u>all</u>*** of the **noncovalent interactions** that can occur *between solute and solvent* for each of the following solutions:

a. oxygen gas dissolved in water

b. carbon dioxide gas dissolved in water

c. potassium iodide dissolved in water

d. pentane dissolved in octane

Solubility

Solubility is a term that refers to the _____________ ***of solute*** that can be dissolved.

Some *liquid-in-liquid* solutions can be made at _______ ratio of the liquids.

- For example, water and ethyl alcohol will mix no matter what the ratio is of water to ethyl alcohol.

When two liquids mix with each other in ***any*** ratio, we say that the substances are "_____________."

Some pairs of liquids ***will not*** mix with each other at all.

- For example, oil will not significantly dissolve in water.
- This is why we see oil floating on the top of water when oil spills occur.

When two liquids ***will not mix*** with each other we say that the substances are "_____________."

For most solute/solvent pairs, there is a limit on how much solute can dissolve in a particular solvent.

a) For example, you can only dissolve so much salt or sugar in water.

At some point, the solution becomes _____________ and the amount of dissolved solute cannot increase.

If you continue to add a solid solute to a **saturated solution**, ***the excess solute will exist as a solid*** in the container.

The Solubility of *Gases* in Water

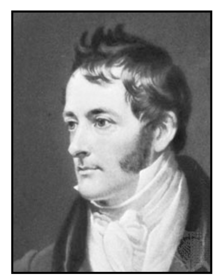
The solubility of a ***dissolved gas*** depends on *both* _______________ and ______________.

Whenever a gas is present above a liquid, some of the gas will dissolve in the liquid.

The higher the partial pressure of a particular gas above a liquid, the more of that gas will dissolve in the liquid.

William Henry was first to report that the amount of gas dissolved in a liquid is directly proportional to the *partial pressure* of the gas.

This relationship between the amount of gas dissolved and pressure is known as "Henry's Law."



William Henry

The **lower** the _______________ of the aqueous solutions, the **greater** the ***solubility of gases***.

- For example, cold water can dissolve more oxygen than warm water.

The Solubility of *Solids* in Water

The solubility of a _________ ***solute*** depends on _______________.

- As the temperature of water increases to its normal boiling point (100°C), the solubility of most ***solid*** *solutes* ____________.

Notice that this is the _______________ of the behavior of ***gaseous*** *solutes*.

The Solubility of *Ionic Compounds* in Water

Some ionic compounds dissolve to a significant extent in water; some do not.

Ionic compounds that **do not** significantly dissolve are categorized as "***water*** ________________."

- *Water insoluble* compounds exist in their crystal/solid form when placed in water.

Ionic compounds that dissolve to a significant extent are classified as "***water*** _____________."

It is convenient to use "**solubility rules"** in order to know which ionic compounds are ***water soluble*** and which ones are ***water insoluble***.

In this table, ionic compounds are first classified as ***water soluble*** or ***water insoluble*** based on their _____________, and then exceptions are listed.

Solubility Rules Table

Water Soluble			
Compound	**Example**	**Exceptions**	**Exception Example**
Nitrates	$NaNO_3$	None	None
Chlorides, Bromides, and Iodides	$NaCl$	Compounds containing Ag^+, Pb^{2+}, or Hg^+, and HgI_2	$AgCl$
Sulfates	K_2SO_4	Compounds containing Pb^{2+}, Sr^{2+}, Ba^{2+}, or Hg^+	$PbSO_4$
Water Insoluble			
Compound	**Example**	**Exceptions**	**Exception Example(s)**
Hydroxides	$Mg(OH)_2$	Compounds containing alkali (Group I) metals *or* Ca^{2+}, Sr^{2+}, Ba^{2+}, NH_4^+	$NaOH$
Phosphates, Carbonates, and Chromates	$FePO_4$	Compounds containing alkali (Group I) metals *or* NH_4^+	K_2CO_3, Li_3PO_4, Na_2CrO_4

Example: Is KNO_3 water soluble?

Solution:

STEP 1: Find the solubility classification in the table based on ***the identity of the anion***.

- We see that ***nitrates*** are in the **water soluble class**.

STEP 2: Check to see if the compound's ***cation*** causes the compound to be an **exception** for the solubility class.

- There are *no exceptions* for ***nitrates***.

This means that ***all nitrates are water soluble***, therefore KNO_3 is ***water soluble***.

Solubility Rules Table

Compound	Example	Exceptions	Exception Example
Water Soluble			
Nitrates	$NaNO_3$	None	None
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Example: Is $Cu(OH)_2$ water soluble?

Solution:

STEP 1: Find the solubility classification in the based on ***the identity of the anion***.

- We see that ***hydroxides*** are in the **water insoluble class**.

STEP 2: Check to see if the ***cation*** causes the compound to be an **exception** for the solubility class.

- There are exceptions for hydroxides, however Cu^{2+} ***is not* one of them**.

Therefore $Cu(OH)_2$ is <u>**not**</u> ***water soluble***.

- If $Cu(OH)_2$ were mixed with water, it would exist as solid crystals submerged in water.

Solubility Rules Table

Compound	Example	Exceptions	Exception Example
Water Soluble			
Nitrates	$NaNO_3$	None	None
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Example: Is $BaSO_4$ water soluble?

Solution:

STEP 1: Find the solubility classification in the table based on ***the identity of the anion***.

- We see that ***sulfates*** are in the **water soluble class**.

STEP 2: Check to see if the ***cation*** causes the compound to be an **exception** for the solubility class.

- There are some exceptions for sulfates and Ba^{2+} ***is one of them***.

Therefore $BaSO_4$ is <u>**not**</u> ***water soluble***.

Solubility Rules Table

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Example: Is KOH water soluble?

Solution:

STEP 1: Find the solubility classification in the table based on ***the identity of the anion***.

- We see that ***hydroxides*** are in the **water insoluble class**.

STEP 2: Check to see if the ***cation*** causes the compound to be an **exception** for the solubility class.

- There are exceptions for hydroxides, and K^+ ***is one of them***; it's an alkali (Group I) metal.

Therefore KOH is ***water soluble***.

Solubility Rules Table

Water Soluble			
Compound	**Example**	**Exceptions**	**Exception Example**
Nitrates	$NaNO_3$	None	None
Chlorides, Bromides, and Iodides	NaCl	Compounds containing Ag^+, Pb^{2+}, or Hg^+, and HgI_2	AgCl
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Understanding Check

Determine which of the following compounds is water soluble.

a. potassium iodide

b. iron(II) nitrate

c. copper(II) hydroxide

d. silver bromide

e. sodium sulfate

f. potassium hydroxide

g. lead(II) chromate

h. ammonium hydroxide

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Electrolytes

Solutions that contain dissolved ***ions*** are capable of conducting electricity and are sometimes referred to as __________________ ***solutions***.

- Dissolved ionic compounds are called ***electrolytes***.

Electrolyte solutions are required in biological functions such as the transmission of nerve impulse signals and muscle actuation.

Our bodies obtain electrolytes from food and drink.

Reactions of Ions in Aqueous Solutions

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:

$$AX + BY \rightarrow AY + BX$$

There are two types of double replacement reactions:

1) **Precipitation Reactions**
2) **Gas Producing Reactions**

1) Precipitation Reactions

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

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

If one of the new pairs formed is______________ a new substance (solid/precipitate) is formed.

For a precipitation reaction to occur, at least one of the ____________ 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 ________________ 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 that was just demonstrated; the reaction of lead(II) nitrate and potassium chromate.

Step 1: Write ***reactants' names*** and arrow for the chemical equation using ***word form*** (not formulas):

lead(II) nitrate + potassium chromate →

Step 2: Add the "*possible*" **products** to the word equation by switching ***anions***:

lead(II) nitrate + potassium chromate → lead(II) chromate + potassium nitrate

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

lead(II) nitrate + potassium chromate → lead(II) chromate + potassium nitrate

$$Pb(NO_3)_2 + K_2CrO_4 \rightarrow PbCrO_4 + KNO_3$$

- **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:

$$Pb(NO_3)_2 + K_2CrO_4 \rightarrow PbCrO_4 + 2\ KNO_3$$

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*)."

$$Pb(NO_3)_2\,(aq) + K_2CrO_4\,(aq) \rightarrow PbCrO_4\,(s) + 2\ KNO_3\,(aq)$$

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).

sodium chloride + silver nitrate →

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

sodium chloride + silver nitrate → sodium nitrate + silver chloride

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

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

Step 4: ***Balance*** the equation:

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

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

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

$$NaCl\,(aq) + AgNO_3\,(aq) \rightarrow NaNO_3\,(aq) + AgCl\,(s)$$

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 formula):

sodium chloride + potassium nitrate →

Step 2: Add the "*possible*" **products** to the word equation by switching ***anions***:

sodium chloride + potassium nitrate → sodium nitrate + potassium chloride

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

$$NaCl + KNO_3 \rightarrow NaNO_3 + KCl$$

Step 4: ***Balance*** the equation:

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

$$NaCl + KNO_3 \rightarrow NaNO_3 + KCl$$

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

$$NaCl\,(aq) + KNO_3\,(aq) \rightarrow NaNO_3\,(aq) + KCl\,(aq)$$

IMPORTANT: If **both** of the "possible" products ***are water soluble***, then **no reaction occurred**.

- There were solvated cations and anions in each 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:

$$NaCl\,(aq) + KNO_3\,(aq) \rightarrow \textbf{No Reaction}$$

You try one: Determine if a precipitation reaction would occur when a silver nitrate solution is mixed with a barium chloride solution ***<u>and</u>***, 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.

The gas producing double replacement reaction that is typically encountered in the health sciences field and, therefore the only gas producing reaction which I would like you to be familiar, is the reaction of aqueous hydrogen monochloride (**HCl**, also know as hydrochloric acid) and aqueous sodium bicarbonate (**$NaHCO_3$**).

$$HCl + NaHCO_3 \rightarrow HHCO_3 + NaCl$$

In this reaction, the ***bicarbonate*** and **chloride** anions switch partners to form aqueous ***carbonic acid*** (**$HHCO_3$**) and ***sodium chloride***.

- In the chemical equation ***above***, I wrote the formula of carbonic acid as **$HHCO_3$** in order to help you see how **Cl^-** and **HCO_3^-** "*switched partners*"; however the correct way to write the formula for carbonic acid is **H_2CO_3**, as described below.

$$HCl\,(aq) + NaHCO_3(aq) \rightarrow \boxed{HHCO_3}(aq) + NaCl\,(aq)$$

carbonic acid

$$HCl\,(aq) + NaHCO_3(aq) \rightarrow \boxed{H_2CO_3}(aq) + NaCl\,(aq)$$

Carbonic acid _______________ to $H_2O\ (l)$ and $CO_2(g)$

The overall gas producing double replacement reaction equation is written in the grey box on the right.

$$HCl\,(aq) + NaHCO_3(aq) \rightarrow \boxed{H_2O\ (l) + CO_2\ (g)} + NaCl\,(aq)$$

Gas Producing Double Replacement Reaction

This particular gas producing reaction is important in medicine because sodium bicarbonate is used as an over-the-counter therapeutic agent to treat acid indigestion (heartburn).

- Sodium bicarbonate is the primary active ingredient in many antacids, such as alka-seltzer.
- Sodium bicarbonate "neutralizes" acid in the stomach to produce water, carbon dioxide gas, and salt.

You will learn much more about acids in later chapters.

The Solubility of Organic Molecules

Molecules will dissolve in a particular solvent when the **solute-solvent** noncovalent interactions are strong enough to _________________ the solute-solute interactions.

The ***more*** **solute-solvent** noncovalent attractive interactions that can occur, the more solute that can be dissolved.

It is for this reason that ***polar solutes*** are capable of dissolving in ***polar solvents***, and ***nonpolar*** solutes are capable of dissolving in ***nonpolar solvents***, as summarized by the phrase "***like dissolves like***."

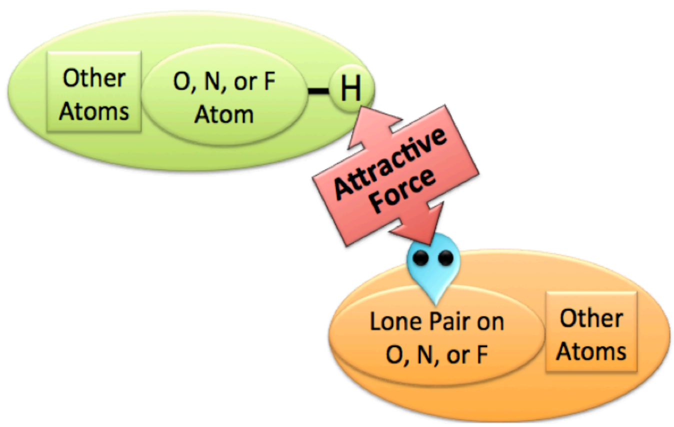
For an organic molecule to have significant **water solubility**, it must be ____________ and/or capable of _______________ ____________ with water.

Dipole-Dipole

Electrostatic attractive force between *two* ***polar*** *molecules.*

Hydrogen Bonding

Electrostatic attractive force between the partially positive charged hydrogen end of an O-H, N-H, or F-H bond and the negative charge of a lone pair on an O, F, or N.



We can use the general rule of "**like dissolves like**" to predict the *relative* ***water solubilities*** of various organic solute molecules.

Let's consider the water solubility of the ***organic molecule families*** that I introduced you to in chapter 4: **hydrocarbons**, **alcohols**, **carboxylic acids**, and **esters**.

Water Solubility of Hydrocarbons

Hydrocarbons are ______________ molecules.

They are **not** capable of *hydrogen bonding* or *dipole-dipole interactions*, therefore they are ______ significantly soluble in water.

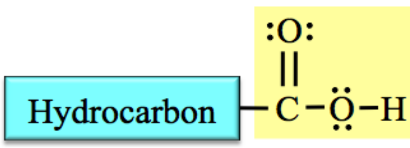
Water Solubility of Alcohols, Carboxylic Acids, and Esters

Alcohols and many other families of organic molecules are attracted to water through hydrogen bonding and/or dipole-dipole interactions.

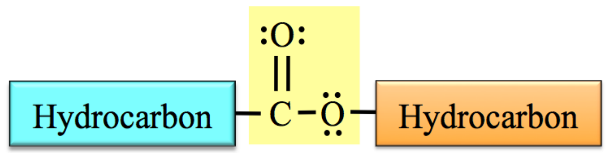
The *general form* of an ***alcohol*** molecule is:

As the ____________ _______ of various alcohol molecules gets ***larger***, *the water solubility* ___________.

This trend of decreasing solubility as the hydrocarbon part of organic molecules gets larger is also seen in **carboxylic acids**, **esters**, and all of the other types of organic molecules that you will be introduced to in later chapters.



general form of a ***carboxylic acid***



general form of an ***ester***

As the hydrocarbon part of a molecule gets __________, London forces become more important (stronger), the molecule becomes ______ polar, and the organic molecules are *more attracted to each other* than they are to water molecules.

- When this occurs, it is *lower in energy* for the organic molecules to be surrounded by other organic molecules and therefore the water solubility drastically decreases.

The table below shows the trend in decreasing water solubility for some alcohol molecules as their hydrocarbon part gets larger.

Molecule Name	Condensed Structure	Solubility in Water
methanol	CH_3OH	miscible in any ratio with water
ethanol	CH_3CH_2OH	miscible in any ratio with water
1-propanol	$CH_3CH_2CH_2OH$	miscible in any ratio with water
1-butanol	$CH_3CH_2CH_2CH_2OH$	slightly soluble
1-pentanol	$CH_3CH_2CH_2CH_2CH_2OH$	insoluble

The Solubility of Organic Molecules in *Non Aqueous* Solutions

Not all solutions involve water as the solvent.

Non water ***polar solvents*** behave quite like water in regard to their ability to dissolve polar solutes better than nonpolar solutes.

On the other hand, ***nonpolar solvents*** dissolve ***nonpolar*** ____________ more readily than polar solutes; ***like dissolves like***.

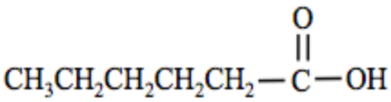
- For example, *pentane* **cannot** be significantly dissolved in water, however it **can** be dissolved in *hexane*.

pentane $CH_3—CH_2—CH_2—CH_2—CH_3$

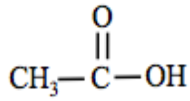
hexane $CH_3—CH_2—CH_2—CH_2—CH_2—CH_3$

Understanding Check

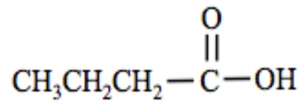
a. List the following carboxylic acids in order of ***increasing solubility in water*** (least soluble to most soluble).



hexanoic acid

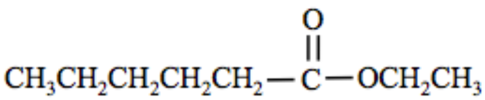


ethanoic acid

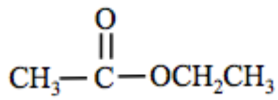


butanoic acid

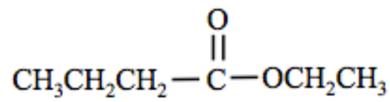
- List the following esters in order of ***increasing solubility in water***.



ethyl hexanoate



ethyl ethanoate



ethyl butanoate

Concentration of Solutions

The term "__________________" refers to the ***amount of a solute in a solution***.

Other *qualitative* (non numeric) descriptions of the amount of solute are sometimes used:

- A solution with a relatively large amount of solute is said to be "__________________."
- A solution with a relatively small amount of solute is said to be "______________."

In this video you will see how to *quantitatively* (with numeric values) describe the ***amount of solute that is present in a solution***.

The ***concentration*** of a solution is the _______________ ______________ of solute that is dissolved in a particular quantity of solution (or solvent).

Various Units Used for Concentration

Percent (%) Concentration

There are three different percent (%) concentration units that are commonly used:

percent weight to weight: % (w/w)
percent volume to volume: % (v/v)
percent weight to volume: % (w/v)

Percent *Weight to Weight*

Percent *weight to weight* - % (w/w) - is defined as the ratio of the **mass** of the ***solute*** to the **mass** of the _______________ ______________, multiplied by **100**:

$$\%\ (\text{w/w}) = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times 100$$

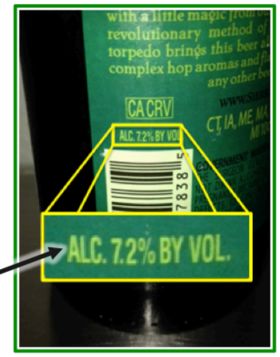
Example: A **10% (w/w)** sodium chloride solution contains ***10 grams*** of dissolved sodium chloride in every ***100 grams of solution***.

Percent weight to weight is also referred to as "*mass percent*" or "*gram percent*."

Percent *Volume to Volume*

Percent *volume to volume* - % (v/v) – is defined as the ratio of the volume of the ***solute*** to the volume of the ***entire solution***, multiplied by **100**:

$$\%\ (\text{v/v}) = \left(\frac{\text{volume of } \textit{solute}}{\text{volume of } \textit{solution}}\right) \times 100$$



Example: Percent volume to volume is commonly used to indicate the concentration of alcohol in adult beverages where % (v/v) is referred to as "*alcohol by volume* (ABV)" or "*alcohol percent by volume*."

In this beer, there are **7.2 mL** of ***alcohol*** for every **100 mL** of ***beer***

Percent *Weight to Volume*

Percent *weight to volume* - % (w/v) - is defined as the number of grams of ***solute*** contained in **100 mL** of ***solution.***

% (w/v) is calculated by multiplying the ratio of the grams of the ***solute*** to the volume (mL) of the ***entire solution***, by **100**.

$$\%\ (\text{w/v}) = \left(\frac{\text{grams of } \textit{solute}}{\text{mL of } \textit{solution}}\right) \times 100$$

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?

$$\%\ (\text{Weight/Vol}) = \left(\frac{\text{grams of Solute}}{\text{mL of Solution}}\right) \text{ x } 100$$

$$\%\ (\text{Weight/Vol}) = \left(\frac{2.0 \text{ g of KI}}{75 \text{ mL of Solution}}\right) \text{ x } 100$$

$$= \mathbf{2.7\ \%\ (w/v)}$$

Percent *Weight to Weight* Example

What is the % (w/w) concentration of a sodium chloride solution prepared by adding 5.0 grams of sodium chloride to 130.0 g of water?

Solution: Calculate the concentration using the equation/definition of % (w/w):

$$\%\ (\text{w/w}) = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times 100$$

a. The mass of the ***solute*** (sodium chloride) was given: **5.0 g**
b. The mass of the ***solution*** is equal to the mass of the ***solute*** (NaCl, 5.0 g) ***plus*** the mass of the ***solvent*** (water, 130.0 g) = **135.0 g**

$$\%\ (\text{w/w}) = \left(\frac{5.0\ \text{g}}{135.0\ \text{g}}\right) \times 100 = 3.7\ \%\ (\text{w/w})$$

In a 3.7 % (w/w) solution, there are 3.7 g of ***solute*** contained in every 100 g of ***solution***.

Note that in this % (w/w) problem, the mass units cancel, therefore we can use any mass units as long as we use the same unit for the mass of solute **and** mass of solution.

Understanding Check: What is the **% (v/v)** concentration of a solution prepared by adding 27 mL of alcohol to enough water to make 552 mL of solution?

Concentration in Parts per Thousand, Parts per Million, Parts per Billion, Parts per Trillion

Parts per thousand, parts per million (ppm), parts per billion (ppb), and parts per trillion (ppt) are defined in a similar way as percent concentration however, _____________ of multiplying the ratios of solute to solution by **100**, they are multiplied by a **thousand**, **million**, **billion**, or **trillion**, respectively.

Just like percent (%) concentration, parts per thousand, ppm, ppb, and ppt units can be **(w/w)**, **(v/v)**, or **(w/v)**.

For example, the definitions or percent, parts per thousand, ppm, ppb, and ppt, ***in*** **(w/w)**, are:

$$\%\ (\text{w/w}) = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times 100$$

$$\text{parts per thousand (w/w)} = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times 1000$$

$$\text{ppm (w/w)} = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times \left(1 \times 10^{6}\right)$$

$$\text{ppb (w/w)} = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times \left(1 \times 10^{9}\right)$$

$$\text{ppt (w/w)} = \left(\frac{\text{mass of } \textit{solute}}{\text{mass of } \textit{solution}}\right) \times \left(1 \times 10^{12}\right)$$

The ppm, ppb, and ppt units are often used for very dilute solutions.

If you were to mix ***one drop*** of alcohol into an Olympic-size pool such as the one shown here, the alcohol concentration would be about **20 part per trillion** (**ppt**).

The volume of an Olympic-size swimming pool is about 2.5×10^9 mL. The volume of one drop of ethyl alcohol is about 0.050 mL. What is the ppt (v/v) concentration of alcohol if 0.050 mL (~1 drop) is mixed into a pool with a volume of 2.5×10^9 mL?

$$\text{ppt (v/v)} = \left(\frac{\text{volume of } \textit{solute}}{\text{volume of } \textit{solution}}\right) \times \left(1 \times 10^{12}\right)$$

- Volume of the ***solute*** (alcohol in this example) was given: **0.050 mL**
- The volume of the ***solution*** is equal to the volume of the ***solute*** (alcohol, 0.050 mL) ***plus*** the volume of the ***solvent*** (water, 2.5×10^9 mL) = **2.5×10^9 mL**

Insert the volume of the solute and the volume of the solution into the equation for ppt (v/v):

$$\text{ppt (v/v)} = \left(\frac{0.050 \text{ mL}}{2.5 \times 10^{9} \text{ mL}}\right) \times \left(1 \times 10^{12}\right) = 20.\ \text{ppt (v/v)}$$

Understanding Check: The legal limit of blood alcohol concentration while driving in most states corresponds to *about* 0.080 grams of alcohol per 100.0 grams of blood (solution).

What is the **parts per thousand (w/w)** blood alcohol concentration at this legal limit?

Concentration in Molarity

Chemists often use a concentration unit of measure called ____________.

Molarity is defined as the number of __________ ***of solute*** per _________ ***of solution***.

It can be calculated by taking the ratio of moles of solute to the volume (in liters) of solution:

$$\text{Molarity} = \left(\frac{\text{moles of } \textit{solute}}{\text{liters (L) of } \textit{solution}} \right)$$

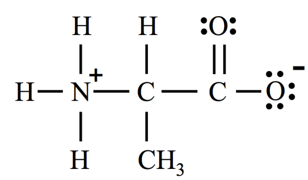
Calculating a solution's molarity using this equation/definition results in units of **moles/L**.

The mole/L unit is often referred to as __________ and abbreviated as "___."

Keep in mind, the "**M**" *can be interchanged with* "**mole/L**."

- For example, 0.030 **moles/L** = 0.030 **M** (molar)

Example: A solution is prepared by dissolving 0.10 moles of the amino acid alanine in enough water to give a final volume of 0.075 L. What is the ***molarity*** of the solution?



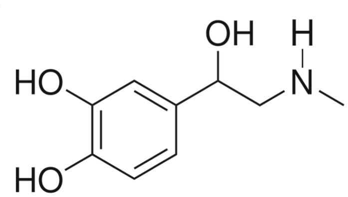
alanine

$$\text{Molarity (M)} = \left(\frac{\text{moles of solute}}{\text{liters of solution}} \right)$$

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

$$= 1.3 \text{ M or } 1.3 \text{ moles/L}$$

Understanding Check: A solution is prepared by dissolving 0.057 moles of adrenaline in enough water to give a final volume of 1.80 L. What is the ***molarity*** of the solution?



adrenaline

In some **molarity** calculations, since mass (grams) is the parameter that is *directly measurable*, you will begin with the *number of* ***grams*** *of solute*.

When this is the case, use the solute's molar mass to convert from grams to moles.

grams of solute $\rightarrow$ moles of solute

$$\text{Molarity (M)} = \left(\frac{\text{moles of solute}}{\text{liters of solution}}\right)$$

You will need to do this in the following *Understanding Check* problem.

Understanding Check: What is the molarity of a solution that is prepared by dissolving 3.83 grams of glucose ($C_6H_{12}O_6$) in enough water to make 5.00 L of solution?

Concentration in Osmolarity

The **osmole** (osmol) unit is used to indicate *the number of moles of dissolved* ______________.

The concentration unit of measure called ______________ is commonly defined as the number of ***moles of dissolved particles (osmoles)*** per ***liter of solution***.

Osmolarity can be calculated by taking the ratio of osmoles of solute to the volume (in liters) of solution:

$$\text{Osmolarity} = \left(\frac{\text{osmoles of } \textit{solute}}{\text{liter (L) of } \textit{solution}}\right)$$

Calculating a solution's osmolarity using this equation results in units of ___________.

The **osmoles/L** unit is often referred to as ___________ and abbreviated as "**osM**."

What is the difference between a solution's osmolarity and molarity?

$$\text{Osmolarity} = \left(\frac{\boxed{\text{osmoles}} \text{ of } \textit{solute}}{\text{liter (L) of } \textit{solution}}\right) \quad \text{Molarity} = \left(\frac{\boxed{\text{moles}} \text{ of } \textit{solute}}{\text{liters (L) of } \textit{solution}}\right)$$

Sometimes it is the **same**, and *sometimes* it is **different**.

Ionic compound solutes dissociate (break apart into ions) into *individual ions* when solvated.

- For every mole of an *ionic compound* that is dissolved, the solution contains ***two or more moles*** of dissolved particles (osmoles), therefore the solution's molarity and osmolarity ***have different values***.

Molecular compound solutes do not dissociate when solvated.

- For every mole of a dissolved *molecular compound*, the solution contains one mole of dissolved particles (one osmole), therefore the solution's molarity and osmolarity ***have the same value***.

Example: Osmolarity for a Molecular Compound

If 0.50 moles of glucose (a molecular compound) is dissolved in enough water to make 2.00 L of solution, what is the ***osmolarity*** of the solution?

$$\text{Osmolarity} = \left(\frac{\text{osmoles of } \textit{solute}}{\text{liter (L) of } \textit{solution}}\right)$$

To calculate the **osmolarity**, we need to know two quantities: the number of **osmoles** of the ***solute*** **AND** the volume **(L)** of the ***solution.***

- **osmoles** of the ***solute*** (glucose): since glucose is a covalent compound, it **does not dissociate** *into ions* when dissolved. For every mole of a molecular compound that is dissolved, the solution will contain one mole of dissolved particles (one osmole).
  - 0.50 moles of glucose were dissolved, therefore the solution contains **0.50 osmoles** of glucose.
- liters (L) of ***solution*** was given: **2.00 L**

Insert the **osmoles** of the *solute* and liters (L) of *solution* into the equation for osmolarity:

$$\text{Osmolarity} = \left(\frac{0.50 \text{ osmoles}}{2.00 \text{ L}}\right) = 0.25 \text{ osmoles/L} \text{ or } 0.25 \text{ osmolar}$$

Note that in this example, since glucose is a **molecular** (non dissociation) compound, the **molarity** and the **osmolarity** have the **same** value.

Example: Osmolarity for an Ionic Compound

If 0.50 moles of sodium chloride (***an ionic compound***) are dissolved in enough water to make 2.00 L of solution, what is the ***osmolarity*** of the solution?

$$\text{Osmolarity} = \left(\frac{\text{osmoles of } \textit{solute}}{\text{liter (L) of } \textit{solution}}\right)$$

- **osmoles** of the ***solute*** (**NaCl**): Sodium chloride dissociates when dissolved, so for every ***mole*** of sodium chloride, the solution will contain *one mole* of dissolved sodium ions ***plus*** *one mole* of dissolved chloride ions.

$$\mathbf{NaCl}(s) \rightarrow \underbrace{\mathbf{Na^+}(aq)}_{\boxed{\textbf{one osmole of Na}^+}} + \underbrace{\mathbf{Cl^-}(aq)}_{\boxed{\textbf{one osmole of Cl}^-}} = \textbf{2 osmoles}$$

Multiply the number of moles of sodium chloride by a ***factor of* 2** to convert ***moles*** of sodium chloride to ***osmoles***:

$$0.50 \text{ moles } \mathbf{NaCl}\left(\frac{\mathbf{2} \text{ osmoles}}{\text{mole } \mathbf{NaCl}}\right) = \textbf{1.0 osmoles}$$

- liters (L) of ***solution*** was given: **2.00 L**

Insert the **osmoles** of the *solute* and liters (L) of *solution* into the equation for osmolarity:

$$\text{Osmolarity} = \left(\frac{1.0 \text{ osmoles}}{2.00 \text{ L}}\right) = 0.50 \text{ osmoles/L} \text{ or } 0.50 \text{ osmolar}$$

Even though we started with the *same number of moles of solute* and the *same volume of solution* in this example as we did in the previous example for a glucose solution, the osmolarity values are different ***because of the difference in the number of particles (osmoles) that are formed upon dissolution of molecular vs. ionic solutes.***

In the *Understanding Check* problem that follows, I will ask you to calculate the osmolarity of a $BaCl_2$ solution.

Keep in mind that for every ***one mole*** of $BaCl_2$ that dissolves, **3 osmoles** are formed.

$$\mathbf{BaCl_2}(s) \rightarrow \underbrace{\mathbf{Ba^{2+}}(aq)}_{\boxed{\text{one osmole of } Ba^{2+}}} + \underbrace{\mathbf{2\ Cl^-}(aq)}_{\boxed{\text{two osmoles of } Cl^-}} = \textbf{3 osmoles}$$

Understanding Check: If 0.50 moles of $BaCl_2$ (***an ionic compound***) is dissolved in enough water to make 2.00 L of solution, what is the ***osmolarity*** of the solution?

Converting between Molarity and Osmolarity

Knowing the number of ______________ (osmoles) that are formed upon dissolution of a solute will enable you to easily convert between ***molarity*** and ***osmolarity***.

The number of osmoles formed per mole of solute dissolved can be used as a ______________ _________.

Example for an Ionic Compound Solute: If the ***molarity*** of a NaCl solution is **1.2 M**, what is the ***osmolarity***?

$$\underbrace{\frac{1.2\ \cancel{\text{moles NaCl}}}{\text{L}}}_{\boxed{\text{Molarity (moles/L)}}} \underbrace{\frac{\mathbf{2}\ \text{osmoles}}{\mathbf{1}\ \cancel{\text{mole NaCl}}}}_{\boxed{\text{Conversion Factor (osmoles/mole NaCl)}}} = \underbrace{\mathbf{2.4}\left(\frac{\text{osmoles}}{\text{L}}\right)}_{\boxed{\text{Osmolarity (osmoles/L)}}}$$

Converting between molarity and osmolarity for **molecular solutes** is simple!

The ***molarity*** is _____________ to ***osmolarity*** for **molecular solutes** because they do not dissociate.

Example for a Molecular Compound Solute:

- Glucose is a molecular compound; if the ***molarity*** of a glucose solution is **1.2** M, then the ***osmolarity*** is **1.2** osmoles/L.

Understanding Check: If the molarity of an $FeCl_3$ (an ionic compound) solution is 0.010 M, what is the osmolarity? **HINT:** Think about how many osmoles are produced when one mole of $FeCl_3$ dissociates.

Concentration in Molality

____________ is defined as the number of ***moles of solute*** per ***kg of solvent***.

Molality can be calculated by taking the ratio of **moles** of solute to the **mass** (in kilograms) of the __________:

$$\text{Molality} = \left(\frac{\text{moles of } \textit{solute}}{\text{kg of } \textit{solvent}}\right)$$

Calculating a solution's molality using this equation/definition results in units of **moles/kg**.

The **moles/kg** unit is often referred to as **molal** and abbreviated as "***m***."

- The "***m***" unit *can be interchanged with* "**moles/kg**."

Calculation of Molality Example

What is the molality of a solution that is prepared by dissolving 0.125 moles of sodium chloride in 1.60 kg of water?

$$\text{Molality} = \left(\frac{\text{moles of } \textit{solute}}{\text{kg of } \textit{solvent}}\right)$$

- The **moles** of the ***solute*** (NaCl) was given: **0.125 moles**
- The **mass** of the ***solvent*** (water) was given: **1.60 kg**

Insert the moles of the *solute* and kg of *solvent* into the equation for molality:

$$\text{Molality} = \left(\frac{0.125 \text{ moles}}{1.60 \text{ kg}}\right) = 0.0781 \text{ moles/kg or } 0.0781\ m$$

In a **0.0781 molal** (***m***) solution, there are **0.0781** ***moles of solute*** contained in every ***kg of solvent***.

Understanding Check: What is the ***molality*** of a solution that is prepared by dissolving 1.34 moles of ethyl alcohol in 0.75 kg of water?

Concentration in Osmolality

Osmolality is defined as the number of ___________ per ______________ ***of solvent***.

It can be calculated taking the ratio of ***osmoles*** (moles of dissolved particles) to the mass (in kilograms) of the solvent:

$$\text{Osmolality} = \left(\frac{\text{osmoles of } \textit{solute}}{\text{kg of } \textit{solvent}}\right)$$

Calculating a solution's osmolality using this equation/definition results in units of **osmoles/kg**.

- The os**moles/kg** unit is often referred to as **osmolal**

Concentration in Equivalents per Liter (Eq/L)

Some properties of solutions depend on the ***total charge*** of the _______ in solution.

- For example, the ability of a solution to conduct electricity depends on the total charge of the ions in solution. In cells, membrane potentials that generated nerve signals depend on the total charge of ions in solution.

An **equivalent** (**Eq**) is defined as a ***mole of*** __________.

The concentration unit of measure called **equivalents per liter (Eq/L)** is defined as the number of **equivalents *(Eq)*** of solute (*moles of charge*) per ***liter of solution***:

$$\text{Eq/L} = \left(\frac{\text{Eq of } \textit{solute}}{\text{liter (L) of } \textit{solution}} \right)$$

At first glance, an **equivalent** (**Eq**) *may appear to be the same* as an ***osmole***, however this is not always the case.

An **equivalent** is a mole of ***charge***; an **osmole** is a mole of ***dissolved particles***.

I will elaborate by comparing the number of osmoles vs. the number of equivalents present when one mole of $BaCl_2$ is dissolved.

When **one mole** of $BaCl_2$ is dissolved, ***3 osmoles*** are formed.

$$BaCl_2(s) \rightarrow \underbrace{Ba^{2+}(aq)}_{\boxed{\text{One Osmole of } Ba^{2+}}} + \underbrace{2\,Cl^{-}(aq)}_{\boxed{\text{Two Osmoles of } Cl^{-}}} = 3 \text{ Osmoles}$$

When **one mole** of $BaCl_2$ is dissolved, **4 equivalents (Eq)** are formed.

$$BaCl_2(s) \rightarrow \underbrace{Ba^{2+}(aq)}_{\boxed{\text{Two Eq of } Ba^{2+}}} + \underbrace{2\,Cl^{-}(aq)}_{\boxed{\text{Two Eq of } Cl^{-}}} = 4 \text{ Eq}$$

- Since barium ions have a "**2+**" **charge**, one mole of barium ions contains *two moles of charge* (***two*** equivalents).
- The **two** moles of chloride ions contain a total of ***two*** equivalents (*two* moles of a "**1-**" charge).

Understanding Check

When **one mole** of $Fe_2(SO_4)_3$ is dissolved in water:

a. How many **equivalents** are present?

b. How many equivalents of **sulfate** are present (equivalents from sulfate only)?

c. How many equivalents of **iron(III)** are present (equivalents from Fe^{3+} only)?

Example for Calculating Eq/L

If 0.50 moles of $BaCl_2$ is dissolved in enough water to make 2.00 L of solution, what is the **Eq/L** ***concentration*** of the solution?

$$\text{Eq/L} = \left(\frac{\text{Eq of } \textit{solute}}{\text{liter (L) of } \textit{solution}} \right)$$

Eq present:

- For every ***one mole*** of $BaCl_2$ dissolved, **4 Eq** are formed (two **Eq** of $\mathbf{Ba^{2+}}$ and two **Eq** of $\mathbf{Cl^-}$).

$$\mathbf{BaCl_2}(s) \rightarrow \underbrace{\mathbf{Ba^{2+}}\,(aq)}_{\boxed{\begin{array}{c}\textbf{Two Eq}\\ \textbf{of } \mathbf{Ba^{2+}}\end{array}}} + \underbrace{\mathbf{2\ Cl^-}\,(aq)}_{\boxed{\begin{array}{c}\textbf{Two Eq}\\ \textbf{of } \mathbf{Cl^-}\end{array}}} \mathbf{= 4\ Eq}$$

- Multiply the number of moles of $BaCl_2$ by a factor of **4** to convert ***moles*** of $BaCl_2$ to **Eq**:

$$0.50 \text{ moles } BaCl_2 \left(\frac{\mathbf{4\ Eq}}{\text{mole } BaCl_2} \right) = \mathbf{2.0\ Eq}$$

liters (L) of ***solution*** was given: **2.00 L**

Insert the **Eq** present and liters (L) of *solution* into the equation for **Eq/L** concentration:

$$\text{Eq/L} = \left(\frac{2.0 \text{ Eq}}{2.00 \text{ L}} \right) = 1.0 \text{ Eq/L}$$

Understanding Check

If 0.015 moles of $Fe_2(SO_4)_3$ is dissolved in enough water to make 2.5 L of solution, what is the **Eq/L** ***concentration*** of the solution?

Calculations for Solutions

In this section, you will learn how do calculations to find how much solute is contained in a specified amount of solution *and* how much solution contains a specified amount of solute.

The key to mastering these calculations is to be aware of the following two statements:

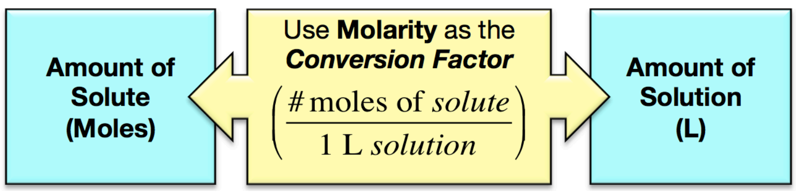
- The __________________ between the *amount of solute* and the *amount of solution* is the ***concentration***.

- You can convert between the *amount of* ____________ and the *amount of* ____________ by ***using the*** __________________ ***as a*** __________________ ____________**.**

Molarity Concentration Calculations for Solutions

The ***molarity*** (____________) of a solution gives us the *relationship* between the ***amount (moles) of solute*** and the ***volume (L) of solution***.

We use the ***molarity*** as a ***conversion factor*** when converting between the moles of solute in a given volume (L) of solution, ***or*** the volume (L) of solution that will contain a given amount (moles) of solute.



Volume of Solution to Amount of Solute Example

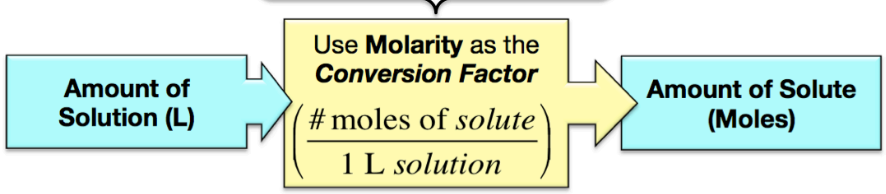
Suppose you know that a patient received 3.50 **L** of a 0.278 **M** glucose IV solution, how many moles of glucose were administered to the patient?

STEP 1) Set up the equation using the given quantity:

$$\frac{3.50\ \text{L solution}}{}\ \Bigg|\ \ \Bigg| = \ ?\ \ \text{moles glucose}$$

STEP 2) Use the ***molarity*** as a **conversion factor** to find the ***number of moles***:

$$\frac{3.50\ \cancel{\text{L solution}}}{}\ \Bigg|\ \underbrace{\frac{0.278\ \text{moles glucose}}{1\ \cancel{\text{L solution}}}}\ \Bigg| = 0.973\ \ \text{moles glucose}$$



3.50 L of a **0.278 M** glucose solution will contain **0.973** moles of glucose.

Moles of Solute to Volume (L) of Solution Example

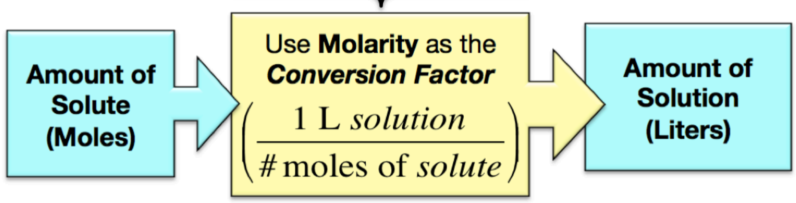
Suppose you wished to administer 0.200 moles of glucose from a 0.278 M glucose IV solution to a patient, what ***volume*** (in liters) of the solution would need to be dispensed?

STEP 1) Set up the equation using the given quantity:

$$\begin{array}{c|c|} 0.200 \text{ moles glucose} & \\ \hline & \end{array} = \ ?\ \text{ L solution}$$

STEP 2) Use the ***molarity*** as a **conversion factor** to find the ***volume (L) of solution:***

$$\begin{array}{c|c|} 0.200\ \cancel{\text{moles glucose}} & 1 \text{ L solution} \\ \hline & 0.278\ \cancel{\text{moles glucose}} \end{array} = 0.719 \text{ L solution}$$



0.719 L of a **0.278 M** glucose solution would be given to the patient in order to provide 0.200 moles of glucose.

Another Molarity Concentration Calculation Example

Next, I want to show you an example problem that combines the concepts of the ***previous two videos*** (calculating a solution's concentration) *and* the concepts of this section (calculations of how much solute is contained in a certain amount of solution *or* how much solution contains a certain amount of solute).

If 1.25 g of acetone (C_3H_6O) is dissolved in enough water to make 0.550 L of solution;

- **a)** What is the **molarity (M)** of the solution?
- **b)** How many ***moles*** of acetone are contained in 0.0679 L of this acetone ***solution***?
- **c)** What **volume (L)** of this acetone solution would contain 0.0079 ***moles*** of acetone?

a) What is the **molarity (M)** of the solution?

$$\text{Molarity} = \left(\frac{\text{moles of } solute}{\text{liters (L) of } solution}\right)$$

- **moles** of the ***solute*** (acetone)

  We were not given the number of moles ***directly***, however, we were given the *grams* of acetone (1.25 g). Use the *molar mass* of acetone (**58.09 g/mole**) to convert from grams to moles.

$$\begin{array}{c|c|} 1.25\ \cancel{\text{grams } C_3H_6O} & 1 \text{ mole } C_3H_6O \\ \hline & 58.09\ \cancel{\text{grams } C_3H_6O} \end{array} = 0.0215 \text{ moles } C_3H_6O$$

- **L** of ***solution*** was given: **0.550 L**

Insert the moles of the *solute* and liters (L) of *solution* into the equation for molarity:

$$\text{Molarity} = \left(\frac{0.0215 \text{ moles}}{0.550 \text{ L}}\right) = 0.0391 \text{ moles/L} \ \text{ or } \ 0.0391 \text{ M}$$

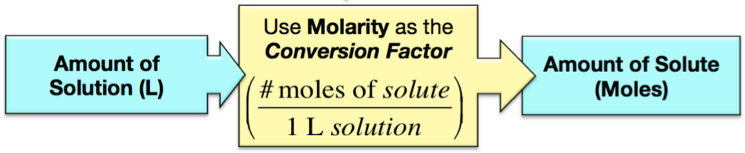
b) How many ***moles*** of acetone are contained in 0.0679 L of this acetone ***solution***?

STEP 1) Set up the equation using the given quantity:

$$\frac{0.0679 \text{ L solution}}{} \Bigg| \frac{}{} \Bigg| =$$

STEP 2) Use the ***molarity*** as a **conversion factor** to find the ***number of moles***:

$$\frac{0.0679 \text{ }\cancel{\text{L solution}}}{} \Bigg| \underbrace{\frac{0.0391 \text{ moles acetone}}{1 \text{ }\cancel{\text{L solution}}}} \Bigg| = 0.00265 \text{ moles acetone}$$



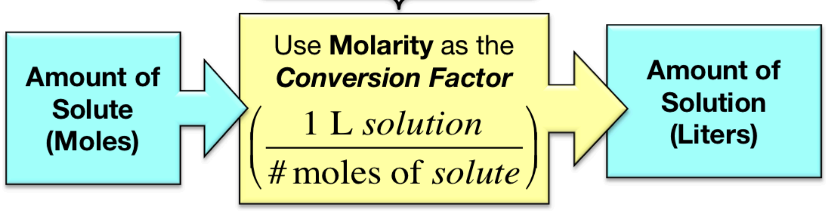
c) What **volume (L)** of this acetone solution would contain 0.0079 ***moles*** of acetone?

STEP 1) Set up the equation using the given quantity:

$$\frac{0.0079 \text{ moles acetone}}{} \Bigg| \frac{}{} \Bigg| =$$

STEP 2) Use the ***molarity*** as a **conversion factor** to find the ***volume (L) of solution:***

$$\frac{0.0079 \text{ }\cancel{\text{moles acetone}}}{} \Bigg| \underbrace{\frac{1 \text{ L solution}}{0.0391 \text{ }\cancel{\text{moles acetone}}}} \Bigg| = 0.20 \text{ L solution}$$



Understanding Check: If a particular wine has an ethyl alcohol molarity concentration of 2.8 M, what volume (in liters) of wine contains 10.4 moles of ethyl alcohol (the lethal dosage)?

Understanding Check: How many ***grams*** of silver nitrate ($AgNO_3$) are contained in 0.384 L of a 0.200 M silver nitrate solution?

Percent (%) Concentration Calculations for Solutions

Percent (%) concentration gives the relationship between the *amount of solute* and the *amount of solution*:

Amount of Solute (Grams) ⟷ Use **%(w/w)** as the ***Conversion Factor*** $\left(\dfrac{\#\text{ grams of } \textit{solute}}{100 \text{ grams of } \textit{solution}}\right)$ ⟷ **Amount of Solution (Grams)**

Amount of Solute (mL) ⟷ Use **%(v/v)** as the ***Conversion Factor*** $\left(\dfrac{\#\text{ mL of } \textit{solute}}{100 \text{ mL of } \textit{solution}}\right)$ ⟷ **Amount of Solution (mL)**

Amount of Solute (Grams) ⟷ Use **%(w/v)** as the ***Conversion Factor*** $\left(\dfrac{\#\text{ grams of } \textit{solute}}{100 \text{ ml of } \textit{solution}}\right)$ ⟷ **Amount of Solution (mL)**

Example: Using % (w/v) to Convert From Volume (mL) of Solution to Grams of Solute

Normal saline intravenous (IV) drips are composed of sterile, 0.90 %(w/v) aqueous sodium chloride (NaCl) solutions. They are used to treat or prevent dehydration and hypovolemia.

If a patient received 2475 mL of a normal saline solution, how many grams of sodium chloride were delivered?

$$2475\ \cancel{\text{mL solution}} \times \underbrace{\frac{0.90 \text{ grams NaCl}}{100\ \cancel{\text{mL solution}}}} = 22 \text{ grams NaCl}$$

Amount of Solution (mL) → Use **%(w/v)** as the ***Conversion Factor*** $\left(\dfrac{\#\text{ grams of } \textit{solute}}{100 \text{ ml of } \textit{solution}}\right)$ → **Amount of Solute (grams)**

If a patient received 2475 mL of a normal saline solution, then **22 grams** of sodium chloride were delivered.

Example: Using %(w/v) to Convert From Grams of Solute to Volume (mL) of Solution

What volume (mL) of a normal saline solution (0.90% (w/v)) contains 12.5 grams of sodium chloride?

$$12.5\ \cancel{\text{grams NaCl}} \times \underbrace{\frac{100 \text{ mL solution}}{0.90\ \cancel{\text{grams NaCl}}}} = 1400 \text{ mL solution}$$

Amount of Solute (grams) → Use **%(w/v)** as the ***Conversion Factor*** $\left(\dfrac{100 \text{ ml of } \textit{solution}}{\#\text{ grams of } \textit{solute}}\right)$ → **Amount of Solution (mL)**

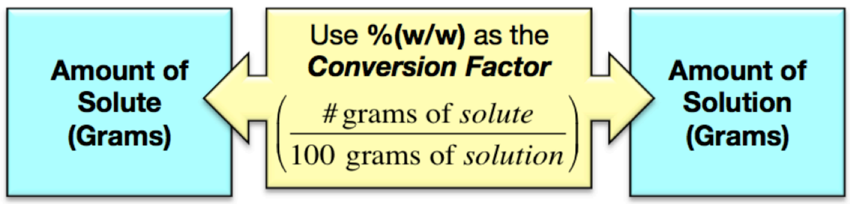
1400 mL of a normal saline solution contain 12.5 grams of sodium chloride.

Understanding Check

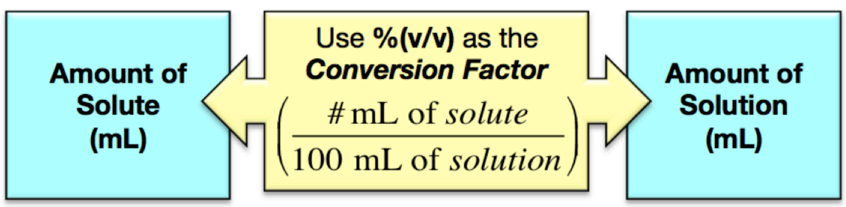
The label of the medication vial tells you that the concentration of morphine sulfate for an intravenous injection is 1.0% (w/v). If you wish to administer 0.0025 grams of morphine sulfate, what volume (mL) would you inject?

The method for converting between the ***amount of solute*** and the ***amount of solution*** can also be used for **%(w/w)** and **%(v/v)**.

%(w/w) is used to convert between the ***mass*** of solute *and* the **mass** of solution:

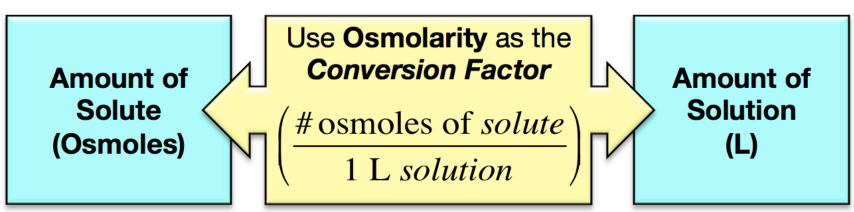


%(v/v) is be used to convert between the **volume** of a liquid solute and the **volume** of the solution:



Osmolarity Concentration Calculations for Solutions

Osmolarity concentration is a relationship between the amount of solute and the amount of solution:



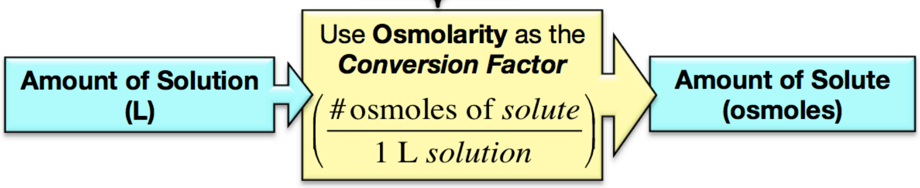
The **osmolarity (osmoles/L)** of a solution gives us the number of ***osmoles of solute*** contained in ***1 L of solution***.

- It can therefore be used to **convert** between _____________ *of solute* and _________ *of solution*.

Example: Using **Osmolarity** to convert between ***L of Solution*** and ***Osmoles of Solute***

How many ***osmoles of solute*** are contained in 2.75 L of a solution that has a concentration of 0.0750 ***osmole/L***?

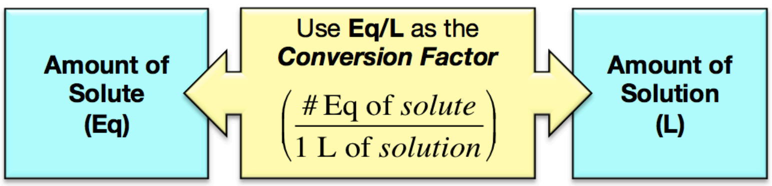
$$2.75\ \cancel{\text{L solution}} \times \frac{0.0750\ \text{osmoles}}{1\ \cancel{\text{L solution}}} = 0.206\ \text{osmoles}$$



0.206 *osmoles* are contained in 2.75 L of a 0.0750 *osmole/L solution*.

Equivalents per Liter (Eq/L) Concentration Calculations for Solutions

Eq/L concentration is a relationship between the amount of solute and the amount of solution:



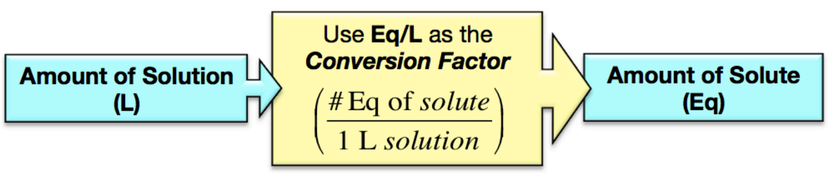
The **Eq/L** of a solution gives us the number of ***equivalents of solute*** contained in ***1 L of solution***.

- It can therefore be used to convert between _____________ ***of solute*** and ***L of solution***.

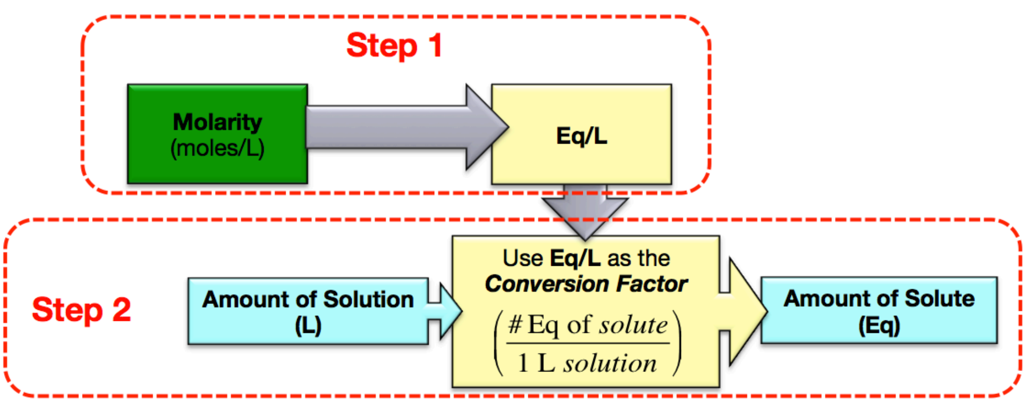
Example: Using **Eq/L** to convert between ***L of Solution*** and ***Equivalents of Solute***

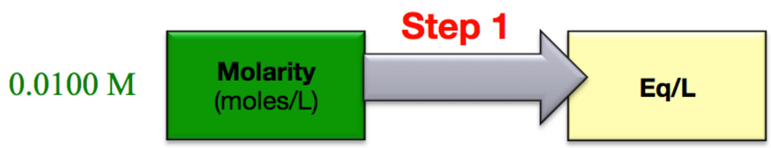
How many equivalents of solute are contained in 0.830 L of a 0.0100 M $Fe_2(SO_4)_3$ solution?

Solution:



We are given the solution concentration in molarity (M = mole/L), but we need to get (Eq/L) in order to solve the problem. First (Step 1) we will convert molarity (mole/L) to (Eq/L), and then (Step 2) we will use the Eq/L concentration as a conversion factor to convert from liters of solution to equivalents of solute.



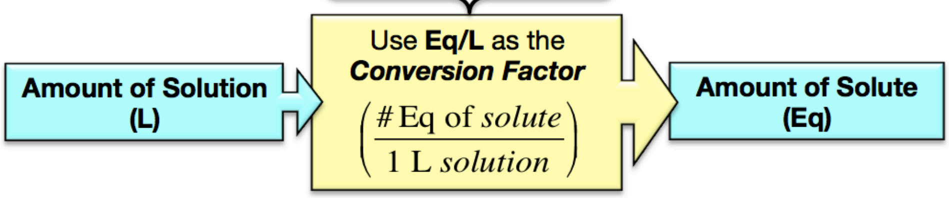


$$Fe_2(SO_4)_3\,(s) \rightarrow \underbrace{\underbrace{2\,Fe^{3+}\,(aq)}_{\boxed{6\text{ Eq of }Fe^{3+}}} + \underbrace{3\,SO_4^{2-}\,(aq)}_{\boxed{6\text{ Eq of }SO_4^{2-}}}}_{\boxed{12\text{ total Eq}}}$$

$$\underbrace{\frac{0.0100\ \cancel{\text{moles }Fe_2(SO_4)_3}}{\text{L}}}_{\boxed{\text{Molarity (moles/L)}}}\left|\frac{12\text{ Eq}}{1\ \cancel{\text{mole }Fe_2(SO_4)_3}}\right| = 0.120\underbrace{\left(\frac{\text{Eq}}{\text{L}}\right)}_{\boxed{\text{Equivalents per Liter}}}$$

Step 2:

$$0.830\ \cancel{\text{L solution}}\left|\frac{0.120\text{ Eq}}{1\ \cancel{\text{L solution}}}\right| = 0.0996\text{ Eq}$$

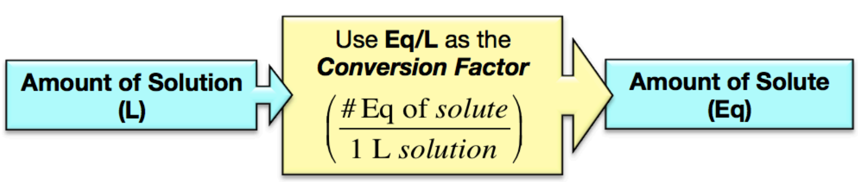


0.0996 Eq of solute are contained in **0.830 L** of a **0.0100 M** $Fe_2(SO_4)_3$ solution.

Example: Using **Eq/L** to convert between ***L of Solution*** and ***Equivalents of Solute***

How many equivalents of **sulfate (*not total equivalents*)** are contained in 0.830 L of a 0.0100 M $Fe_2(SO_4)_3$ solution?

Solution:



We will do this problem in the same way as we did for the previous example problem, ***with one exception***: we will convert the given **molarity** (0.0100 **M**) to **Eq of SO_4^{2-}/L**.

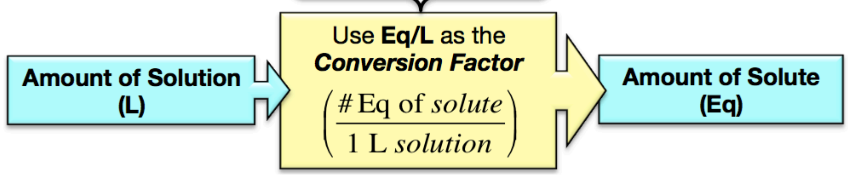
0.0100 M

$$Fe_2(SO_4)_3\,(s) \rightarrow 2\,Fe^{3+}\,(aq) + \underbrace{3\,SO_4^{2-}\,(aq)}_{\boxed{6\text{ Eq of }SO_4^{2-}}}$$

$$\underbrace{\frac{0.0100\ \cancel{\text{moles }Fe_2(SO_4)_3}}{\text{L}}}_{\boxed{\text{Molarity (moles/L)}}} \times \frac{6\text{ Eq }SO_4^{2-}}{1\ \cancel{\text{mole }Fe_2(SO_4)_3}} = 0.0600\left(\frac{\text{Eq }SO_4^{2-}}{\text{L}}\right)$$

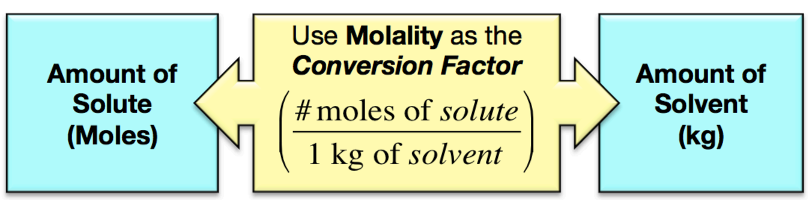
Step 2:

$$0.830\ \cancel{\text{L solution}} \times \underbrace{\frac{0.0600\text{ Eq }SO_4^{2-}}{1\ \cancel{\text{L solution}}}} = 0.0498\text{ Eq }SO_4^{2-}$$

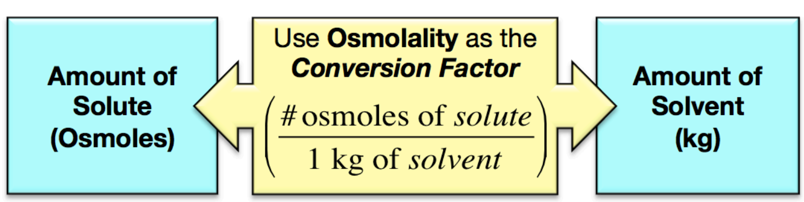


Molality and Osmolality Concentration Calculations for Solutions

Molality is used to covert between ***moles of solute*** and ***kg of*** _______________:



Osmolality is used to covert between ***osmoles of solute*** and ***kg of solvent***:



In practice, it is more useful to know how much solute is contained in a particular **amount of solution *(not solvent)*** or how much **solution (*not solvent*)** contains a particular amount of solute.

In order to work with the *amount of solution* **instead** *of the amount of solvent*, a very useful approximation can be made ***for dilute aqueous solutions***.

In the case of dilute aqueous solutions, *the solution is almost entirely solvent*.

Since 1 kg of water has a volume of 1 L, it is a reasonable approximation to equate ***the amount of solution*** to ***the amount of solvent;*** 1 kg of solvent is ***assumed*** to be the same as 1 L of solution.

Using this approximation, the ***molality*** is _____________ to the ***molarity***.

For *Dilute Aqueous Solutions*:
Molality Units Can Be Replaced by Molarity Units:

$$\underbrace{\left(\frac{\#\text{ moles of } solute}{1\text{ kg of } solvent}\right)}_{\textbf{molality}} \Rightarrow \underbrace{\left(\frac{\#\text{ moles of } solute}{1\text{ L } solution}\right)}_{\textbf{molarity}}$$

For the same reason, using this approximation, the ***osmolality*** is *equal* to the ***osmolarity***.

For *Dilute Aqueous Solutions*:
Osmolality Units Can Be Replaced by Osmolarity Units:

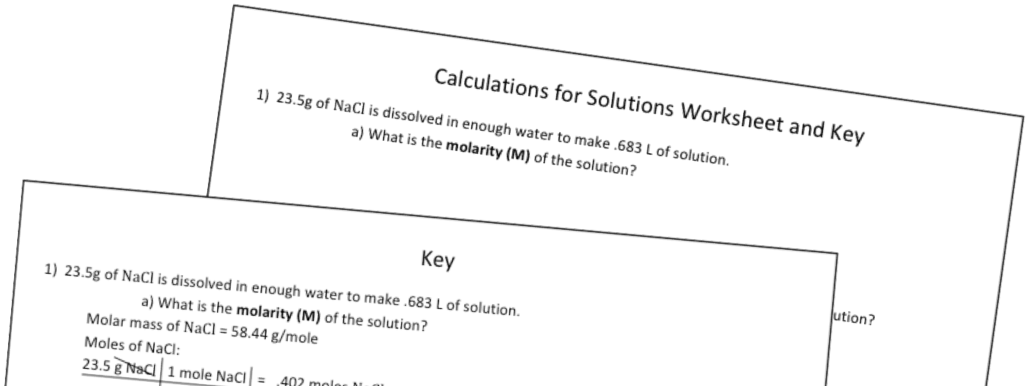
$$\underbrace{\left(\frac{\#\text{ osmoles of } solute}{1\text{ kg of } solvent}\right)}_{\textbf{osmolality}} \Rightarrow \underbrace{\left(\frac{\#\text{ osmoles of } solute}{1\text{ L } solution}\right)}_{\textbf{osmolarity}}$$

Summary of Conversion Factors for Solution Calculations

Amount of Solute	When converting Between / Use One of the Following Concentrations as the Conversion Factor:	Amount of Solution
moles of solute	**molarity** (moles/L)	liters (L) of solution
osmoles of solute	**osmolarity** (osmoles/L)	liters (L) of solution
equivalents (Eq) of solute	**equivalents/L** (Eq/L)	liters (L) of solution
mass of solute (typically grams)	**% (w/w)** typically (g solute/100 g solution)	mass of solution (typically grams)
volume of solute (typically mL)	**% (v/v)** typically (mL solute/100 mL solution)	volume of solution (typically mL)
grams of solute	**% (w/v)** (g solute/100 mL solution)	mL of solution

Understanding Check

Before Watching the Next Video: Do the Problems in the **Calculations for Solutions Worksheet**



Calculations for Solutions Worksheet and Key

1) 23.5g of NaCl is dissolved in enough water to make .683 L of solution.
a) What is the **molarity (M)** of the solution?

Key

1) 23.5g of NaCl is dissolved in enough water to make .683 L of solution.
a) What is the **molarity (M)** of the solution?
Molar mass of NaCl = 58.44 g/mole
Moles of NaCl:
23.5 g NaCl | 1 mole NaCl | = .402 moles NaCl

The Solubility of Biological Compounds

Biological compounds are the __________ _____________ that occur in biological organisms.

Examples of biological compounds that you will learn about in this book are: steroids, fatty acids, bile salts, phospholipids, glycolipids, cholesterol, triglycerides (animal fat and vegetable oil), proteins, carbohydrates, RNA, and DNA.

Biological compounds can be put into one of three categories based on their __________ ____________.

1) **Hydrophilic**
2) **Hydrophobic**
3) **Amphipathic**

1) Hydrophilic Compounds

Hydrophilic compounds ____________ in water.

- The word *hydrophilic* is derived from an ancient Greek word that is translated as "*loving water*."

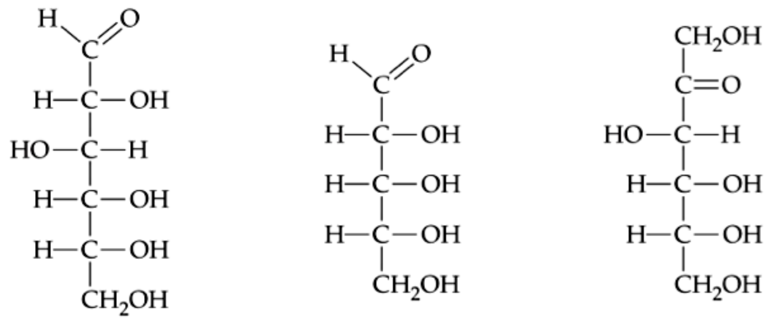
Compounds that are significantly ________ and/or can ____________ ________ with water tend to be water soluble.

As a general rule, molecules that have at least ______ ***polar functional group*** for every _______ ***carbon atoms*** are water soluble, and therefore classified as **hydrophilic**.

- You saw *four* ***polar functional groups*** in chapter 4: the hydroxyl group (-OH), the carbonyl group (C=O), the carboxyl group (-COOH), and the carboxylate group (COO).
- There are a few other polar functional groups that you will see in later chapters.
- The presence of ____________ __________will also help a biological compound to dissolve in water because of the attraction of water molecules' dipoles to the charged region of the compound (ion-dipole interactions).

Examples of Hydrophilic Compounds: Monosaccharides

The Structural Formulas of Three Monosaccharides



D-Glucose (makes up starch and cellulose; source of energy)

D-Ribose (occurs in ATP, RNA, and coenzymes)

D-Fructose (sweeter than table sugar, a component of corn syrup, occurs in fruit)

2) Hydrophobic Compounds

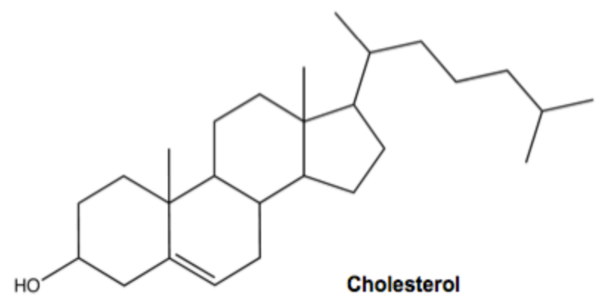
Hydrophobic compounds ____ _____ ***dissolve*** in water.

- The word *hydrophobic* is derived from an ancient Greek word that is translated as "*having a horror/fear of water*."

As a general rule, molecules that have less than **one *polar functional group*** for every ***five carbon atoms*** do not dissolve in water and are therefore *hydrophobic*.

An Example of a Hydrophobic Compound: Cholesterol

Note that cholesterol does have *one* polar hydroxyl (-OH) functional group, however the nonpolar part of the molecule is so large that the ratio of polar functional group to total carbons is ***much less*** than 1:5 and therefore cholesterol does not dissolve in water.



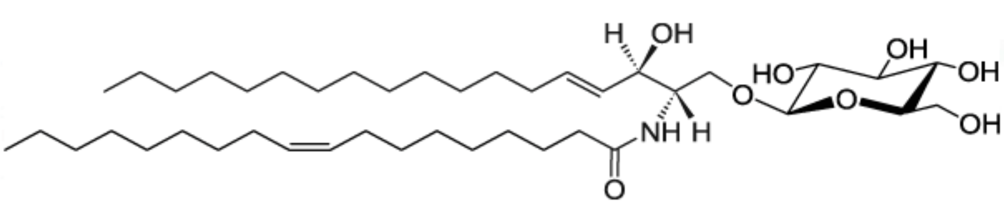
Cholesterol

3) Amphipathic Compounds

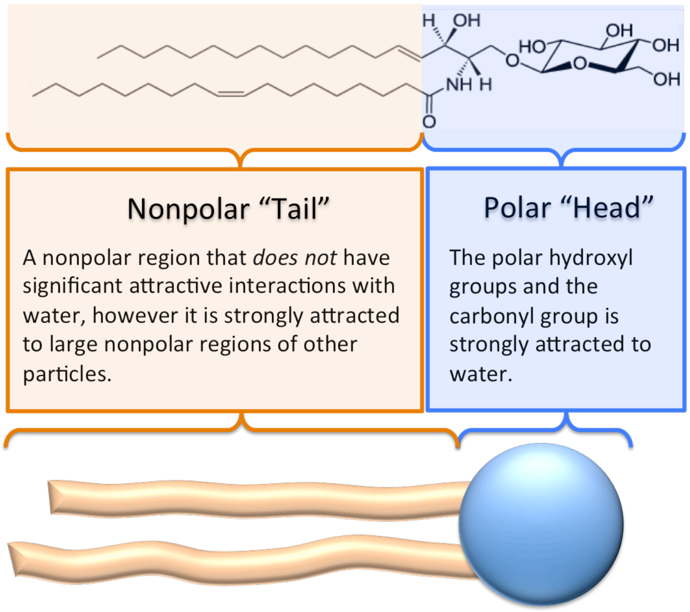
Amphipathic compounds have _________ a large ***nonpolar*** region, which is *not* strongly attracted to water, ***and*** an ***extremely polar*** and/or ***formally-charged*** region, which is quite strongly attracted to water.

An Example of an Amphipathic Compound: A Glycolipid

The particular glycolipid shown in the structure below is one of the most prevalent of the glycolipids that make up cell membranes within the brain.



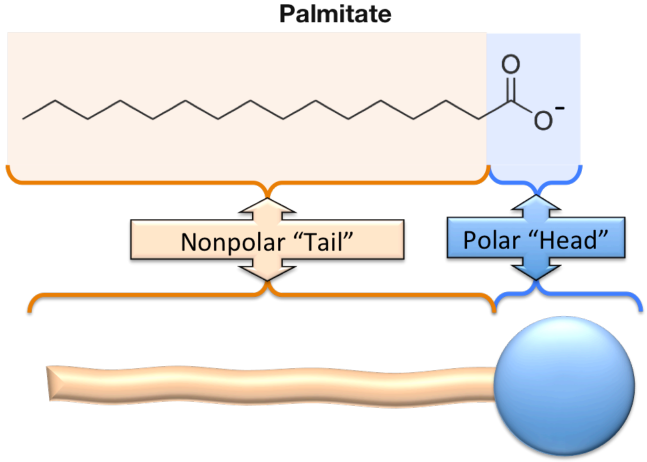
Amphipathic compounds are often illustrated using a ___________ for the ***polar head*** that is attached to ***one or more long tubular structures*** that represent the carbon chains in the ***nonpolar tail***.



In some amphipathic compounds, such as the glycolipid shown above, there are *two* carbon chains that make up the nonpolar tail; in other amphipathic compounds, the tail is composed of *only one* carbon chain.

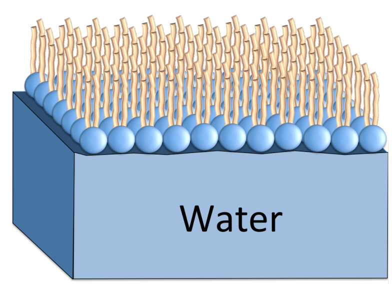
An example of an amphipathic compound that has a *single* carbon chain tail is *palmitate*.

Palmitate

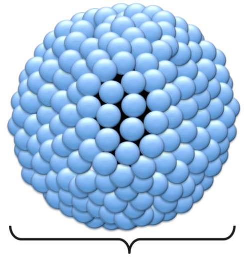


When amphipathic molecules are put into water they do not dissolve; they exist as **monolayers** and/or **micelles.**

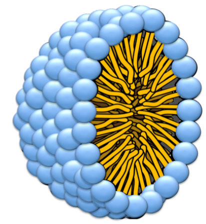
- As _____________, amphipathic compounds form a single (mono) layer of individual particles oriented with their polar heads toward the water and their nonpolar tails pointing upward.



Amphipathic compounds can also exist in water as _____________ arrangements called ______________.



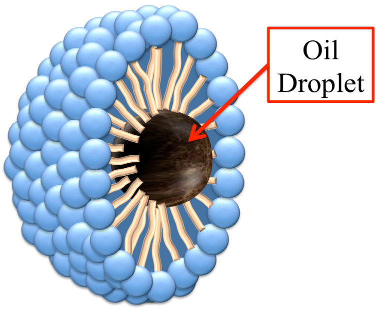
a micelle



cross-sectional view of a micelle

The amphipathic compounds making up micelles are oriented with their ***polar heads*** outward, toward the water, and their ***nonpolar tails*** inward, away from the water.

Soap consists of palmitate and/or similar *amphipathic compounds.*



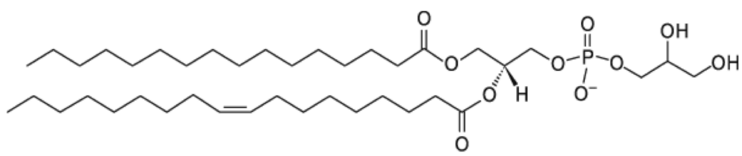
As you know, soap *__and__* water are much more effective at removing oil from skin than is just water alone. This is because there is no strong attraction between the nonpolar oil molecules and water; however, soap forms micelles that *encapsulate* the oil within their nonpolar tail interiors. Micelles containing the oil can move into the rinse water and away from the skin.

When a liquid contains compounds that are *encapsulated* by amphipathic compounds in micelles, the mixture is called an ___________.

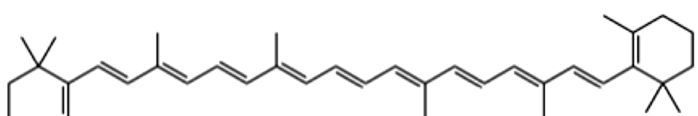
Understanding Check

Predict whether each of the following biological compounds is **hydrophobic** or **amphipathic**?

a.



b.

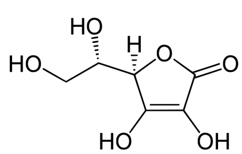


beta-carotene (a molecule in the vitamin A group)

Understanding Check

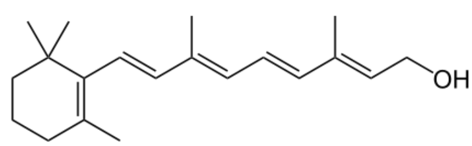
Predict whether each of the following biological compounds is **hydrophilic** or **hydrophobic**?

a.



vitamin C

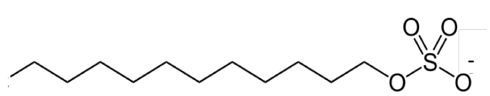
b.



retinol (a molecule in the vitamin A group)

Understanding Check

The ion shown below is called *laurel sulfate*. Laurel sulfate is *amphipathic* and is often used in shampoo.Is the ***polar head*** located on the ***right-hand <u>or</u> left-hand*** end of the molecule as it is illustrated below?



laurel sulfate

Dilutions

Dilution is the process of **adding** more ***solvent*** to a solution.

A series of dilutions (left to right) of an aqueous solution containing a colored solute.

When considering ***dilutions***, the **concentration** of the solution is ***inversely proportional*** to the volume of the solution.

- For example, if enough solvent is added to ***double*** the volume, then the **concentration** is ***decreased by a factor of 1/2***.

As you saw in chapter 5 with Boyle's gas law, when properties are *inversely proportional*, the **product** of the *initial* and *final* properties are **equal**.

In the case of dilution, the **product** of the initial (un-diluted) concentration (C_1) and the volume before dilution (V_1) ***is equal to*** the **product** of the final (diluted) concentration (C_2) and the final volume (V_2):

$$C_1 \bullet V_1 = C_2 \bullet V_2$$

This equation is called the "__________ ____________" and it can be used with any of the concentration units of measure that include the *volume of solution* (molarity, osmolarity, % (w/v), % (v/v), or Eq/L).

- For example, the dilution equation can be written using **molarity (M)** concentration:

$$M_1 \bullet V_1 = M_2 \bullet V_2$$

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?

$$\boxed{\textbf{Strategy: } M_1 \bullet V_1 = M_2 \bullet V_2}$$

$$\frac{M_1\,V_1}{V_2} = \frac{M_2\,\cancel{V_2}}{\cancel{V_2}}$$

$M_1 = 1.8$ M	$M_2 = ?$
$V_1 = 25$ mL	$V_2 = 35$ mL

$$M_2 = \frac{M_1\,V_1}{V_2} = \frac{(1.8\text{ M})\ \ (25\ \cancel{\text{mL}})}{(35\ \cancel{\text{mL}})} = \mathbf{1.3\ M}$$

Understanding Check

If 1.70 L of a 1.50 M solution is diluted to a final volume of 3.50 L, what is the final concentration?

Colloids and Suspensions

When particles that are larger than typical molecules or ions are put into another medium, typically water, the resulting mixture is classified as either a **colloid** or a **suspension** depending on the _______________ ________.

In ______________, the dispersed particles (***colloidal particles***) are small enough that they ***do not*** *settle to the bottom of their container*.

Conversely, in _________________, the solid particles are large enough that gravity causes them to *settle to the bottom of their container* unless the mixture is repeatedly or constantly stirred or shaken.

Colloids

Colloidal particles are typically in the size range of ***1 nanometer up to 1 micrometer***.

Because of their relatively small size, the kinetic energy from collisions with the particles making up the medium, typically water molecules, overcomes the force of gravity and the particles remain evenly dispersed in the medium.

If the particles are very small, the colloid will not "scatter" light and it will therefore appear clear or colored (but not cloudy). As the particle size gets larger, a colloid mixture will appear cloudy since the light entering the medium is scattered in many different directions by particles.

The ***micelles*** that you learned about in this chapter are examples of colloids.

Another example of a colloid is milk.

- Milk contains small agglomerations of many individual protein molecules (these particles are called "casein") as well as particles composed of emulsified fat (triglyceride) molecules. It is these fat particles that are separated from milk to make butter. Milk is classified as a *colloid* rather than a *suspension* because the colloidal particles do not settle to the bottom. Note that milk containers do not say "*shake well before using*" because the particles are small enough to remain evenly dispersed.

Other examples of colloids are mayonnaise and hand lotion.

Suspensions

The solid particles contained in **suspensions** are typically larger than 1 micrometer.

Because of their relatively large size compared to colloidal particles, the force of gravity causes the particles to settle to the bottom of the container. The settling process may take seconds or several hours. When stirred or shaken, the suspension will appear cloudy since the particles are large enough to scatter light.

An example of a suspension is muddy water.

- If muddy water is constantly stirred, the clay/dirt particles are evenly distributed throughout the container; however, if the stirring is discontinued, the particles will settle to the bottom of the container.

Another example of a suspension is orange juice.

If a liquid contains solid particles and is labeled "________ _____ before using," then ***it is a suspension***.

Understanding Check
If sand is added to a glass of water, is the resulting mixture a colloid or suspension?

Diffusion and Osmosis

Diffusion

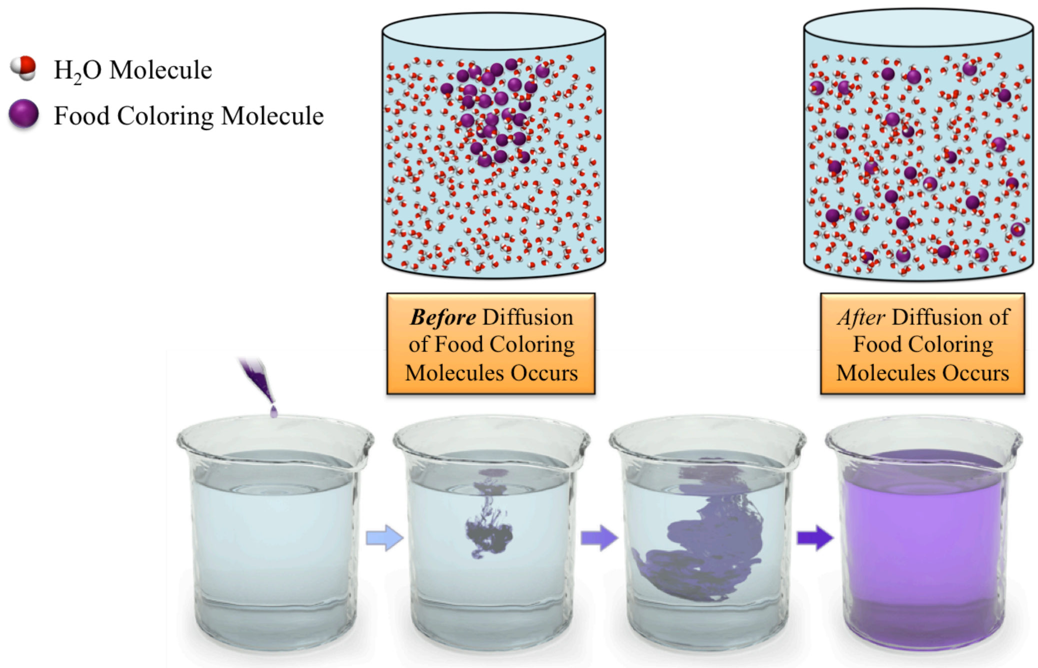
Just like gases, solute and solvent particles in liquid phase solutions travel in random directions until they collide with other particles or the container wall.

This type of random movement is called ***Brownian motion*** and results in a process called __________.

Diffusion is defined as the net transport of a substance, due to *Brownian motion,* from a region of __________ concentration of the substance to a region of __________ concentration of the substance.

It is the random movement of particles that causes them to be evenly mixed.

In the diffusion process, substances will ***spontaneously*** move from an area of ***greater concentration*** (of the particular substance) **to** ***lesser concentration*** until it is evenly distributed.



When a substance is not evenly distributed and has a ***greater concentration*** in one region and a ***lesser concentration*** in another region, we say that there is a "______________ ______________."

__________ a ***concentration gradient*** is present, and there is not a physical barrier preventing transport, diffusion will occur.

- We say that the diffusing species move "________ ***the concentration gradient***."

There can only be a net movement of dissolved particles from areas of ***lesser concentration*** to areas of ***greater concentration*** (***against*** the concentration gradient) ***when external*** __________ ***is supplied***.

Osmosis

A _______________ _____________ is any type of physical barrier through which only certain substances can pass.

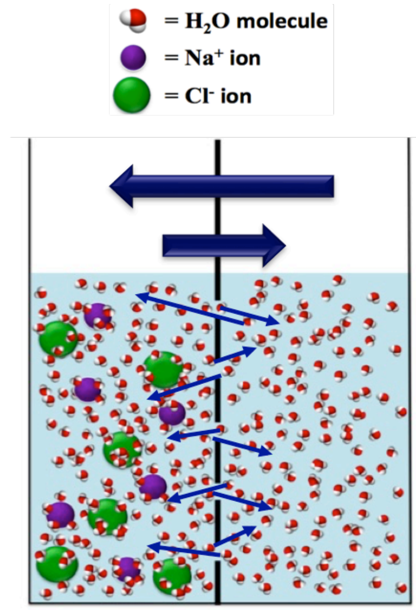
- For example, many membranes, both natural and synthetic, are ***permeable*** to water (allow water to pass) but are ***impermeable*** to ions (do not allow ions to pass).

As a general rule, *biological membranes* in cells, are ***permeable*** to nonpolar molecules and small polar molecules, and are ***impermeable*** to ions and large polar molecules.

Water molecules pass through holes (pores) in biological membranes called ***aquaporins***.

Osmosis *is the net transport of* ___________ from *a* solution with a ***lesser solute*** particle concentration *through a semipermeable membrane t*o a solution with a ***greater solute*** particle concentration.

Note that a semipermeable membrane allows ***solvent*** to *continuously* move *back and forth* ***between both sides of a membrane***; however, in ***osmosis***, there is a greater amount of solvent transported in the direction from the side of the membrane with ***lesser solute*** particle concentration to the side of the membrane with ***greater solute*** particle concentration, resulting in a ***net (overall) transport of solvent*** in that direction.

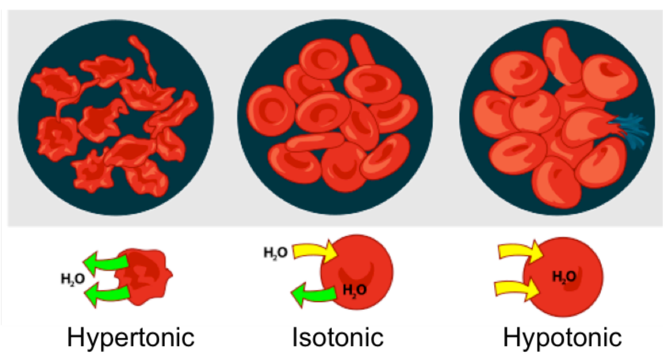


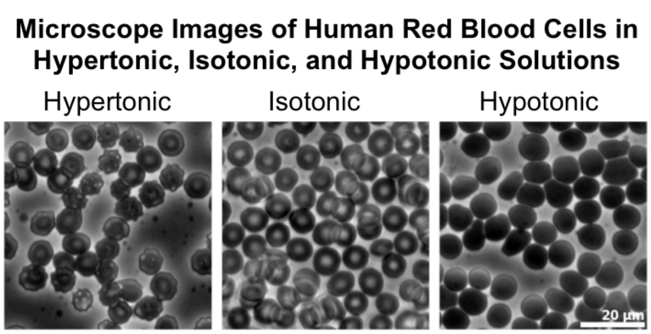
We will only discuss osmosis for ***aqueous solutions***, therefore for our purposes, **osmosis** *is the net transport of* __________ ____________ *from a* solution with a ***lesser solute*** particle concentration *through a semipermeable membrane* to a solution with a ***greater solute*** particle concentration.

Osmosis is very important in biology because cell membranes are semipermeable.

The difference in solute particle concentration (osmolarity) between the ___________ of the cell and the ___________ solution has important implications in maintaining the viability of the cell.

Consider the three different cases for *the solution that surrounds a cell*:





1. In a _________________ ***solution***, there is a greater solute particle concentration outside the cell than inside of the cell, so there is a net flow of water from the inside to the outside of the cell. This results in the shrinking of the cell.

2. In an _________________ ***solution***, the concentration of solute particles is the same on the inside and outside of the cell, therefore the flow of water in and out of the cell *are equal* and the cell maintains it natural and healthy (viable) shape.

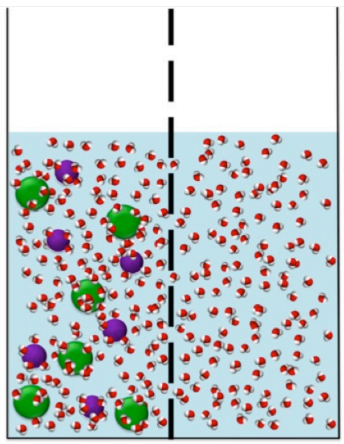
3. In a _________________ ***solution***, there is a lesser solute particle concentration outside the cell than inside the cell, and there is a net flow of water from the outside to the inside of the cell. This results in the swelling and possible bursting of the cell.

The pressure associated with the ***transport of water*** in the osmosis process is called __________ __________.

Initial State:
Equal amounts of liquid are placed on opposite sides of a membrane. Saltwater on one side and pure water on the other side.

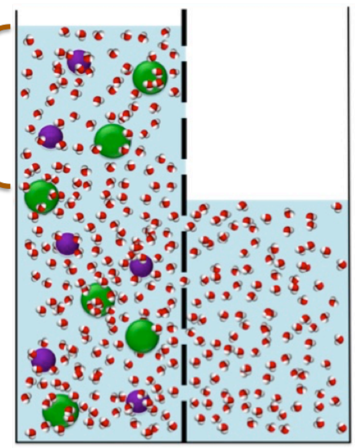
Osmosis

Final State:
Water molecules moved from the chamber with pure water (right side) to the side with greater solute concentration (left side):

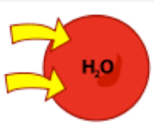


The water levels change ***until*** the pressure caused from the difference in water column heights **is equal to the *osmotic pressure*.**

The greater the difference in ____________ between each side of the semipermeable membrane, the greater the ***osmotic pressure***.



For membranes that form a continuous enclosure around a solution, such as those of biological cells, the *osmotic pressure* is the pressure required to stop the net transport of water into *or* out of cells.



Hypotonic

When the solution around a cell becomes ***hypotonic***, there is a lesser solute particle concentration outside the cell than inside the cell, and therefore there is a net flow of water from the outside to the inside of the cell.

As water flows into the cell, the pressure on the inside of the cell increases (imagine inflating a balloon).

The pressure in the cell will continue to increase until either (1) the ***osmotic pressure*** is reached and osmosis stops, **or** (2) the cell bursts. The bursting of red blood cells is called *hemolysis* and is evidenced by the appearance of red color in the blood *plasma*.

Understanding Check: Which of the following systems (#1 or #2) would have a ***greater*** **osmotic pressure**:

System #1: 1.00 M sodium chloride on one side of a semipermeable membrane and pure water on the other side.

or

System #2: 0.500 M sodium chloride on one side of a semipermeable membrane and pure water on the other side.

HINT: The greater the **difference in osmolarity** between each side of the semipermeable membrane, the **greater** the *osmotic pressure*.

Understanding Check:

Which of the following systems (#1 or #2) would have a ***greater*** **osmotic pressure**:

System #1: 1.00 M NaCl on one side of a semipermeable membrane and pure water on the other side.

or

System #2: 1.00 M $Fe_2(SO_4)_3$ on one side of a semipermeable membrane and pure water on the other side.