Chapter 13: Solutions

SOLUTION: A **homogeneous** mixture of two or more substances

- Composition can vary from one sample to another
- Appears to be one substance, though really contains multiple materials

The components of a solution are:

a) **Solute** - the solution **component present in small amount** as compared to the solvent.

b) **Solvent** - the solution **component present in greatest amount**. (Usually the solvent's physical state is retained).

- In some cases, when a solute is dissolved in a solvent, the solute seems to disappear and take on the state of the solvent (i.e. dissolving a solid solute in a liquid solvent).

- Solutions in which the solvent is water are called **Aqueous Solutions**

Solution phase	Solute phase	Solvent phase	Example
Gaseous solutions	Gas	Gas	Air (mostly N_2 and O_2)
Liquid solutions	Gas Liquid Solid	Liquid Liquid Liquid	Soda (CO_2 in H_2O) Vodka (C_2H_5OH in H_2O) Seawater (NaCl in H_2O)
Solid solutions	Solid	Solid	Brass (Zn in Cu)

Solubility

The **solubility** of a solute is the maximum amount of a solute that will dissolve in a given amount of a solvent. (Common units are g solute/100 g solvent)

- A **soluble substance** - dissolves to an appreciable extent in a given solvent
- An **insoluble substance** - none (or close to none) of the substance dissolves in a given solvent

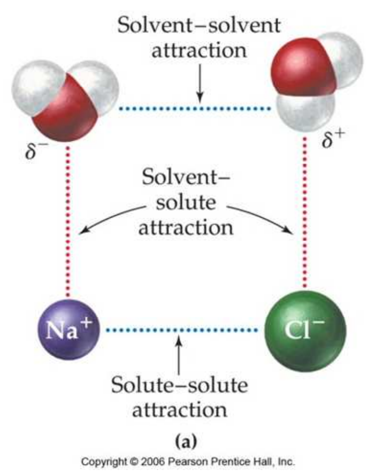
Will It Dissolve??

For a solute to dissolve in a solvent two types of interparticle attraction **must be overcome**:

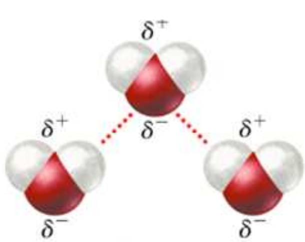
1) Attractions between solute particles
   SOLUTE-SOLUTE attractions

Example: sodium chloride dissolving in water

Ionic Bonds in solid NaCl



2) Attractions between solvent particles
   SOLVENT-SOLVENT attractions



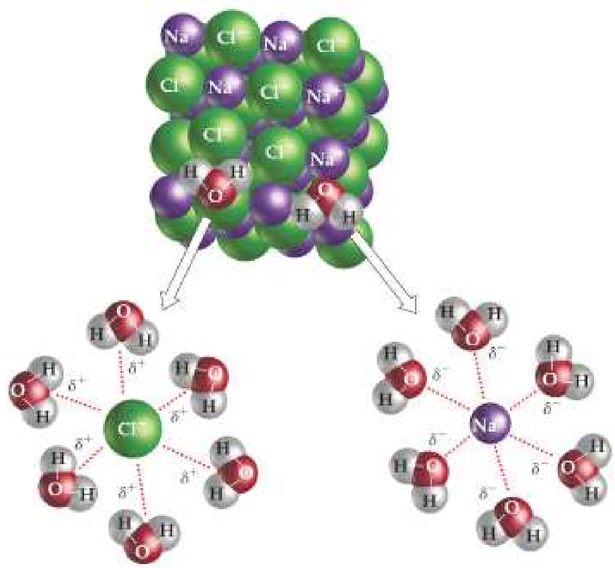
Hydrogen-bonding between water molecules

New attractive force - the attraction between solute and solvent particles

Solubility of Ionic Compounds

When a solute dissolves in a solvent, the new attractive force is the attraction between solute and solvent particles

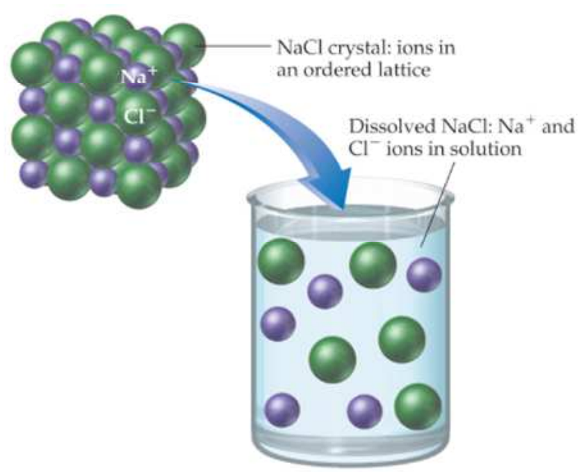
SOLUTE-SOLVENT attractions



When materials dissolve, the solvent molecules surround the solute particles due to the solvent's **attractions for the solute** - called **SOLVATION**

Solvated ions are effectively isolated from each other

In order for a compound to be soluble in a particular solvent, the solute-solvent attractive forces **must be HIGHER IN ENERGY** than the sum of the solute-solute and solvent-solvent attractive forces that are disrupted when the solution is formed.



If the solute-solvent attractive forces are **lower in energy** than the sum of the solute-solute and solvent-solvent attractions, then the compound WILL NOT dissolve appreciably - the compound will be insoluble.

Solubility of Ionic and Covalent Compounds

In general, a chemical will dissolve in a solvent **if it has a similar structure to the solvent**

- When the solvent and solute structures are similar and the solute dissolves, the solvent molecules will attract the solute particles at least as well as the solute particles to each other.

LIKE DISSOLVES LIKE

Ions are attracted to polar solvents

- Many ionic compounds dissolve in water

    If the total charges of the ions **< 4** then usually the ionic compound will dissolve, if > 4 then usually it will not dissolve

Polar molecules are attracted to polar solvents

- Examples: table sugar (sucrose), ethyl alcohol, and glucose all dissolve well in water

    These molecules have either multiple OH groups and/or a smaller hydrocarbon region to allow hydrogen bonding to readily occur with water

Nonpolar molecules are attracted to nonpolar solvents

- Example: I_2 dissolves in hexane

Many molecules have both polar and nonpolar structures—whether they will dissolve in water depends on the kind, number, and location of polar and nonpolar structural features in the molecule.

Saturation and Miscibility

Decide if each of the following will be significantly soluble in water:

Potassium iodide, KI   octane, C_8H_{18} methanol, CH_3OH

Copper, Cu   cetyl alcohol $CH_3(CH_2)_{14}CH_2OH$ iron(III) sulfide Fe_2S_3

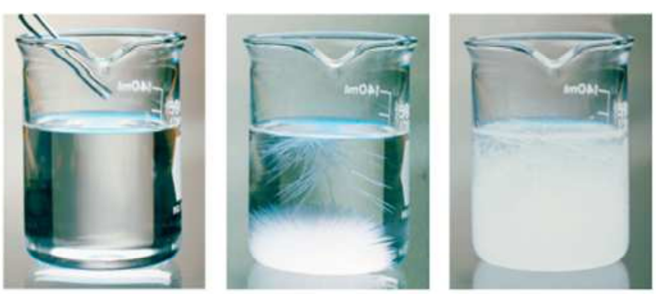
There is usually a limit to the solubility of one substance in another:

SATURATION

Saturated solutions - have the maximum amount of solute that will dissolve in that solvent at that temperature.

Unsaturated solutions - can dissolve more solute - contains less than the maximum amount possible.

Supersaturated solutions are **temporarily** holding more solute than they should be able to at that temperature (unstable).



(a) (b) (c)

A supersaturated solution has more dissolved solute than the solvent can hold. When disturbed, all the solute above the saturation level comes out of solution.

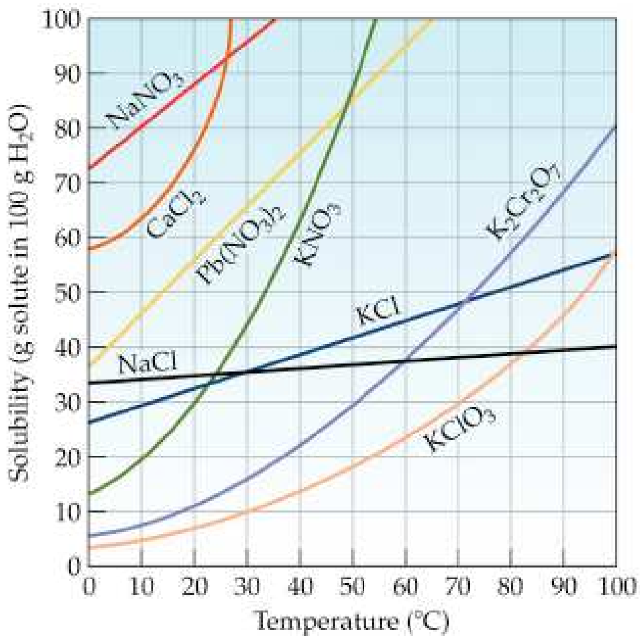
Miscibility - ability of one liquid to dissolve in another liquid

- Can have completely miscible, partially miscible or immiscible liquids

Relationship Between Solubility and Temperature

The solubility of the solute in a solvent depends on the temperature

- HIGHER temperature = HIGHER solubility of a **solid solute in a liquid solvent**



As temperature INCREASES, solubility INCREASES

Magnitude of the effect of temperature depends on the identity of the solute

Effect of Temperature on the Solubilities of Some Ionic Compounds

Because of this temperature dependence, solubility data **must** be expressed with the temperature at which the solubility of the substance was measured:

Ex. Solubility of KNO_3 in H_20 at 30°C = 43.0 g/100 g H_2O

Effect of temperature on the solubility of a **gas solute** in a **liquid solvent** is OPPOSITE to that of solids:

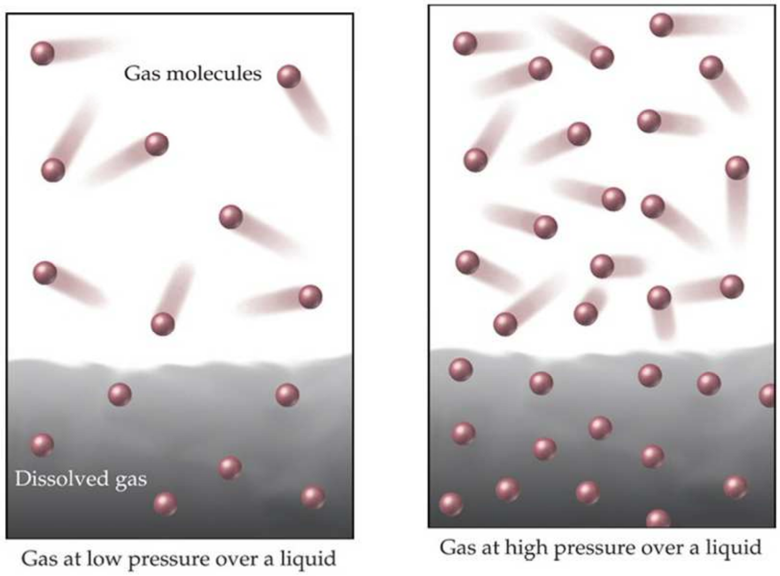
HIGHER temperature = **LOWER** solubility of gas in liquid

As temperature INCREASES solubility of a gas DECREASES

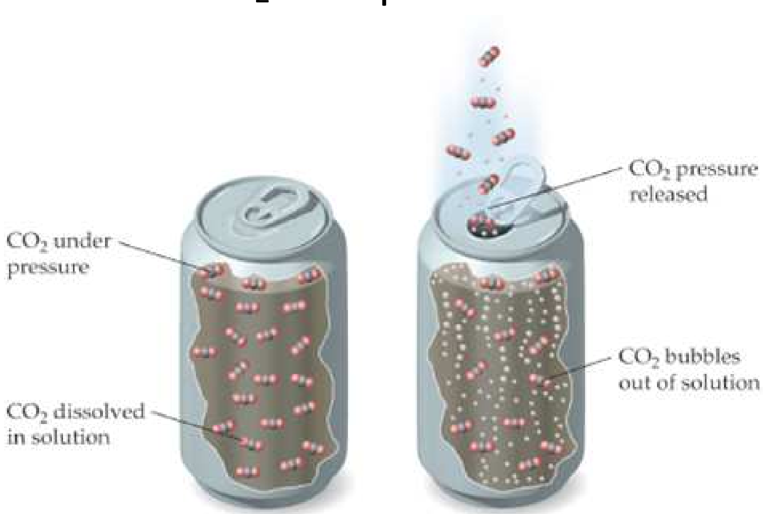
Relationship Between Pressure and Gas Solubility

The <u>solubility of a gas</u> in a liquid **INCREASES** as the **pressure INCREASES**

Higher pressure = higher solubility



Example: Dissolved CO_2 in Pop



When a can of pop is sealed, the CO_2 is under pressure. Opening the container lowers the pressure, which decreases the solubility of CO_2 and causes bubbles to form.

Solution Concentration: Mass Percent

Concentration - amount of solute present in a specified amount of solution

$$\text{mass \%} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100\%$$

Mass of Solute + Mass of Solvent = Mass of Solution

Calculate the mass percent of a solution that contains 11.5 g of NaCl in 155 mL of water at 25°C. (Density of H_2O at 25°C = 1.00 g/mL)

Bleach is a 6.25% NaOCl (sodium hypochlorite) solution in H_2O. What mass NaOCl is in 265 g bleach?

- When you're given a mass %, make a conversion factor out of it. **100 g** solution contains ___ g NaOCl

How would you prepare 250.0 g of a **2.00% by mass** aqueous sugar solution?

1st - Calculate how much sugar solute is needed:

2nd - Determine the mass of the solvent need to prepare the solution:

g(solution) = g(solute) + g(solvent)

3rd- Describe how to prepare this solution:

Molarity

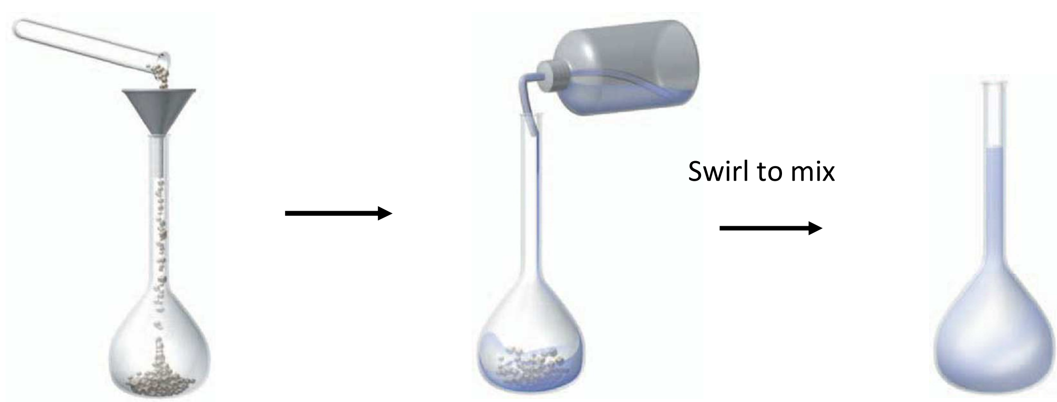
Molarity is the most common concentration unit in chemistry

$$\underline{\textbf{Molarity}}\ (M) = \frac{\text{Moles of Solute}}{\text{Liters of Solution}} = \frac{\text{mol}}{\text{L}}$$

Determine the molarity of a solution where 1.56 moles of NaCl is dissolved in water to give 556 mL of solution.

What is the molarity of a solution prepared by dissolving 25.0 g of NaCl in enough water to give 2.50 L of solution?

How would you prepare 1.00 L of a 1.00 M NaCl solution?



Weigh out 1.00 mole (58.45 g) of NaCl and add it to a 1.00 L volumetric flask

Add water to dissolve the NaCl, then add water to the 1.00 L mark

You have prepared 1.00 L of a 1.00 M NaCl solution

Molarity and Dilution

Can use Molarity as a conversion factor:

Liters → mol   or   mol → Liters

How many grams of NaCl are present in 53.0 mL of a 0.95 M solution of NaCl?

How many liters of 0.10 M NaCl can be prepared from 15.0 g of NaCl?

SOLUTION DILUTION

Dilution: adding solvent to an existing solution

Dilution will increase or decrease the concentration?

The dilution equation:

$$M_1V_1 = M_2V_2$$

M_1 and V_1 = initial molarity and volume
M_2 and V_2 = final diluted molarity and volume

What is the molarity of a solution made by diluting 65.0 mL of a 0.950 M HCl solution to a final volume of 120.0 mL?

What volume of 1.20 M HCl is needed to prepare 1.40 liters of a 0.300 M HCl solution?

Solution Stoichiometry

Recall from Chapter 8: Stoichiometry

Grams A ⟷ **Moles A** ⟷ **Moles B** ⟷ **Grams B**

Conversion factor: Use molar mass of A | Use **mole ratio from the balanced equation** | Use molar mass of B

When you have solutions involved, can use the outline above along with **an additional conversion**:

Grams A ⟷ **Moles A** ⟷ **Moles B** ⟷ **Grams B**

Use **molarity of A** as a conversion factor: **Moles A** ⟷ **Volume of A**

Use **molarity of B** as a conversion factor: **Moles B** ⟷ **Volume of B**

Determine the volume (in mL) of 0.500 M HCl that is required to react with 4.50 g of $Ca(OH)_2$ (molar mass = 74.10 g/mol)

$$2\ HCl_{(aq)} + Ca(OH)_{2(s)} \rightarrow CaCl_{2(aq)} + 2\ H_2O_{(l)}$$

Solution Stoichiometry

Grams A $\leftrightarrow$ **Moles A** $\leftrightarrow$ **Moles B** $\leftrightarrow$ **Grams B**

Use **molarity of A** as a conversion factor

Use **molarity of B** as a conversion factor

Volume of A **Volume of B**

How many liters of 0.105 M KI is needed to react with 0.0850 L of a 0.260 M $Pb(NO_3)_2$?

$$2\,KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2\,KNO_{3(aq)} + PbI_{2(s)}$$

Freezing Point Depression and Boiling Point Elevation

- We spread salt on icy roads and walkways to melt the ice
- We add antifreeze to car radiators to prevent the water from boiling or freezing (antifreeze = ethylene glycol)

When we add solutes to water, it changes the **freezing point** and **boiling point** of the water.

Colligative Properties

Any property of a solution whose value depends **ONLY** on the number of dissolved solute particles (and **not** on the type of solute) is called a

COLLIGATIVE PROPERTY

Freezing Point Depression: A Colligative Property

The freezing point of a solution is ALWAYS LOWER than the freezing point of a pure solvent

The difference between the freezing points of the solution and pure solvent (ΔT_f) is directly proportional to the **molal** concentration:

$$\Delta T_f = m \times K_f$$

where K_f = freezing point depression constant for the solvent
m = molality of the solution (**mol solute/kg solvent**)

Boiling Point Elevation: A Colligative Property

The boiling point of a solution is ALWAYS HIGHER than the boiling point of a pure solvent

The difference between the boiling points of the solution and pure solvent (ΔT_b) is directly proportional to the **molal** concentration:

$$\Delta T_b = m \times K_b$$

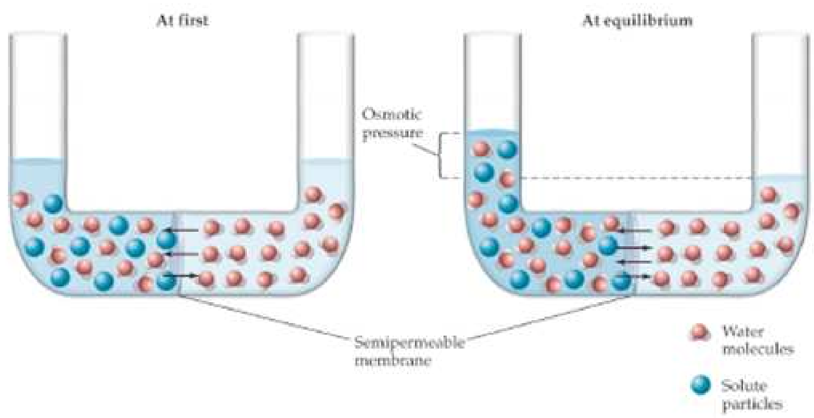
where K_b= boiling point elevation constant for the solvent
m = molality of the solution (mol solute/kg solvent)

Note: we will NOT do calculations with the equations above, but you should understand qualitatively what the equations mean

Osmosis: A Colligative Property

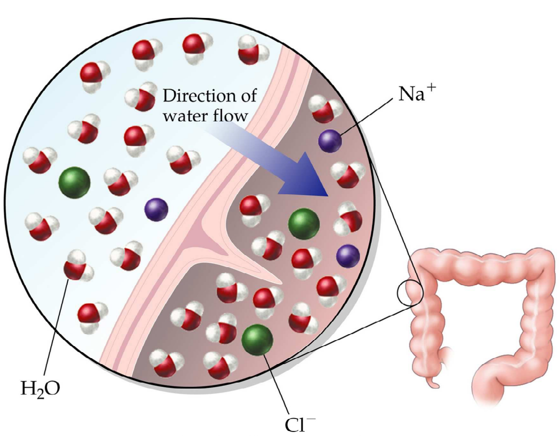
Osmosis is the process in which solvent molecules pass through a semipermeable membrane that <u>does not allow solute particles to pass through</u>

- Solvent flows **<u>to try to equalize</u>** the solute concentration on both sides
- Solvent flows from side of **low concentration of solute** **to high concentration of solute**



Osmotic pressure - pressure that is needed **to prevent osmotic flow of solvent**

Why does drinking seawater cause dehydration?



Because seawater has a higher salt concentration than your cells, water flows out of your cells into the seawater (in your intestine) to try to decrease its salt concentration.
The net result is that instead of quenching your thirst you become dehydrated.