SOLUTIONS

Chapter 13

Solutions

- At this point in the course, we have dealt with solutions on several occasions.
- We should address some formal definitions to clarify our understanding of this important topic.

Solution Terms

- A solution is composed of two parts:
  - The solute, which is the part of the solution being dissolved.
    - The solute need not be a solid.
  - The solvent, which is the medium in which the solute is dissolved.
- For example, sugar (a solute), can be dissolved in water (a solvent), to form a sugar-water solution.
- If the solute completely dissolves in the solution (usually the case), the solution is a homogeneous mixture

Solution Terms

- If the solution contains a relatively large amount of solute, we say that the solution is concentrated.
- If the solution contains a relatively small amount of solute, we say the solution is dilute.
- The key word here is *relative*.

Liquid Water

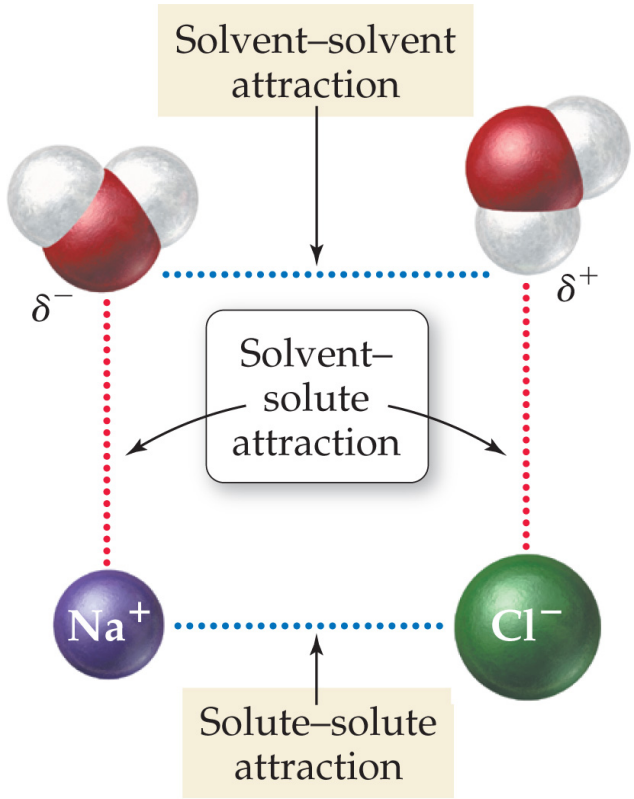
- The structure of water as a bent molecule is already familiar to us
- Among the most interesting properties of water is the strong attraction water molecules have for one another
- We know that oxygen is quite electronegative, while hydrogen is much less so
- The oxygen atom of a water molecule is strongly attracted to the hydrogen atoms of other water molecules
  - The attraction between different molecules is called an <u>intermolecular force</u> (more on these in Chapter 12)

Polarity of Water

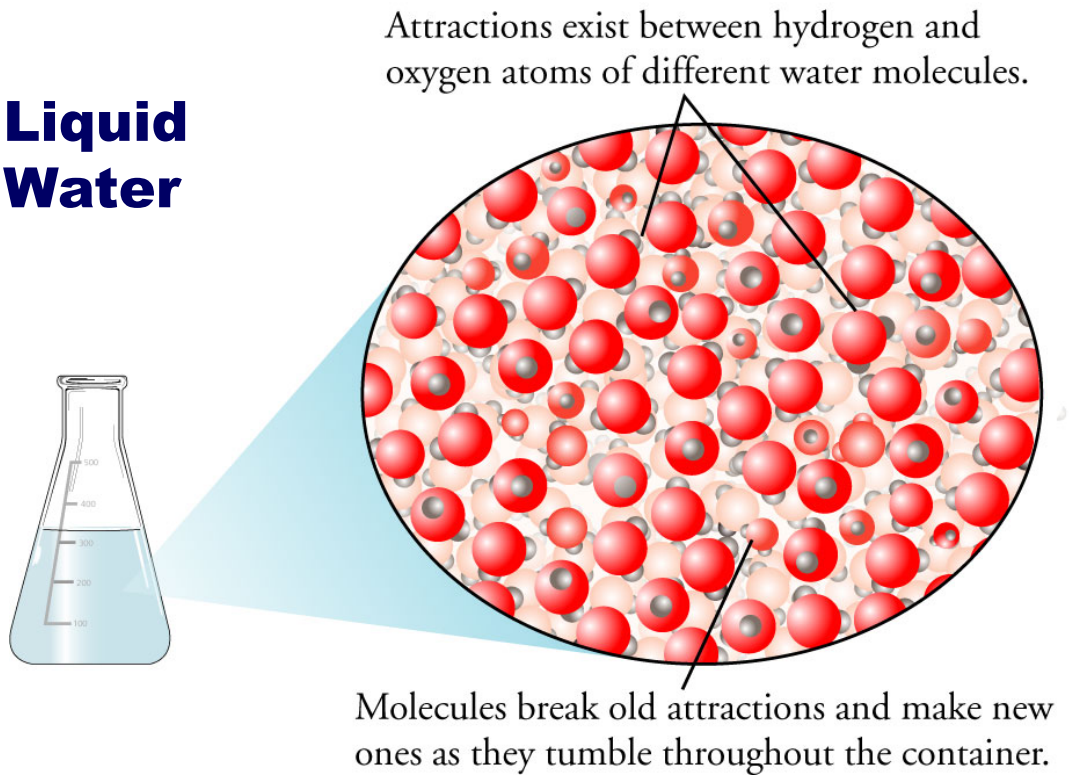
- We say that the oxygen atom in water bears a <u>partial negative charge</u> since the bonding electrons in water are more likely to be found near the electronegative oxygen
- Similarly, we say that the two hydrogen atoms bear a <u>partial positive charge</u>, as they do not get as great a share of the electrons
- These partial charges are indicated with the symbols $\delta+$ and $\delta-$

Polarity of Water

- Substances like water which have distinct positive and negative regions are called <u>polar</u>
  - Think of a magnet, with its positive and negative poles
- Substances which share electrons more equally (and do not have permanent partial negative and positive charges) are said to be <u>nonpolar</u>
- Polar compounds tend to mix well with other polar compounds, whereas nonpolar compounds generally mix well with other nonpolar compounds
  - We often say "Like dissolves like"



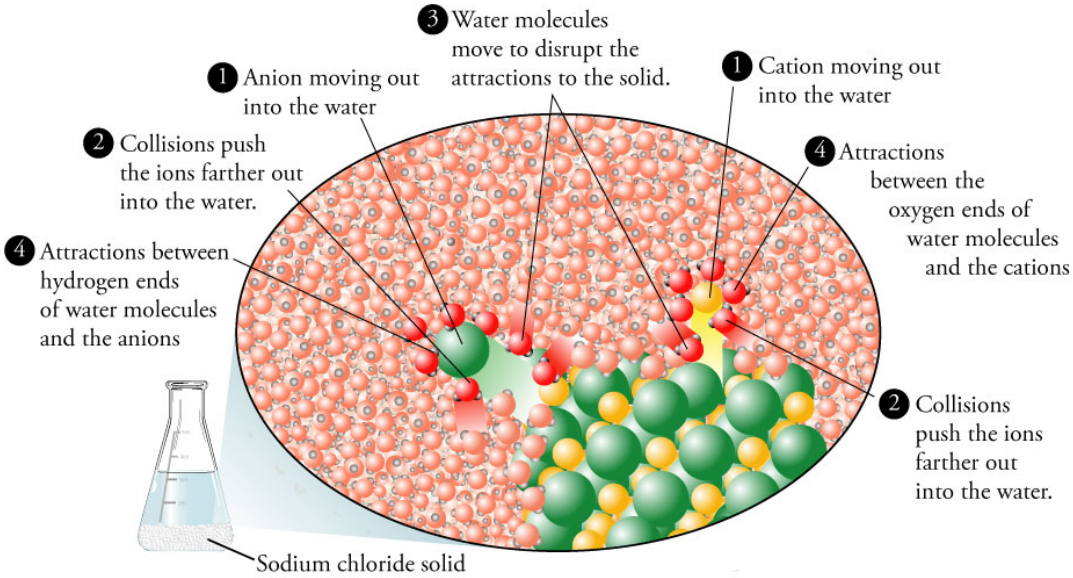
Liquid Water



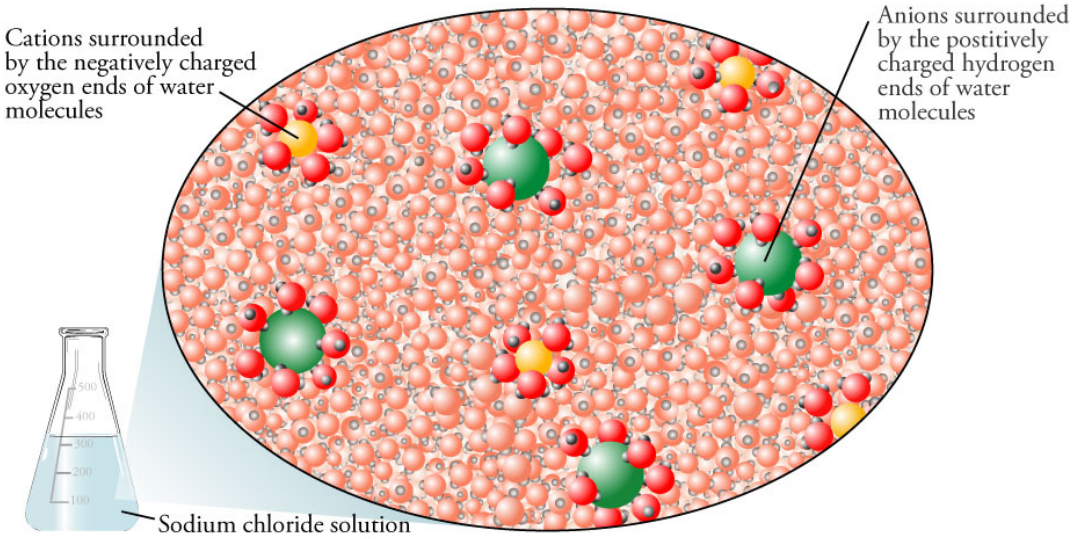
Water as a Solvent

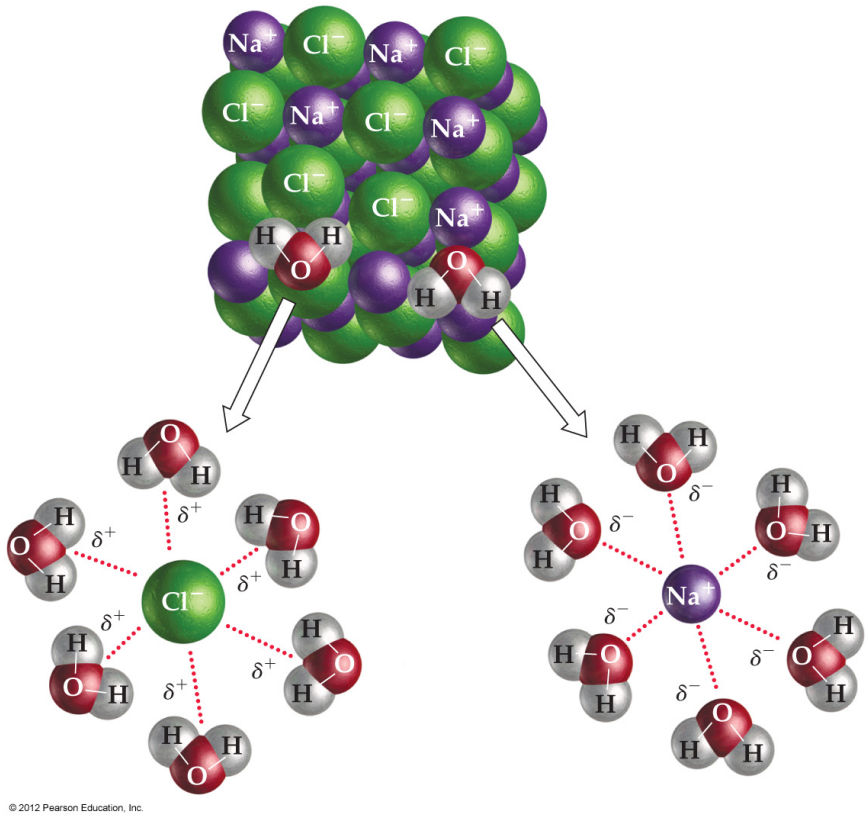
- Many ionic compounds are said to be <u>soluble</u> in water
  - We call the substance being dissolved (in this case, the ionic compound) a <u>solute</u>
  - The solute is dissolved in water, which is called a <u>solvent</u>
  - These two parts together give us a <u>solution</u>
- When the ionic compound is placed in water, the water molecules surround the ions in the crystal lattice
  - The partially-negative oxygen atoms of water help to pull the cations free of the crystal structure, while the partially-positive hydrogens pull away the anions
- When a compound is separated into ions in water, we say that the compound has <u>dissociated</u>

Solution of an Ionic Compound



Solution of an Ionic Compound (cont.)



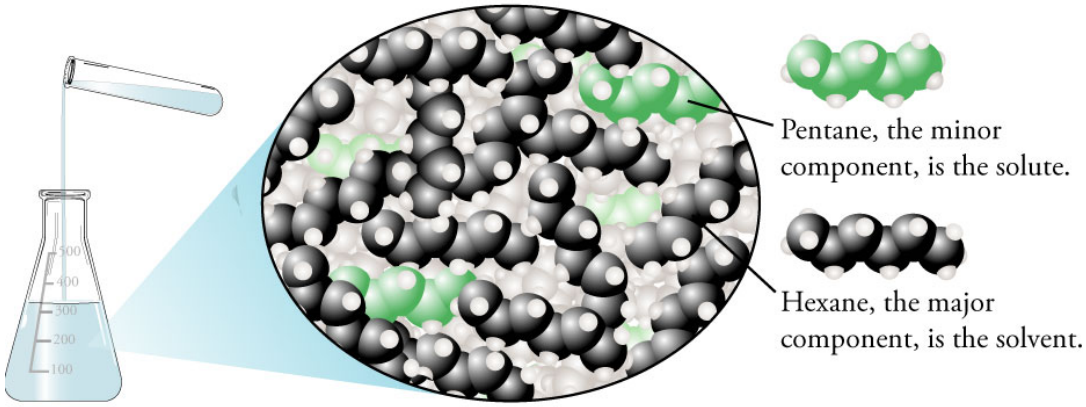


Liquid-Liquid Solution

Pentane and hexane are both non-polar and mix well with one another.

Would they mix well with water?

Would ionic compounds dissolve in pentane or hexane?

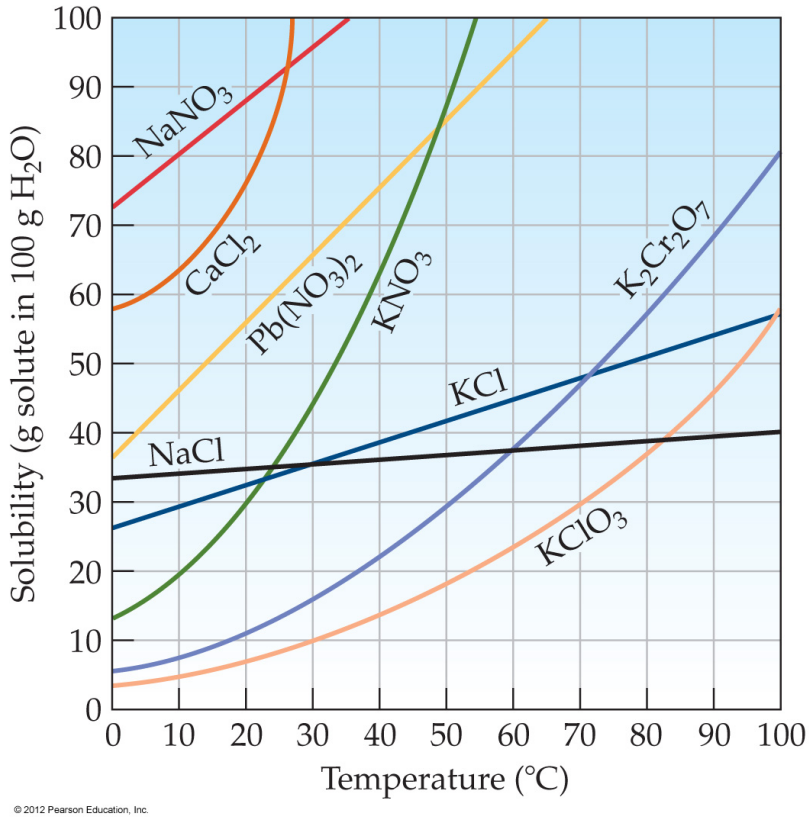


Solubility

- The solubility of a compound quantifies the amount of solute which can be dissolved in a given solvent.
- If a compound is soluble in a given solvent, this means that a significant amount can be dissolved.
- On the other hand, if a compound is insoluble, very little or none of it will dissolve.
- The exact solubility of the compound is often temperature dependent, with most non-gas compounds dissolving in greater amounts at higher temperatures.

Solubility

- Solubility is usually measured in grams of solute per 100. grams of solvent.
- These values can usually be read off of a graph comparing temperature with solubility.
- Example: What mass of potassium chloride can be dissolved in 2.50 kg of water at 40 °C if the solubility of KCl at this temperature is 42 g/100 g water (see graph)?



Saturation

- As we have seen, there is a maximum amount of solute which can be dissolved in a solvent at a certain temperature.
- If more solute has been added than the solution can dissolve, the solution is said to be <u>saturated</u>.
- If this solubility limit has not been reached for a solution, it is said to be <u>unsaturated</u>.

Supersaturation

- It is sometimes possible to dissolve more solute in a solution than it should be able to contain at a given temperature; if this is the case, the solution is supersaturated.
- To prepare a supersaturated solution, the following steps are carried out:
  - The solvent is heated, and a fairly large amount of solute is dissolved in it.
  - The solution is then slowly cooled.
  - If solute does not precipitate out when the solubility limit is exceeded, the solution is supersaturated.
  - If the solute does precipitate, it is usually in a much purer form than the original solute. This purification technique is called recrystallization.
- Disturbing a supersaturated solution by adding a "seed crystal" of the solute may cause excess solute to precipitate out.

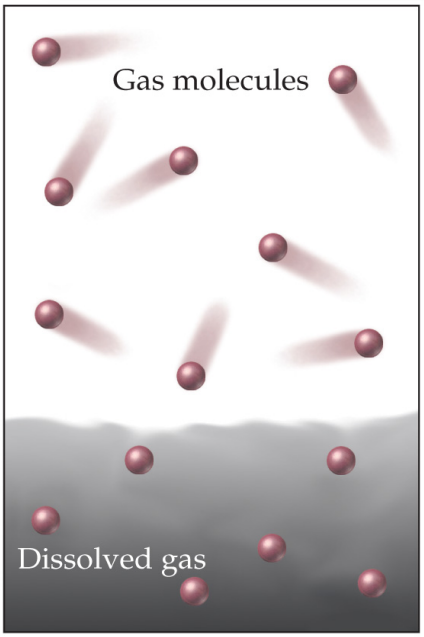
Solubility of Gases

- Unlike solids, gaseous solutes are usually more soluble at lower temperatures rather than higher ones
- The solubility of a gas in a solution also depends on the pressure of any gas (typically air) above the liquid
- According to Henry's Law, increasing the air pressure above a solution increases the solubility of the gas

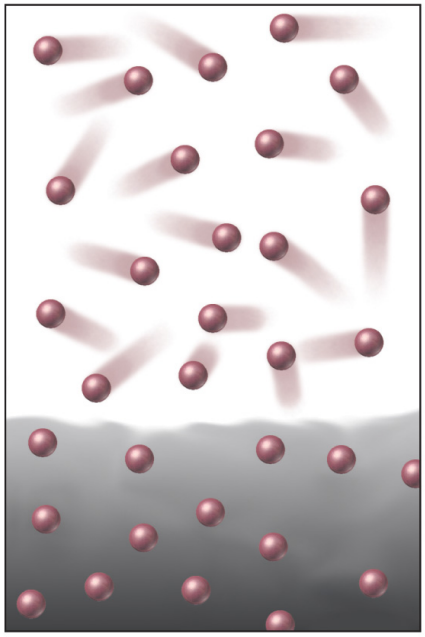
Cold soda pop: carbon dioxide more likely to stay in solution

Warm soda pop: carbon dioxide more likely to bubble out of solution

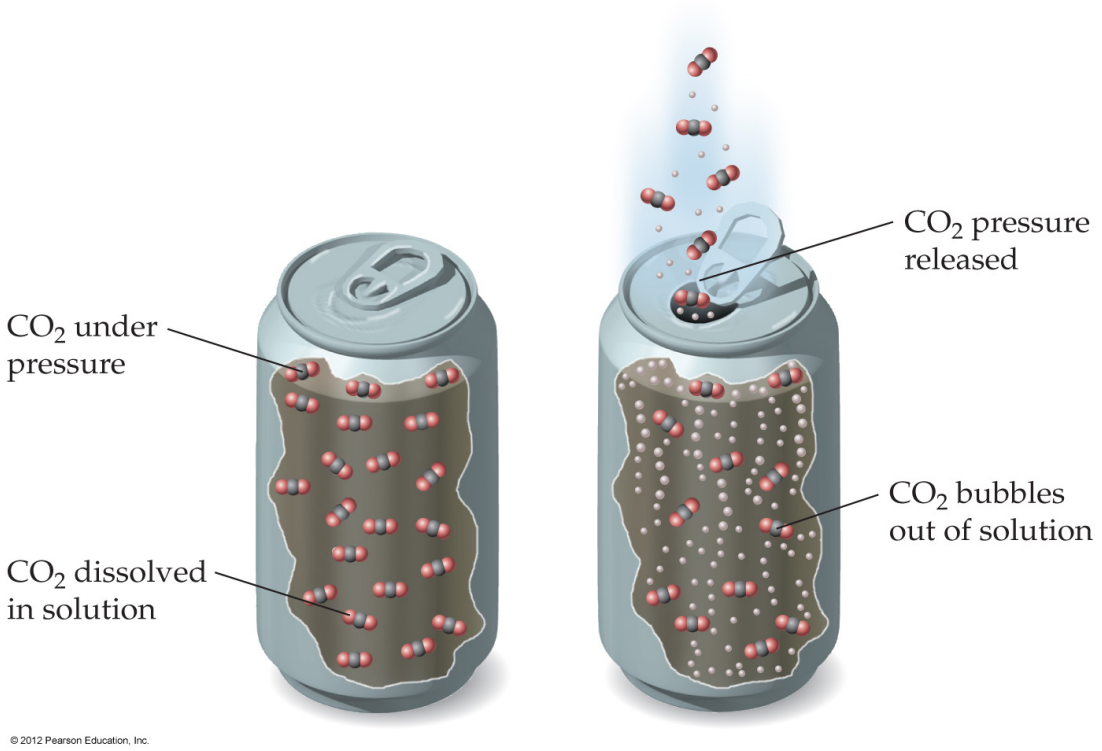
The solubility of a gas in a liquid increases with increasing pressure.



Gas at low pressure over a liquid



Gas at high pressure over a liquid



Concentration

- The <u>concentration</u> of a solution is a measurement of how much of a solute it contains
- There are several different methods we can use to measure concentration, including
  - Mass Percent
  - Molarity
  - Molality
  - Mole Fraction

Mass Percent

- The mass percent tells us what percent of the mass is derived from the solute:

$$\text{mass percent} = \frac{\text{g solute}}{\text{g solute} + \text{g solvent}} \times 100\% = \frac{\text{g solute}}{\text{g solution}} \times 100\%$$

- The symbol %(m/m) is used to specify the fact that we are measuring percents by mass
- For example, a 10.0%(m/m) aqueous NaCl solution contains 10.0 g of NaCl per 90.0 g of water.

Example

15.00 grams of sodium hydroxide is added to 185.00 grams of water.

a. What is the mass percent of NaOH in this solution?

b. Suppose that the volume of the solution was 196.2 mL. What is the density of the solution?

c. What volume of this solution contains 50.00 grams of NaOH?

Molarity

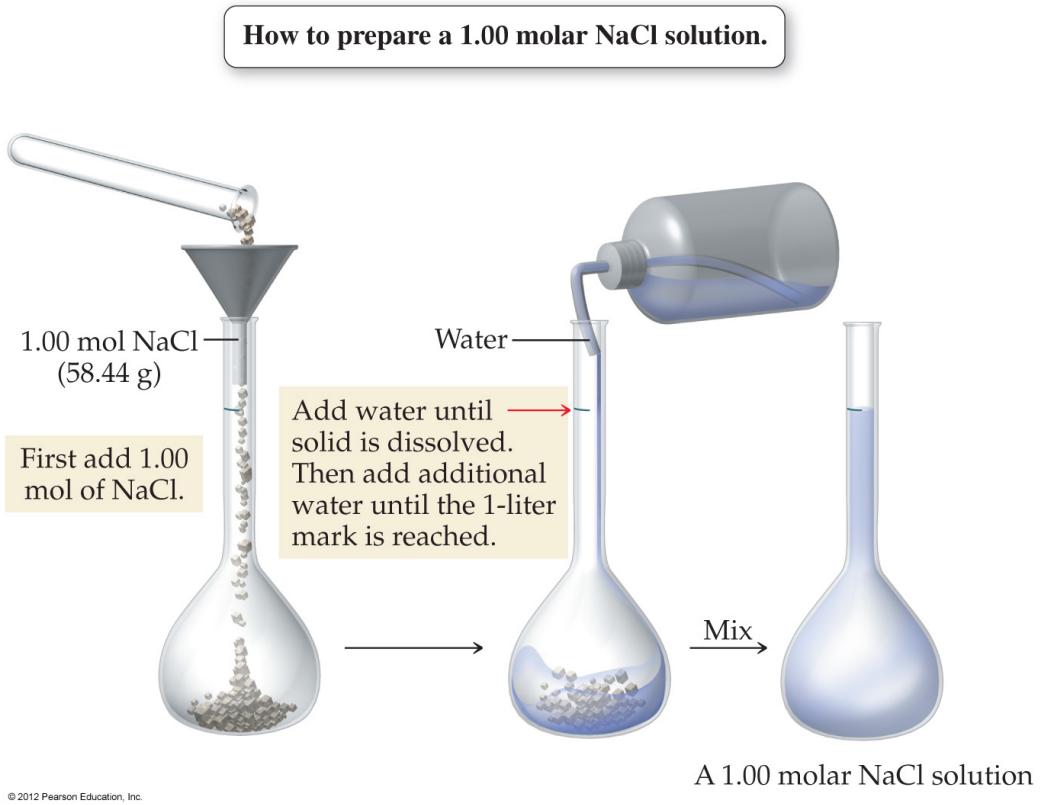
- The most useful term for concentration relates the number of moles of solute in a liter of solution; it is called the molarity, and abbreviated "M":

$$\text{molarity} = \frac{\text{moles of solute}}{\text{Liters of solution}}$$

- We often use square brackets around a compound or ion to indicate molarity
  - For example $[Na^+]$ means "molarity of sodium ion"

Concentration and Molarity

- For example, to prepare a 1.0 M NaCl solution, we take 1.0 mol of NaCl (58.44 g) and add enough water to it to bring the total volume of the solution to 1.0 L.
- We cannot simply say "add 1.0 L of water," because the presence of the salt will effect the final volume of the solution.



Example

Explain how to prepare 2.00 L of a 0.550 M solution of potassium chloride.

Mole-Volume Problems

- Since the molarity of a solution relates the volume to the number of moles of a solute, we can use it to help us calculate the amount of product formed in a reaction.
- First, we must convert the volume of a solution to moles of solute, using the molarity.
- From there, we can use the mole ratios of the chemical equation to figure out how much product is produced.

Examples

25.00 mL of a 0.45 M $AgNO_3$ solution is treated with excess NaCl solution. What solid precipitate is formed, and what should its mass be in grams?

Examples

What volume of 0.224 M $Ba(OH)_2$ solution is required to completely neutralize 15.00 mL of a 0.855 M HNO_3 solution?

Examples

25.00 mL of a 0.425 M potassium iodide solution is added to about 90 mL of water. To this is added 25.00 mL of a 0.724 M lead(II) nitrate solution. What precipitate is formed, and what is its mass?

Millimoles

- Since we often carry out volume measurements in the lab in milliliters, it is often convenient to use <u>millimoles (mmol)</u> rather than moles in calculations
- A millimole is defined in the same way other metric units are
  - 1 mmol = 10^{-3} mol, or equivalently
  - 1000 mmol = 1 mol
- We can define molarity as

$$molarity(M) = \frac{mol\ solute}{1\ L\ solution} = \frac{mmol\ solute}{1\ mL\ solution}$$

and use the form which is most convenient

Examples

Let's redo an earlier problem, using millimoles instead of moles in the calculation:

What volume of 0.224 M $Ba(OH)_2$ solution is required to completely neutralize 15.00 mL of a 0.855 M HNO_3 solution?

Ion Concentration

- For many chemical calculations, it is essential that we know the concentration of *each ion* in the solution
- For example, consider a 1.00 M $CuCl_2$ solution
- What is the molarity of Cu^{2+} in this solution?
- What is the molarity of Cl^-?

Example

10.00 mL of 2.45 M NaCl solution is combined with 10.00 mL of 1.50 M $BaCl_2$ solution. What is the concentration of each ion in solution after the two have been combined?

Example

How could we prepare 500. mL of a solution with $[Ni^{2+}] = 0.100$ M? We will use nickel (II) nitrate as the source of the nickel ion. (Why is it not possible to add only Ni^{2+} ion to the solution without adding any other ion?)

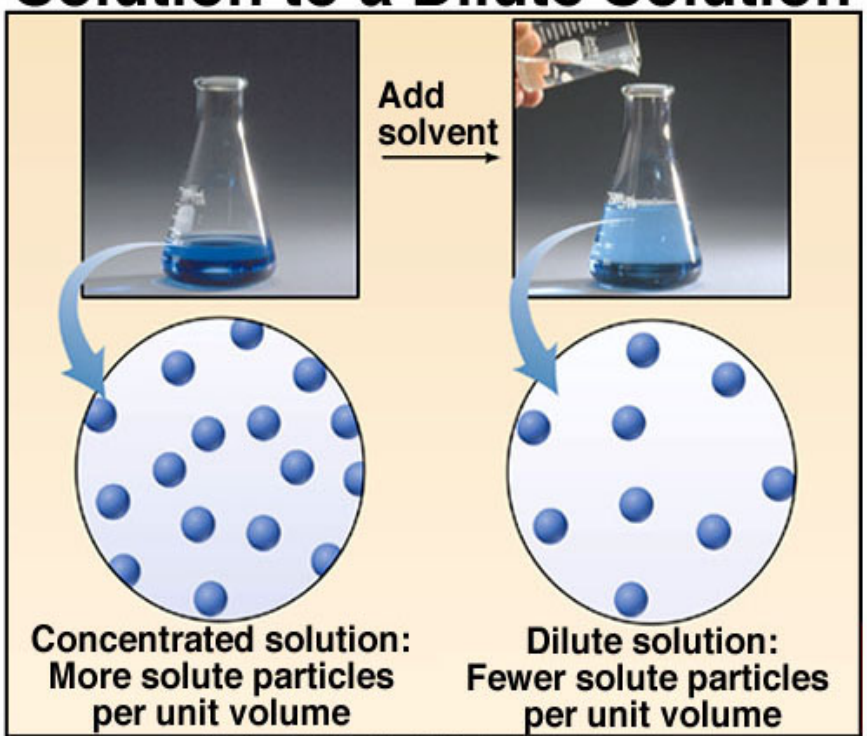
Dilutions

- A common laboratory procedure is the dilution of a concentrated solution to one of a lesser molarity.
- The necessary calculations for this can be accomplished either by using dimensional analysis, or by using the relationship

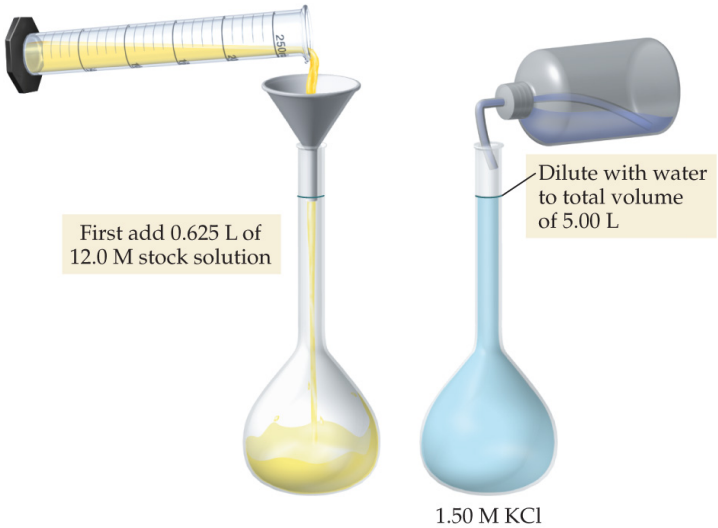
$$C_1V_1 = C_2V_2$$

where C stands for concentration, and V stands for the volume of the solution.

Converting a Concentrated Solution to a Dilute Solution



How to make 5.00 L of a 1.50 M KCl solution from a 12.0 M stock solution.



$$M_1V_1 = M_2V_2$$

$$\frac{12.0\ \text{mol}}{\cancel{\text{L}}} \times 0.625\cancel{\text{L}} = \frac{1.50\ \text{mol}}{\cancel{\text{L}}} \times 5.00\cancel{\text{L}}$$

$$7.50\ \text{mol} = 7.50\ \text{mol}$$

Example

- A chemical stockroom has a 6.00 M NaOH solution on hand (called a "stock solution"), which a chemist must use to prepare 10.0 L of 0.15 M solution. What volume of the stock solution is required?

Titrations

- A titration is usually carried out to determine the molarity of an unknown solution.
- A specific amount of a compound whose concentration is well known, called the primary standard is measured out.
- An indicator, a compound with a different color under different chemical environments, is added to the primary standard.
- The unknown solution is added until the indicator changes color, signaling the end of the titration (called the end point).
- From the collected data it is possible to determine the molarity of the unknown solution.

Titrations

- At the end point, we assume that there are "equivalent" amounts of each chemical present
  - This is *not* precisely true, but this definition is adequate for now
- By equivalent, we mean that the ratio of the moles of each reactant added equals the ratio in the balanced equation

Titrations

Examples

$HCl + NaOH \rightarrow H_2O + NaCl$

Ratio of acid to base: 1 to 1, so we need 1 equivalent of acid for every 1 equivalent of base

$H_2SO_4 + 2\ NaOH \rightarrow 2\ H_2O + Na_2SO_4$

Ratio of acid to base: 1 to 2, so we need1 equivalent of acid for every 2 equivalents of base

Example

10.00 mL of an HCl solution is pipetted into a beaker containing about 25 mL of deionized water. This solution is then titrated with 0.4575 M NaOH. It is found that 17.85 mL of base is required to reach the end point. What is the molarity of the acid solution?

Example

4.2554 grams of solid oxalic acid ($H_2C_2O_4$) is dissolved in about 50 mL of deionized water. It is found that 37.55 mL of a KOH solution is required to titrate this acid to the end point. What is the concentration of the base solution?

Optional Material

The remaining material is covered in the textbook but is not included in the students' notes. It should generally not be covered in Chemistry 120.

Molality

- The <u>molality</u> (symbolized by m) relates the amount of solute (in moles) to the *mass* of the solvent (not the total solution):

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

- So, an aqueous 5.0 m NaCl solution would contain 5.0 moles of NaCl per kilogram of water.

Example

45.8 g sodium nitrate is combined with 54.9 g distilled water

a. What is the mass percent of the resulting solution?

b. What is the molality of the resulting solution?

Example

Consider an aqueous solution which contains 35.5% KCl by mass. What is the molality of this solution?

Example

What is the molality of a 0.250 M NaCl solution, assuming the density of the solution is 1.034 g/mL?

Mole Fraction

- The mole fraction of a compound in a mixture is the number of moles of that compound divided by the total number of moles of all components in the mixture.
- We use the letter X (or the Greek letter χ, called "chi") to designate mole fraction
- For example suppose we combine 15.0 grams of NaCl with 200. grams of water:
  - 15.0 g NaCl = 0.257 moles NaCl
  - 200. g H_2O = 11.1 moles H_2O

$$\chi_{NaCl} = \frac{0.257}{(0.257 + 11.1)} = 0.0225$$

Example

What is the mole fraction of sodium chloride for a solution which is 1.15%(m/m) NaCl?

What is the molality of an aqueous solution of NaCl with $\chi_{NaCl} = 0.050$?

Colligative Properties

- Colligative properties describe how the presence of a solute effects the solvent.
- The specific colligative properties we will consider are
  - Lowering of the solvent's vapor pressure
  - Boiling point elevation
  - Freezing point depression
  - Change in osmotic pressure
- You will not be responsible for calculations based on colligative properties in this course.

The Effect of Solute Particles

- When solute particles are added to a solvent, they "get in the way" of the solvent's molecules.
- The vapor pressure of a solvent may be lowered when a solute is added, as the solute particles may block the surface of the solution, making it more difficult for the solvent particles to evaporate.
- For the same reason, the boiling point is raised, as more energy is required for solvent molecules to break through the surface.

Freezing Point Depression

- The freezing point of a solution is lower than that of the pure solvent.
- As the solvent molecules in a solution try to form a solid in the freezing process, the solute particles may "get in the way."
- Extra cooling is required to solidify the solution.

Osmotic Pressure

- When two solutions of unequal concentration are separated by a semi-permeable membrane (one that only allows water to flow through it), solvent will flow from the less concentrated solution, through the membrane, into the more concentrated solution to equalize the two. This process is called osmosis.
- The pressure that would have to be applied to prevent this from occurring is called the osmotic pressure.
- In general, the greater the difference in the concentration of the solutions, the greater the osmotic pressure.