CHAPTER 15

Solutions

15.1 What are Solutions?

Solutions are homogeneous mixtures that are grouped according to physical state.

Mixtures have no bonding
Mixtures involve intermolecular forces

Characteristics of Solutions

Solute: substance that dissolves
Solvent: dissolving medium
Solutions can be: *(see table 15.1 for examples)*
- Gases: Air
- Liquids: salt water, carbonated water
- Solids: steel, brass, bronze (alloys)

The nature of solvent and solute affects whether a substance will dissolve.

- **Soluble: a substance that dissolves in a solvent**
- **Insoluble: a substance that does not dissolve in a solvent**

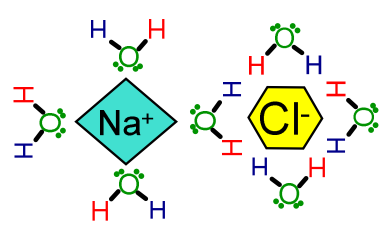
- **Immiscible: liquids that do not mix *(oil and vinegar)***
- **Miscible: liquids that mix and are soluble in each other**

Solvation

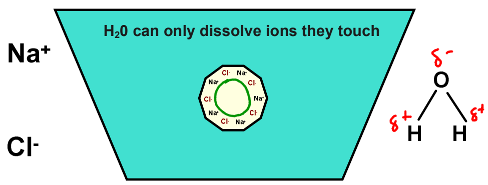
- **Solute particles are separated and mixed throughout the solvent.** ***Dissolving***
- **Negatively and positively charged ions become surrounded by solvent molecules.**
  - > Polar solvents dissolve ionic and polar compounds.
  - > Non-Polar solvents dissolve non-polar compounds.

Water molecules surround the Na^+ and Cl^- ions.

Making the solution uniform when dissolving occurs.



- **Solvation is a surface phenomenon!**
- **Water dissolving NaCl.**



<u>Aqueous Solutions in Ionic and Molecular Compounds</u>

- **Like dissolves like.**
- **Salt and water: ionic & polar**
- **Sugar (molecular) and Water: polar & polar**
- **Gasoline and grease: non polar molecular molecules**

- **In some ionic compounds, the attraction of the ions is stronger than the attraction from the water, therefore, dissolving will not occur.**

<u>What determines how fast solutes dissolve</u>?

1. amount of surface area
2. force of collisions with solvent
3. frequency of collisions with solvent

3 factors increase the rate of solvation:

- Agitation
- Temperature
- Particle Size

} **These increase the collisions between solvent and solute**

<u>Agitation</u>

solvent is brought into contact with the solute more frequently with greater force.

- **Shaking and stirring**

<u>Temperature</u>

- **increases the force and frequency of solvent and solute**

- **Heat makes particles move faster**

Particle Size

powder has a large surface area:volume ratio, therefore has most of its particles exposed to a solvent.

Large Crystals have a lot of surface area, but a lot of volume as well, which make it difficult for solvent to be exposed to the particles inside.

Heat of Solution

The overall energy change that occurs during solvation.

Solute and solvent particles require energy to separate them. (Endothermic)

When they mix together, energy is released. (Exothermic)

Solubility

The maximum amount of solute for a given amount of solvent at a constant temperature & pressure.

- How much solute can a given amount of solvent hold?
- Table 15-2 on page 457
- Solubility changes as temperature changes.

For Example:

- **36.2 grams of NaCl in 100 grams of water @ 25 °C.**

$$\frac{36.2\text{- g NaCl}}{100\text{-g } H_2O} \text{ at 25 °C}$$

Solvation in an unsaturated solution

- **Rate of Solvation is greater than Rate of Desolvation**
- **Particles move from solid to solvated state and back to solid again, but the amount of dissolved solute is increasing.**

Dynamic Equilibrium in a saturated solution.

- **Rate of Solvation = Rate of Desolvation**
- **Dissolving = Crystallization**
- **Particles move from solid to solvated state and back to solid again, but the amount of dissolved solute remains constant.**

Solubility is generally expressed in grams of solute per 100 g of solvent.

$$100\text{-g} = 100\text{-mL} \qquad \frac{\text{grams of solute}}{100\text{-g } H_2O} \text{ at 20 °C}$$

Saturated Solution: contains maximum amount of dissolved solute for a given amount of solvent.

Unsaturated Solution: less then the maximum

Factors Affecting Solubility

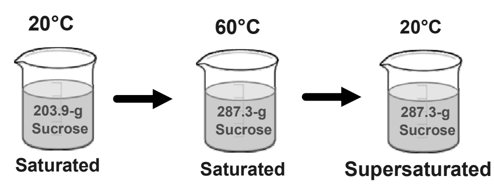
- For liquids and solids, the solubility of most substances increases with temperature because the particles are colliding with more energy.
- Gases are less soluble at high temperatures because they escape into the air as kinetic energy increases.

- **The solubilities of gases are greater in cold water than in hot water**
- **Hotter water has higher vapor pressure, which allows gas to escape**
- **The components of air become less soluble as the temperature of the water rises.**

Thermal Pollution

- **Industrial plants remove cold water and dump hot water back into the lake.**
- **This increases the temperature of the water, decreasing the dissolved O_2.**
- **Warm vs. Cold Soda can**

Supersaturated Solutions



Pressure and Solubility

- **Gases are more soluble in liquids when the external pressure is higher**
- **Why does pop fizz when you open it?**
- **How do they get the carbon dioxide to dissolve in the solvent?**
  - **If the pop fizzes at sea level, what could bottlers do to increase the solubility of a gas?**

Henry's Law

- **At a given temperature the solubility of a gas in a liquid is directly proportional the pressure of the gas above the liquid**
- **In other words, as the pressure of the gas above the liquid increases, the solubility of the gas increases**

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Example

A gas has a solubility in water of 10.5 g/L at 15 °C and 4.49 atm of pressure. What is the solubility of the gas in water at 15 °C and 6.07 atm of pressure?

$$\frac{10.5\text{ - g/L}}{4.49\text{ atm}} = \frac{\quad}{6.07\text{ atm}}$$

15.2 Solution Concentration

The concentration of a solution is a measure of the amount of solute that is dissolved in a given quantity of solvent.

- **The unit for concentration is Molarity (*M*)**

Concentration= amount of solute dissolved in a solvent.

Molarity

Dilute solutions have a *small* amount of solute.

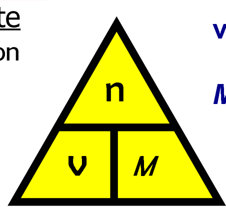
Concentrated solutions have a *large* amount of solute.

How do I compare 1.0-M HCl with 2.0-M HCl?

How do I compare 1.0-M HCl with 1.0-M H_2SO_4?

Molarity (*M*)

$$\frac{\text{Moles of solute}}{\text{Liters of solution}}$$



n = moles

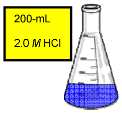
v = volume

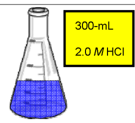
M = molarity

Key Concept!

Is it possible to have 1-mL of 5.0*M* solution?

If you poured 5.0*M* HCl into 3 different size containers, would the molarity in each container change?

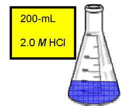


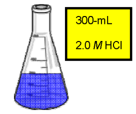


Are the substances different because they have different amounts in them?

Does each flask have 2 moles of HCl dissolved?

Which flask has more moles of HCl dissolved?





$$\frac{2.0\text{-moles}}{1.0\text{-L}} = \frac{\quad}{0.2\text{-L}}$$

$$\frac{2.0\text{-moles}}{1.0\text{-L}} = \frac{\quad}{0.3\text{-L}}$$

<u>Example</u>

What is the molarity of NaCl in sea water if it contains 4 moles of NaCl per 500-mL?

$$\frac{4\text{ moles}}{0.5\text{-L}} =$$

<u>Example</u>

What is the molarity of NaCl in sea water if it contains 20.5 g of NaCl per 750-mL?

20.5-g NaCl	1 mole
	58.5-g

=

$$\frac{.350\text{ moles}}{0.75\text{-L}} =$$

<u>Example</u>

How many moles of NaCl are dissolved in 100.0 mL of a 5.0 *M* solution?

$$n = M \times V = \frac{5.0\text{-moles}}{1.0\text{-L}} \times 0.1\text{-L} =$$

<u>Example</u>

How many grams of NaCl are contained in 450.0 mL of a 2.0 *M* solution?

$$n = M \times V = \frac{2.0\text{-moles}}{1.0\text{-L}} \times 0.45\text{-L} =$$

0.9 moles	58.5-g
	1 mole

Example

What volume (in mL) of 12.0 *M* HCl is needed to have 5.00 moles of HCl?

$$V = \frac{n}{M} = \frac{5.0 \text{ moles}}{\frac{12.0\text{-moles}}{1.0\text{-L}}} =$$

Making Molar Solutions

- **Remember, solute has some volume, so you cannot always add 1.0-L of solvent**

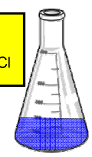
1. **Convert the moles needed into grams**
2. **Mass the correct mass of solute**
3. **Put the solute in a 1-L graduated cylinder**
4. **Add solvent until you reach 1-L.**

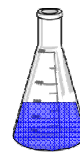
Making Dilutions

- **Adding water to an existing solution.**
- **Increasing the volume to lower Molarity**
- **The number of moles of solute does not change when a solution is diluted**

0.5-L

8.0 *M* HCl





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2.0 *M* HCl

How much water must be added to make 2.0 M HCl?

- $M_1 \times V_1 = M_2 \times V_2$

- Volume can be in mL or L as long as they are the same on both sides.
- The number of moles are the same on both sides.
- Stock solution is M_1 or the solution of greater concentration.

Example

Concentrated H_2SO_4 is 18.0 *M*, what volume is needed to make 4.50 L of 1.00 *M* solution? How much water must be added to it?

M_1: 18.0 *M*
V_1:
M_2: 1.0 *M*
V_2: 4.50 - L

The number of moles in each solution are the same!!!
The difference between the 2 solutions is the amount of water you add.

Mixing Solutions

Calculate the final concentration if 5.0-L of 2.0*M* NaCl and 3.0-L of 4.0*M* NaCl are mixed?

Step 1: Find total volume *(Same units)*

Step 2: Find total moles

Step 3: Calculate molarity

Example

Calculate the final concentration if 4.00 L of 5.0*M* NaCl and 6.00 L of 2.0*M* NaCl are mixed?

Example

Calculate the final concentration if 4.00 L of 5.0*M* NaCl and 6.00 L of 2.0*M* NaCl are mixed, then you added 5.0-L of water to the mixture.

Percent Solutions

- Determining the percentage of solute in any given solution.
- Not the same as molarity because every solute has a different molar mass
- Important when comparing different solutions!!
  -comparing 1.0-*M* HCl vs. 1.0-*M* H_2SO_4

Percent Mass vs. Percent Volume

$$\frac{\text{mass of solute}}{\text{mass of solution}} \qquad \frac{\text{volume of solute}}{\text{volume of solution}}$$

1-mL of water = 1-gram of water

Percent Mass Examples

% mass is always expressed in g/g.

If you are given molarity, then you can always determine the mass per 1.0-L.

% m/m of 2.0-*M* NaCl Molar Mass NaCl = 58.5	% m/m of 2.0-*M* $MgCl_2$ Molar Mass $MgCl_2$ = 95	100-g of NaCl is dissolved in 2.00-L of solution.

Percent Volume Example

What is the %(v/v) of ethanol when 75-mL is diluted to a volume of 250 mL with water?

Example

What is the percent (mass/mass) of NaCl in a 200.0 mL solution containing 14 grams of NaCl?

Percent Mass Example

How many grams of glucose would you need to prepare 2.0 L of 2% (m/m) glucose solution? How much water was used?

Example

What is the %(v/v) of ethanol when 50-mL is diluted with 150 mL of water?

Molality

- **Moles of solute dissolved in 1-kilogram of solvent**
- **Often more useful than *molarity* because volume can change with temperature, mass will not.**

$$\text{Molality } (m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Molality Example

What is the molality of a solution containing 13.7 grams of NaCl dissolved in 1250.0 grams of water

1. **Grams of NaCl to moles**
2. **Grams of H_2O to kilograms**

Mole Fractions

When you know the moles of solute & solvent

The sum of the mole fractions is always 1, since they make up 100% of the solution.

This is a ratio of moles of solute to moles of solute + solvent. ***(X=mole fraction)***

$$X_A = \frac{n_A}{n_A + n_B} \quad X_B = \frac{n_B}{n_A + n_B}$$

Mole Fraction Example

What is the mole fraction of HCl in an aqueous solution consisting of 29.3% HCl by mass?

1. **Convert 29.3-g HCl to moles**
2. **Convert 70.7-g H_2O to moles**
3. **Solve for mole fraction of HCl**

$$X_{HCl} = \frac{n_{HCl}}{n_{HCl} + n_{water}}$$

15.3 Colligative Properties of Solutions

- Colligative properties depend on the number of particles dissolved in a given mass of solvent
- The collection of particles

1. Freezing Point depressoin
2. Vapor Pressure lowering
3. Boiling Point elevation

Electrolytes in Aqueous Solution

- Compounds that conduct an electric current in aqueous solution or the molten state are electrolytes
- All ionic compounds: the strength of an electrolyte is determined by the amount of dissociation of the ions.
- Insoluble ionic compounds are ionic compounds in molten state.
- Barium sulfate only in molten state.
- **1 mole of NaCl dissociates as 1 mole of Na^+ and 1 mole of Cl^-, therefore, 2 moles of dissolved particles.**

- When a weak electrolyte is in solution, only a fraction of the solute exists as ions
  - > (Molecular compounds, Ammonia, Water)
- When a strong electrolyte is dissolved, almost all of the solute exists as ions. (More dissociation)
  - > (Acids, Bases, Soluble Salts)

Nonelectrolytes in Aqueous Solution

- Do not conduct an electric current in either aqueous solution or the molten state.
- Many molecular compounds are nonelectrolytes because they are not composed of ions, hence, do not ionize or dissociate
- Compounds made of carbon.

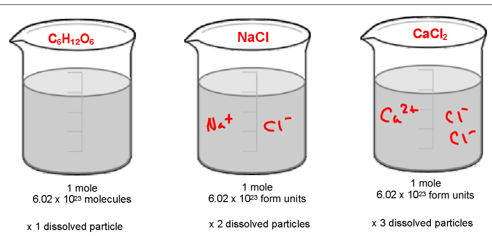
3 Colligative Properties

- Factors that change when Particles dissolve
- For the same substance, we compare Molarity!
- For different substances, we compare the number particles dissolved for 1 mole of a substance.

Dissolved Particles

The more dissolved particles, the greater change in the colligative properties.

- Glucose: 1 molecule *(1 particle)*
- NaCl: 2 ions *(2 particles)*
- $CaCl_2$: 3 ions *(3 particles)*



Which will have the most dissolved particles in 1-L of solution?

3.0-*M* NaCl	1.8-*M* $AlCl_3$
3.0 moles x (6.02 x 10^{23}) x 2 particles (Na^+ & Cl^-)	1.8 moles x (6.02 x 10^{23}) x 4 particles (Al^{3+}, Cl^-, Cl^-, Cl^-)

3 Colligative Properties

1. vapor pressure lowering
2. boiling point elevation
3. freezing point depression

30º Celsius

Water	Salt Water
Vapor Pressure 4.24-kPa Boiling Point = 100ºC Vapor pressure of the water must be 101.3-kPa to boil.	Vapor Pressure 3.69-kPa Boiling Point = 102 ºC Vapor pressure of the water must be 101.3-kPa to boil.

Vapor Pressure

- Substances with high vapor pressure evaporate quickly. Alcohol + Ammonia (molecular)
- Substances with low vapor pressure evaporate slowly. Water + salt water

Freezing Point

- The more dissolved particles, the colder it needs to be for a solvent to freeze.
- Antifreeze in cars.
- Why is $CaCl_2$ better to melt ice than NaCl?

Boiling Point

- Vapor pressure of liquid equals vapor pressure of the air.
- The higher percentage of solute, the higher the boiling point.
- As water evaporates, the concentration of solute rises.

Boiling Point

- Salt water has lower vapor pressure than fresh water.
- Salt water will boil at 102 °C at a location with an atmospheric pressure of 1 atm.
- Pasta will cook faster at 102 °C than 100 °C.

High Altitude Directions

- At high altitudes, the atmospheric pressure is less than 101.3 kPa. (Denver = 84.0 kPa)
- In Denver, water boils at 94 °C so water reaches a boil faster but food cooks slower than at sea level.

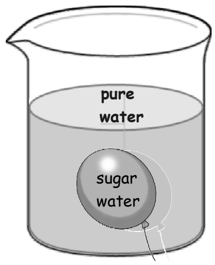
This is the solubility of salt in water at 25 °C.

$$\frac{\text{36.2- g NaCl}}{\text{100-g } H_2O} \text{ at 25 °C}$$

- **What happens to the solubility if temperature increases or decreases?**
- **How many grams of NaCl can be dissolved in 300 grams of water?**

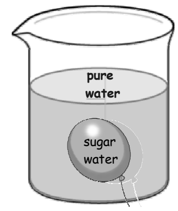
Osmosis

- The diffusion of solvent particles across a semipermeable membrane
- Plays a big role in biological systems as water moves back and forth across the membrane
- Sugar molecules cannot cross the membrane, therefore, more water goes into the balloon that out of it.



Osmotic Pressure

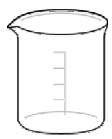
- Is a colligative property
- Pure water consists of all water molecuiues, therefore, it is easier for them to move across the membrane than the sugar & water molecule mixture.
- Osmotic pressure is based on the number of molecules traveling across the membrane.



15.4 Heterogeneous Mixtures

No Dissolving

- **Suspensions:** mixtures from which some of the particles will ***settle*** slowly upon standing
- Heterogeneous You can Tell the difference
- Can be filtered particle are large
- Ex. muddy water, italian dressing



Colloids

- Mixtures containing particles that are intermediate in size between those of suspensions and true solutions.
- No settling, but no dissolving either.
- Cannot be filtered.
- Ex. Glue, jello, milk, butter

- The properties of colloids differ from those of solutions and suspensions.
- Many colloids are cloudy or milky in appearance but look clear when they are very dilute.
- The particles in a colloid cannot be retained by filter paper and do not settle out with time.

- Emulsions are colloidal dispersions of liquids in liquids.
- Soaps and other emulsifying agents allow the formation of colloidal dispersions between liquids that do not ordinarily mix, such as oil and water

Tyndall Effect

- **Exhibited by dilute colloids.**
- **Colloidal particles exhibit this by scattering visible light in all directions.**
- **Suspensions also exhibit the Tyndall effect, but solutions never do.**
- **Why don't solutions exhibit the Tyndall Effect?**

Brownian Motion

- The chaotic movement of colloidal particles.
- Brownian motion is caused by the water molecules of the medium colliding with the small, dispersed colloidal particles.
- This motion causes collisions of particles with electrostatic forces, which prevents settling.
- Heating actually can cause a colloid to settle, since the kinetic energy becomes so high, particles cannot stay suspended.

Solution	Colloid	Suspension
Dissolving Cannot be filtered Small Particles Uniform Throughout Homogeneous No Settling No Tyndall Effect No Brownian Motion		