> *"It is clear that under these circumstances the classical theory can not be retained. All experimental material indicates that its fundamental starting point should be abandoned, and that, in particular, an equilibrium calculated on the basis of the mass action law does not correspond to the actual phenomena."*
>
> *P. Debye and E. Hückel*

Solutions

To a chemist, a solution is nothing more than a homogeneous mixture. The defining phrase is comprised of two words, each of which has a very specific meaning in chemistry: Homogeneous, meaning that the sample has a uniform appearance and composition throughout, plus mixture, a sample that consists of two or more substances. If both of these definitions are met, a sample is a solution.

Solutions are frequently, both in the chemistry laboratory and in everyday life. In the laboratory, solutions are an excellent medium for promotion of chemical reactions and growing crystals. The particles are much closer together than in a gas, and they have more freedom of movement than in a solid. Outside of the laboratory, the process of life itself depends on solutions. The air we breathe and the oceans, lakes, and streams that cover most of our planet are examples of solutions.

Solution Terminology

Solutions are not limited to ionic and molecular solids being dissolved in water as in the workshop *"Ions in Solution"*.

Terms, which relate to solutions, need to be expanded. In general, the component of the solution that is present in the greatest amount is called the *solvent*. The substance with the smaller amount in the solution is called the *solute*. These terms are not precise, however, and their usage varies among different specialties in chemistry. For example in water solutions, water is almost always referred to as the solvent no matter its relative amount. Also when a solid or a gas is dissolved in a liquid, the liquid is generally called the solvent.

It is often important to know the maximum amount of a solute that will dissolve in a given solvent at a specified temperature. This measure is known as the *solubility* of that solute. Reference sources often report solubilities in grams of solute per 100 grams of solvent. When a solution contains a solute amount less than the solubility limit, it is said to be *unsaturated*; if it is at the solubility limit, it is *saturated*. Under certain special conditions, a solution can contain more solute than it normal solubility limit, and in this case, it is called *supersaturated*.

The terms concentrated and dilute are often used to describe solutions. It is important to keep in mind that these terms are valid only in a relative sense. A *concentrated* solution has a relatively large amount of solute per given amount of solvent when compared with a *dilute* solution. The comparison is only valid for systems of the same solute and solvent.

When discussing solutions of liquids in liquids, the term *miscible* is used to describe two liquids that will dissolve in one another in all possible combinations. When liquids will not dissolve in one another, they are said to be *immiscible*. A chemist would say that oil and water are immiscible, whereas alcohol and water are miscible.

Solution Concentration Units

A number of different units are used to express the quantity of solute dissolved in a given amount of solvent. Molarity was introduced in the workshop *"Ions in Solution"*.

$$\text{molarity} = \text{M} = \frac{\text{moles solute}}{\text{liters solution}}$$

Some other common units include percentage by mass, percentage by volume, (which relates to alcoholic proof), parts per million, parts per billion, and molality. The definition of each provides the basis for calculations with that unit.

$$\%\text{ by mass} = \frac{\text{mass solute}}{\text{mass solution}} \times 100$$

$$\%\text{ by volume} = \frac{\text{volume solute}}{\text{volume solution}} \times 100$$

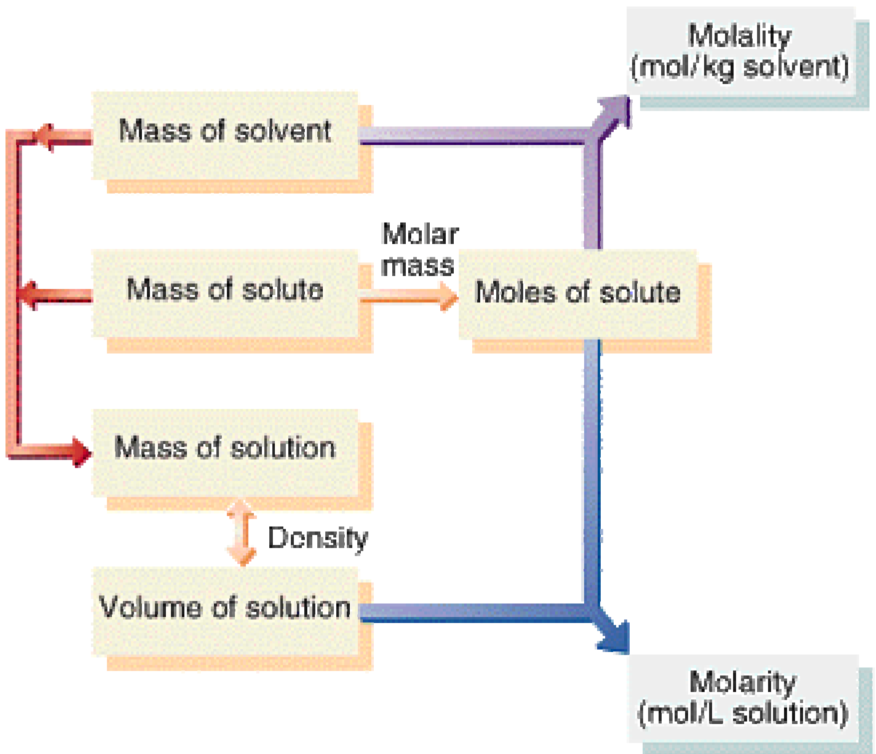
[Proof = % by volume x 2]

$$\text{parts per million} = \text{ppm} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^6$$

$$\text{parts per billion} = \text{ppb} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^9$$

$$\text{molality} = \text{m} = \frac{\text{moles solute}}{\text{kilograms solvent}}$$

The choice of concentration unit is largely a matter of application and convenience; other units beyond these may be encountered. There are some technical factors that must be considered. Percentage by mass, parts per million, parts per billion, and molality are applicable at any temperature. However, molarity and percent by volume must specify a temperature, since the volume varies with temperature.


Molality
(mol/kg solvent)
Mass of solvent
Molar
mass
Mass of solute
Moles of solute
Mass of solution
Density
Volume of solution
Molarity
(mol/L solution)

Solution Formation

Consider the process by which table salt, sodium chloride, dissolves in water. Positively-charged sodium ions and the negatively-charged chloride ions are attracted to one another, this attraction is termed to be an *ionic bond*. When a solution forms, the ions must be separated from each other, and thus energy is required to separate the solute particles. A similar situation occurs with the solvent particles. The water molecules are clumped together in hydrogen-bonded groups. These groups must be “pulled apart” so that solute ions can fit between them. This process also requires an input of energy. Energy is released when the solute and solvent particles interact. The negative ends of water molecules surround positively-charged sodium ions and the positive ends of water molecules surround negatively-charged chloride ions and pull apart the respective ions.

http://207.10.97.102/chemzone/lessons/03bonding/dissociate.htm

Crystal Formation (Crystallization)

If the solute and solvent do not chemically react, solution formation can be reversed. It is possible to recover the solute by removing all or some of the solvent. This can be particularly useful in purifying solids through "recrystallization", where the solid has contaminants that are more soluble in the solvent than the wanted solid.

Another useful application is in growing "single" crystals from saturated and super-saturated solutions. Single crystals are the exact geometrical shape as the salt's unit cell. There are seven different types of unit cells found in nature.

http://webphysics.davidson.edu/alumni/MiLee/JLab/Crystallography_WWW/intro.htm

Name(s): ________________________________

Workshop: Solutions

1. Solution A is made by dissolving 0.250 g of an ionic salt (X) in 1.00 L of water, and solution B is made by dissolving 2.50 g of a different ionic salt (Y) in 0.100 L of water. Which solution is dilute and which is concentrated? What is the solute and solvent in each solution? Which solution is saturated which is unsaturated? Can all of these questions be answered? If not, why not?

2. A solution is made by dissolving 1.0 g of potassium dihydrogen phosphate in 1.0 L of water. Assume that the volume of the resulting solution is 1.0 L and that the density of water and the resulting solution is 1.0 g/mL. Determine the concentration of the solution as: a) percentage by mass, b) parts per million, c) parts per billion and d) molality.

a)

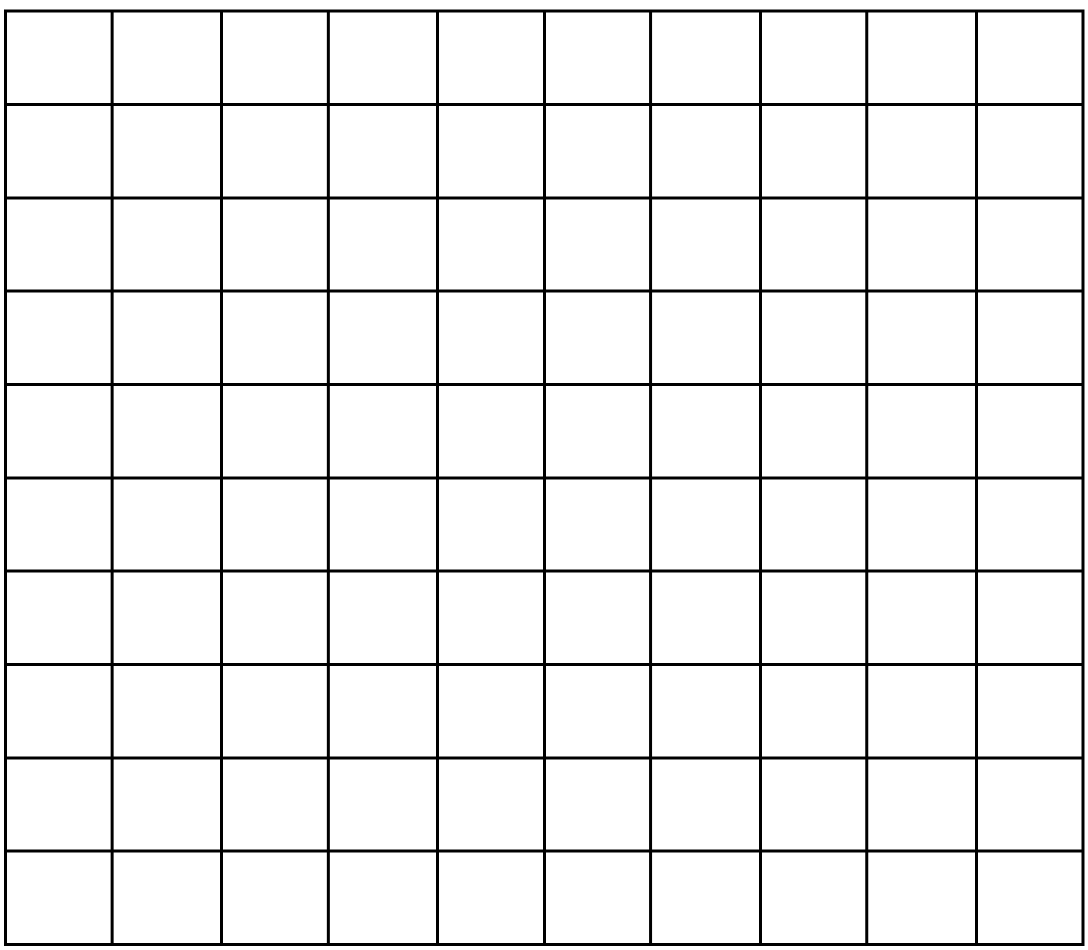
b)

c)

d)

3. Construct a solubility curve for copper (II) sulfate from the experimental data provided. Draw, title and label the curve in the blank graph that follows the datatable.

Water temperature (° C)	$CuSO_4$ dissolved (grams/ 100g H_2O)
0	23.1
10	27.5
20	32
30	37.8
40	44.6
50	53.2
60	61.8
70	72.8
80	83.8
90	98.9
100	114



4. From your graph in question 3, determine how much $CuSO_4$ will be in solution at 25 ^oC if 100g were dissolved in 150 g of water at 70 ^oC.

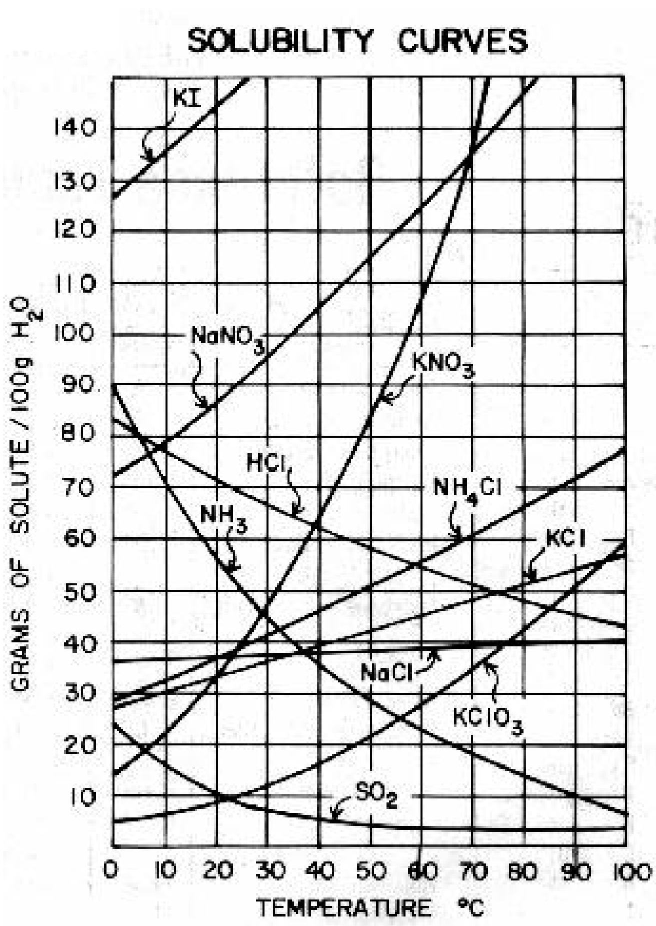
Using the following solubility curves answer questions # 5- 7

5. At what temperature is the solubility of ammonium chloride equal to ammonia?

6. If a power plant's smokestack gas were passed through water, would more SO_2 be removed at room temperature or near the boiling point of water?

7. If 20 g of sodium chloride were dissolved in 50 g of water at 70 ^oC and the solution then cooled to room temperature, how much solid sodium chloride should appear?

Pressure = 1 atm



8. A solid is dissolved in a liquid. How can they be separated? Brainstorm to come up with a list of possibilities, and then discuss each item on the list until the group arrives at least two feasible methods to separate liquids and solids.

9. How can you speed up the process of dissolving a specified amount of solute in a given quantity of solvent? Brainstorm to come up with a list of possibilities, and then narrow your list to three methods. For each of the three methods, analyze what happens at the particulate level to make it effective in speeding up the dissolving process. List the three methods, briefly describe your logic for each and explain if a solubility curve is important to your answer.

   I.

   II.

   III.

10. The solubility of magnesium hydroxide in water is 0.0009 g/mL at 18°C. What is the minimum amount of solution necessary–how much water will you add?–to dissolve 1.0 g $Mg(OH)_2$? What mass of magnesium hydroxide will dissolve in 1.0 L of water?

Questions 11 through 14: Use a group round-robin method to answer each question.

11. You need to make 1.00 kg of a 22.5% ammonium chloride solution. How many grams of ammonium chloride do you use? What volume of water will be used?

12. A sample of fruit juice is analyzed for its lead concentration. If a sample removed for analysis is found to have 0.477 ppm lead, how many grams of lead would be in a 8.0-fl oz serving of the juice?

13. Determine the molarity of a solution that is 88.2 ppm in urea, $CO(NH_2)_2$.

14. Describe how to prepare 5.00 L of a 0.75 molal solution of table sugar, $C_{12}H_{22}O_{11}$.

15. Each mole of particles added to one kilogram of water lowers the freezing point by 1.9°C. Determine the freezing point of solutions made by adding the following quantities to one kilogram of water:

    a. 16.0 g methanol, CH_3OH

    b. 58.44 g sodium chloride, NaCl

    c. 166.5 g calcium chloride, $CaCl_2$

    If each substance above had the same cost per kilogram, which one would be the best choice for melting ice in the winter? Explain.