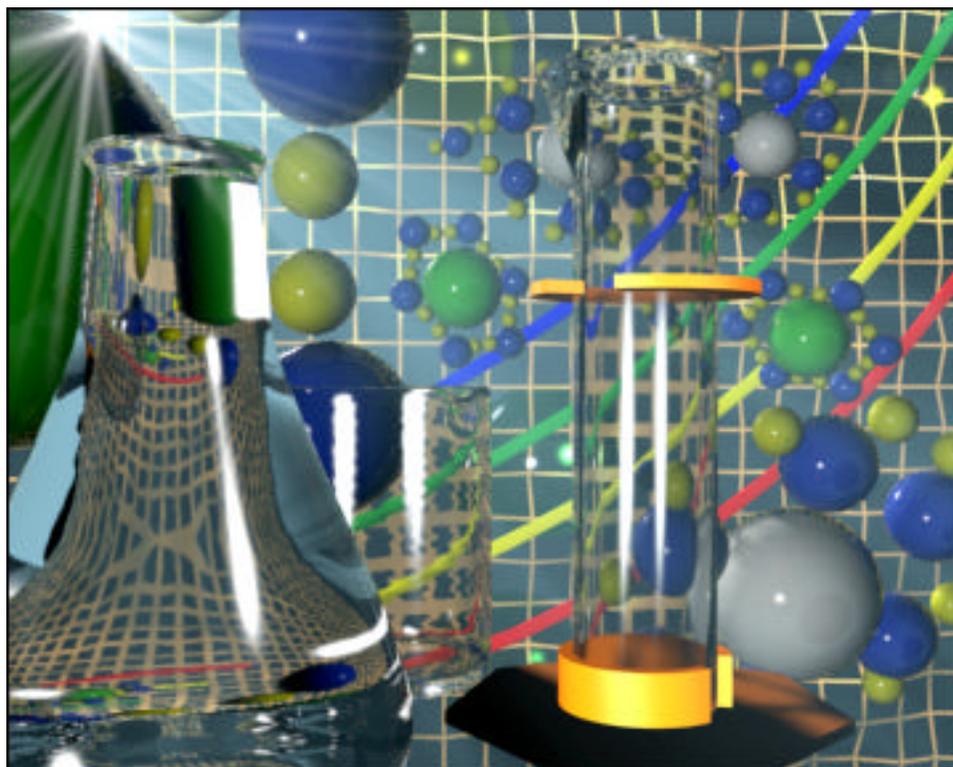


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Solutions
Program Supplement



Solutions Table of Contents

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Scenes 1 – 9
Introduction to Solutions

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Scene 1.

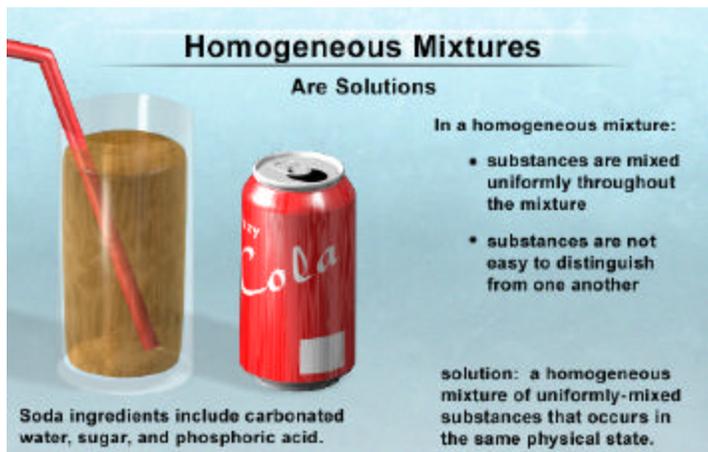
What do the sound you hear when you open a can of soda and the odor you smell when you open a bottle of cleaner have in common? Although you may not realize it, your senses detect that opening these containers causes chemicals to come out of **solution**. In this program, you will learn about solutions and how they form. You will become familiar with various ways to calculate solution concentrations. In addition, you will learn what colligative properties of solutions are and how solutes influence them.

Scene 2.

Two or more combined **substances** that do not react form a **mixture**. The word substance is a generic term describing **matter** that has definite composition, such as an **element** or a **compound**. Mixtures can be either heterogeneous or homogeneous. In a **heterogeneous mixture**, substances do not occur uniformly throughout the mixture. Therefore, substances in heterogeneous mixtures can in essence be distinguished from one another. A chocolate chip cookie can be considered a heterogeneous mixture because some substances, in this case chocolate chips, are not evenly distributed throughout the cookie. For example, you may have one chocolate chip in your first bite of cookie, but none in the second bite.

Scene 3

Unlike heterogeneous mixtures, the substances in **homogeneous mixtures** occur uniformly throughout the mixture. If you take several samples of a solution, you will find that the proportions of the substances within each sample are equal. In addition, the substances are not readily distinguished from one another. In fact, the substances in solutions remain together when run through a filter. Soda is an example of a homogeneous mixture because it contains several ingredients, including carbonated water, sugar, and phosphoric acid mixed uniformly throughout the soda. Your taste buds attest to the uniform distribution of this solution, since the first gulp of soda tastes the same as the last. Soda is also an example of a **solution**. Any homogeneous mixture is a solution. Heterogeneous mixtures cannot be solutions since the substances in them are not uniformly mixed.



Scene 4

For solutions to form, one or more substances must dissolve in a medium, which is often a liquid. Using Kool-Aid as an example of a solution, you see that the Kool-Aid powder dissolves in water when stirred. Since the powder dissolves in water, the powder is called a solute. A **solute** is any substance that dissolves when making a solution. The medium in which a solute dissolves is called a **solvent**. Water is the solvent in this example. A solute that dissolves in a specific solvent is said to be **soluble** in that solvent. As you have just seen, Kool-Aid® powder is soluble in water. However, not all substances dissolve to form solutions in a given solvent. Compare what happens when you place a water soluble vitamin C tablet into a beaker of water and another into a beaker of vegetable oil. The vitamin C tablet begins dissolving shortly after being placed into the water. In contrast, the vitamin C tablet placed in the beaker of vegetable oil does not dissolve, even when prodded with a stirring rod. A solute that does not dissolve in a specific solvent is said to be **insoluble** in that solvent. In this example, vitamin C is insoluble in vegetable oil.

Scene 5

Solutions can be liquids, gases, or solids. Classification of a solution is based on the state of matter of the solution after the solute and the solvent mix. For instance, liquid solvents are capable of dissolving gases, solids, and other liquids to form liquid solutions. In fact, soda is a good example where the three states of matter are blended into one state. To produce carbonated water, which is the main ingredient in most sodas, carbon dioxide gas and water are bottled under great pressure, causing the gas to dissolve into the water. To sweeten the soda, solid sugar and liquid corn syrup are dissolved into the liquid soda solution. Since this solution and many other solutions contain water as the solvent, a term distinguishing these solutions from solutions where water is not the solvent is used. Liquid solutions that have water as the solvent, such as soda, are called **aqueous solutions**. Liquid solutions that contain a different solvent are simply called liquid solutions.

Scene 6

Not all liquids dissolve into one another to form a solution. You are probably quite familiar with the fact that oil and water do not mix. The rainbow colored beads or streaks you see on the pavement after a rainstorm result from leaked engine oil not dissolving into the water on the pavement. Liquids, such as these, that do not dissolve are said to be **immiscible**. Liquids that dissolve, such as acetic acid and water, which make the vinegar solution used to color Easter eggs, are **miscible**. Liquids capable of dissolving into one another in any amount are said to be completely miscible. For example, 50mL of acetic acid dissolve in 5mL of water as easily as 5mL of acetic acid dissolve in 50mL of water.

Scene 7

Many liquid solutions are common and familiar. However, solutions can also be solids or gases. Remember that the state of matter of a solution is determined once the solute and solvent are mixed. For example, 18-karat gold is a solid solution composed of 75 % gold and 25 % silver. In order to blend pure gold and pure silver into a solution, the metals must be melted, mixed, and then cooled. Once cooled, the blend of the two metals is a **solid solution**, in which the gold and silver atoms remain fixed in a uniform distribution.

Scene 8

Metals, such as silver and gold, are often combined into solid solutions to create products with properties different from the pure metals composing the solution. The resulting solution is called an **alloy**. For example, pure gold is a soft metal. Consequently jewelry made of pure gold bends, twists, and scratches easily. However, when gold is mixed with silver or nickel, the resulting alloy is much harder. Other common alloys are used in cromoly bike frames, stainless steel pots, sterling silver jewelry, and brass instruments.

Scene 9

Examples of gaseous solutions include the compressed air one breathes while scuba diving and the laughing gas administered during dental surgery. Substances in a gaseous solution are uniformly distributed and constantly in motion. Any homogeneous mixture of gases that do not react together forms a solution. As with all solutions, the properties of the solution can differ from the properties of the substances forming the solution.

Scenes 10 – 16
How Solutions Form and Factors Affecting Formation

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Scene 10

Now that you know what solutions are, take a look at how solutions form. This demonstration shows how a solid solute dissolves in a liquid solvent. Here you see a **crystal lattice** of the sodium chloride solute begin to dissolve when placed in water. Since water is a **polar** molecule, its oppositely charged ends can overcome attractions keeping the ions within the salt crystal together. The partial positive charge on the hydrogen end of several water molecules draws the negative chloride ion from the solute and the partial negative charge on the oxygen atom in several water molecules draws the positive sodium **ion** from the solute. As a result, the sodium and chloride ions are stripped from the crystal and enter the solution. As the ions enter the solution, **intermolecular** attractions between water molecules are broken, providing room for the ions within the solvent. Once the ions are in solution, intermolecular attractions form among the ions and the water molecules in a process called **solvation**. When water is the solvent, as in this example, the process of **solvation** is called **hydration**.

Scene 11

Here is a summary of the three processes that occur when a solid solute dissolves in solution. Although the processes are discussed and summarized in a particular order, they occur simultaneously when forming a solution. In one process, the solute breaks apart. If the solute is an **ionic compound**, such as sodium chloride, attractions among ions are overcome. If the solute is a **molecular substance**, such as the sugar glucose, attractions among the individual molecules are overcome. In a second process, attractions among molecules composing the solvent must be broken to provide space for the solute. In a third process, attractions form among the solvent particles and solute particles to form the solution.

Scene 12

In step one, energy is required to break the attractive forces holding the solute together. Energy is also required in step two to break the attractive forces within the solvent, which provides space for the solute. Processes requiring energy in the form of heat, such as those occurring in steps one and two, are called **endothermic** or energy absorbing processes. In contrast, the process in step three releases energy as attractions form between solvent and solute particles. Energy releasing processes are called **exothermic** processes. If the combined amount of energy required to break the attractions within the solute and within the solvent is greater than the energy released when new attractions form between solute and solvent, then the formation of the solution is endothermic. If the opposite situation occurs, where the energy required to break the attractions within the solute and within the solvent is less than the energy released when attractions form between solute and solvent, then solution formation is exothermic.

Scene 13

A solution formed through an exothermic reaction, which can be made easily in your lab, is shown here. The solute, calcium chloride, is added to 100 mL of room temperature water in a beaker. As calcium chloride dissolves you can feel heat coming from the solution, as the increasing temperature reading indicates. An endothermic reaction occurs when making an ammonium nitrate solution. Ammonium nitrate dissolves when placed into 100 mL of water. As it dissolves, the absorbed energy causes the beaker to feel cold to the touch. The decreasing temperature reading indicates the cooling of the solution. You may be more familiar with this endothermic process than you realize if you have ever used a cold pack to reduce the soreness and swelling of an injury.

Scene 14

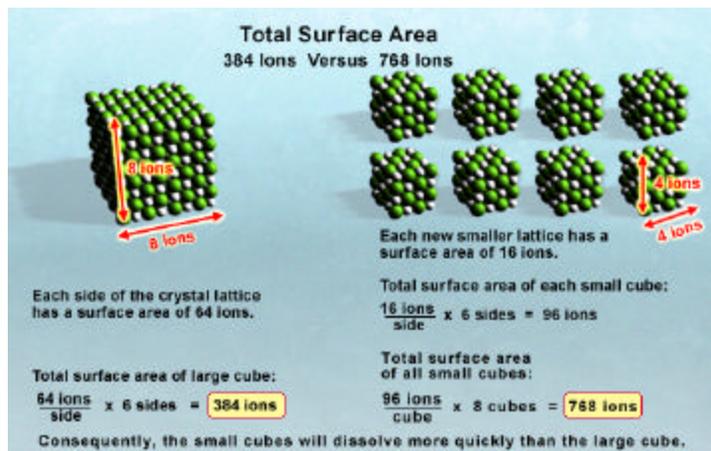
Imagine wanting to increase how quickly the ammonium nitrate mentioned in the last scene dissolves. How can you influence this? There are three factors you can use that affect how quickly a solid dissolves. These factors are temperature, stirring, and surface area. Temperature has a fairly consistent effect on the dissolving rate of a solid. When a solid solute is placed in a liquid solvent and subsequently heated, the increase in kinetic energy from the heat causes the solvent molecules to move quickly. As a result, the solvent molecules strike the solute's surface and generally strip the particles off the solute more rapidly at higher temperatures. Therefore, heating a solvent usually causes the solute to dissolve more quickly.

Scene 15

Another way to increase how quickly a solid dissolves is to stir the solvent. Stirring increases how fast a solid dissolves because stirring moves stripped solute particles away from the solid and into solution. If a solid dissolves without being stirred, the stripped solute tends to linger near the solid. This hinders solvent molecules from approaching the solid, therefore, dissolving is slow. Stirring moves dissolved solute away from the solid, allowing the solvent molecules to quickly strip more solute from the solid's surface.

Scene 16

You just saw that dissolving occurs when solvent particles strip solute particles from the surface of a solid. Increasing the surface area of a solute provides a larger surface upon which the solvent can act. As a result, the rate at which the solid dissolves increases. If you are given a salt block, how could you increase the surface area



of the block so it could dissolve more quickly? One answer is to break the salt block into many smaller pieces. Although the total amount of salt remains the same, the many smaller pieces of salt have a combined surface area larger than the single block. Consequently, the smaller pieces provide a larger total surface area on which the solvent can pull off solute at any time. The more pieces you break off the block, the greater the total surface area. Therefore, when you put the pieces of salt into the solvent, they dissolve more quickly than would a large block of salt.

Scenes 17 – 33
Solution Concentrations

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Scene 17

The amount of solute dissolved in an amount of solvent is referred to as a solution's concentration. Sometimes your senses can detect the concentration of solutions. If you consider cranberry juice too tart, then the juice is too concentrated for your preference. Three terms describe the relative amount of solute within a solution: unsaturated, saturated, and supersaturated.

Scene 18

A solution in which solute particles continue to dissolve in the solvent is said to be **unsaturated**. For example, three bags of sweetener poured into a glass of unsaturated sun tea dissolve. If you continue adding more sweetener, the sweetener eventually stops dissolving. At this point, the solution contains as much solute as the solvent can hold at a specific temperature, and the solution is **saturated**. Thus, addition of solute to a saturated solution does not cause more solute to dissolve. However, a solution can sometimes be **supersaturated**, meaning that the solution holds more solute than a saturated solution at a specific temperature. Supersaturated solutions occur only under manipulated conditions, when the temperature of the solvent is raised to allow more solute to dissolve. The solution is then slowly cooled so that solute crystals do not immediately form, as is shown in this example of sodium acetate. Supersaturated solutions are unstable. It takes just a small crystal or a tap on the container to make the excess solute come out of solution, which returns the solution to a saturated solution.

Scene 19

Solution concentrations usually need to be measured more precisely than in the relative terms just discussed. A patient admitted into a hospital for dehydration needs an I.V. to hydrate them. It is important to know the exact concentration of the saline solution in the I.V. because too high of a saline solution could further dehydrate the patient. Without knowing exact concentrations, how would you determine what is too high? There are a number of ways to determine solution concentration. One way is to determine the solution's molarity, which is designated by a capital M. **Molarity** is expressed as the number of moles of solute in one liter of solution. You may remember from your previous chemistry studies that the number of particles in a mole of anything, including various solutes, is 6.02×10^{23} .

Solution Concentrations

Molarity

Molarity is one way to measure solution concentration.

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

The number of particles in a mole of any kind is 6.02×10^{23} .

A mole of gummy bears is 6.02×10^{23} gummy bears.	A mole of salt grains is 6.02×10^{23} salt grains.	A mole of stars is 6.02×10^{23} stars.
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Scene 20

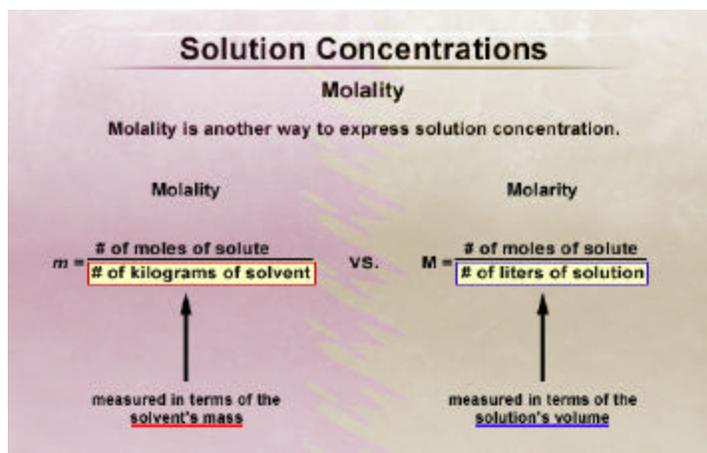
When determining the molarity of a solution, remember to express the answer in units of molarity -- moles of solute per liter of solution. How would you determine the molarity of a 100.00mL solution that contains 6.00 grams of potassium dichromate? Notice that this example begins with six grams of the solute per 100.00mL of solution instead of moles of solute per liter of solution. Therefore, you need to convert the six grams of potassium dichromate to moles. To convert grams to moles you must first determine the **molar mass** of the solute. Molar mass tells how much one mole of a substance weighs, and is expressed in grams per mole. Calculate the molar mass by adding the molar masses of all the atoms that comprise the solute. Each potassium atom has a mass of 39.10 grams per mole, each of the two chromium atoms has a molar mass of 51.99 grams per mole, and each of the seven oxygen atoms has a mass of 16.00 grams per mole, therefore the molar mass for potassium dichromate is 294.18 grams per mole. Dividing the 6.00 grams of potassium dichromate by its molar mass gives a total of 0.0204 moles of potassium dichromate.

Scene 21

The next step to determine molarity in this example is to convert milliliters to liters. The 100 milliliters of solution need to be converted to liters. Make the conversion using the conversion factor 1000 milliliters equal one liter. To be sure the equation is set up correctly, make sure the units cross cancel for everything but moles in the numerator and liters in the denominator. The calculation shows the molarity of this potassium dichromate solution is 0.204 moles of potassium dichromate per liter of solution. When making a solution measured in molarity, first place the solute into a volumetric flask, then add roughly 25% of the desired solvent to the flask. Swirl the flask until the solute dissolves, then add the solvent until the solution reaches the desired amount. In this example, 6.00 grams of potassium dichromate are placed in a volumetric flask, which is subsequently filled with 25.00mL of water. The contents are then swirled until the potassium dichromate is evenly distributed. More solvent is added, and the solution is swirled again. Finally, water is added to the 100.00mL mark, giving a concentration of 0.204 moles of potassium dichromate per liter of solution.

Scene 22

A second way to express solution concentration is **molality**, which is denoted by a lower case italicized *m*. Molality expresses the number of moles of solute per kilogram of solvent. Note that the units in the denominator position differ between molarity and molality. Molality measures the mass of solvent whereas molarity measures the volume of the entire solution.



Scene 23

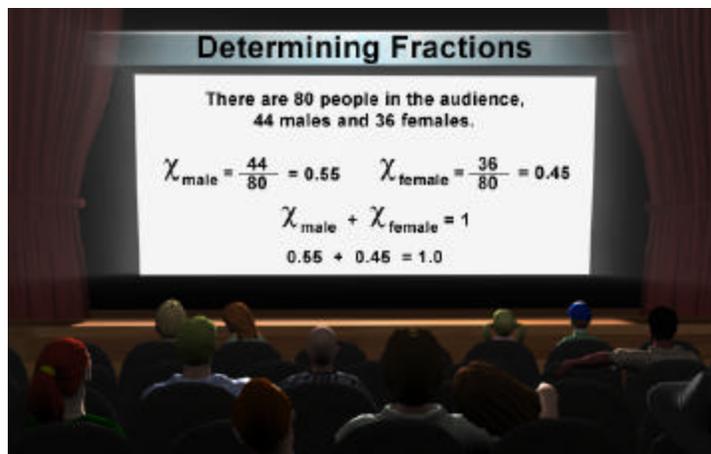
What is the molality of a solution containing 125.0 grams of sodium chloride in 2.00 kg of water? To determine the molality, first determine the molar mass of sodium chloride, which is 58.44 grams per mole. Converting grams to moles, you see that 125.00 grams of sodium chloride equal 2.139 moles of sodium chloride. Now set up the remainder of the problem so that your answer will be left in units of molality, or moles of solute per kilogram of solvent. This is simply done by placing the 2.00 kilograms of water in the denominator position. The molal concentration of this sodium chloride solution is 1.07 moles of sodium chloride per kilogram of water.

Scene 24

Mole fraction is a third way to express a solution's concentration, and it is represented by the Greek symbol chi. The mole fraction is the number of moles of one component, either the solute or the solvent, divided by the total moles of the solution. Total moles of the solution equal the number of moles of solute plus the number of moles of solvent. A subscript indicating whether the solute fraction or the solvent fraction is being solved for is placed after chi. For example, if the mole fraction of a solute is being determined, the equation is moles of solute divided by total moles of solution, and the equation is denoted as chi with the term solute written as a subscript. If the mole fraction of a solvent is being determined, the equation is moles of solvent divided by total moles of solution. This equation indicates what portion or fraction of the entire solution is solvent. To verify that your answer is correct, make sure the solvent mole fraction and the solute mole fraction total one when added. This ensures that each part of the solution was taken into account.

Scene 25

Before calculating mole fractions, let's use a familiar example of fractions to illustrate the concept. Imagine that you and your friends go to the movies. You notice that the audience is composed of a mixture of males and females. You count 80 people in the audience, 44 males and 36 females. The audience fraction of males is 44 out of 80 or 0.55 and the audience fraction of females is 36 out of 80 or 0.45. When the audience fractions are added, their sum equals one. Adding each component is a good way to determine if your work is correct.



Scene 26

Here is a sample of a problem using mole fractions. Find the mole fraction of the solute in a solution prepared by mixing 53.00 grams of sodium acetate into 100.00 mL of water. Determine the number of moles of sodium acetate and water. The composition of sodium acetate is $\text{NaC}_2\text{H}_3\text{O}_2$. Therefore, its molar mass is 82.04 grams per mole. Using the conversion factor 82.04 grams per mole, calculate the number of moles in 53.00 grams of sodium acetate. The number of moles is 0.6460. Next, convert 100 milliliters of water to moles of water. Remember that one liter of water equals one kilogram of water. Therefore the conversion factor equating one milliliter of water to one gram of water can be used because 1000mL equal one liter and 1000g equal one kilogram. Now the units are in terms of mass and not volume, but you still need to determine the number of moles of water. The composition of water is H_2O ; therefore, its molar mass is 18.02 grams per mole. Finishing the conversion of 100.00mL of water into moles of water, gives 5.549 moles. In other words, there are roughly five and a half moles of water in 100mL of water. Now that you know the number of moles of solute and solvent, you can determine the mole fractions of each in the next scene.

Scene 27

Adding the number of moles of water with the number of moles of sodium acetate determined in the previous scene gives the total number of moles of solution. This value, which becomes the denominator when determining mole fractions, equals 6.195. Now, find the mole fraction for sodium acetate. Simply plug the number of moles you just calculated for sodium acetate, 0.6460, into the equation for the mole fraction of the solute. You see that 0.6460 divided by 6.195 equals 0.1043, which is the mole fraction of sodium acetate. You can determine the mole fraction of the solvent at this point in one of two ways. One way to determine the mole fraction of water is by dividing the number of moles of water,

5.549, by the total moles of solution, 6.195. The mole fraction of water is 0.8957. To insure that your answer is correct, add both mole fractions together. If their sum equals one, then you have done the calculations correctly. The sum of 0.1043 and 0.8957 equals one; therefore you can be assured you solved this problem correctly. The other way to determine the mole fraction of water is to subtract the value of the solute's mole fraction from one. Subtracting 0.1043 from one gives a solvent mole fraction of 0.8957. Although the second method is quicker, there is no way to guarantee that your calculations are correct. Remember that the original question only asks for the mole fraction for the solute, which was calculated to be 0.1043.

Scene 28

Percent solution is the fourth and final way solution concentration will be presented in this program. However, you will learn two ways of calculating percent solution. The first, called percent solution by volume, is rather simple to calculate and is used when both the solute and solvent are liquids, such as methanol and water. Percent solution by volume measures the volume of a solute in relation to the volume of the entire solution. The fraction is converted to a percent by multiplying it by 100%.

Scene 29

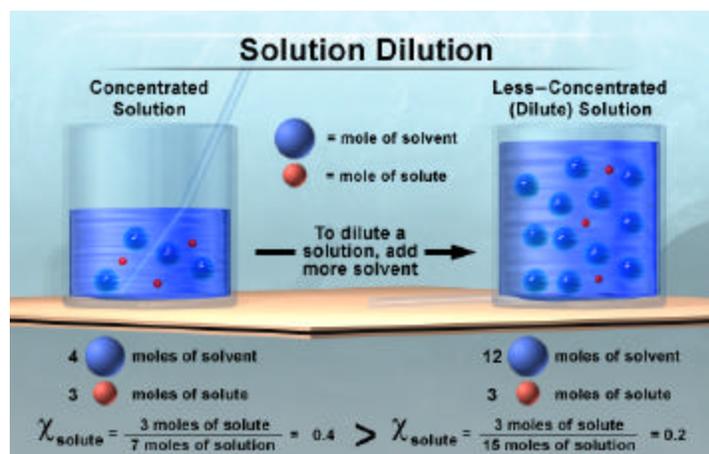
Suppose you add enough water to 25.00 milliliters of methanol to bring the final solution volume to 125.00mL. What is the percent methanol by volume? To answer this question, list the details that you know. One, the volume of solute is 25.00 milliliters. Two, the volume of solution is 125.00mL. The volume of solute divided by the total volume of solution multiplied by 100 % gives a value of 20 %. Since the units for volume are the same, they cancel. This solution contains 20% methanol. Remember to indicate that this measurement of percent solution is volume to volume, which is abbreviated as v over v.

Scene 30

When working with a solid solute in a liquid solvent, percent solution can be expressed in terms of mass of solute in grams over volume of solution in milliliters times 100%. It is often simple and convenient to weigh the quantity of solute needed to make a solution. For example, you are asked to make an aqueous sucrose solution that is 3 percent mass by volume. This indicates that 100.00 milliliters of solution contain 3.0 grams of sucrose. To make this solution, place the 3.00 grams of sucrose into a flask and add enough water to bring the volume to 100.00 milliliters. Whenever dealing with percent solutions, be sure to report the percent solution in the appropriate units. Notice that mass per volume is abbreviated as m over v.

Scene 31

Sometimes solutions are more concentrated than needed or desired. It is rather simple to dilute solutions. In fact, you have probably done this often. If cranberry juice is too tart, you add water to make it less concentrated. To dilute a solution, add more solvent. As the example shows, the amount of solvent in the dilute solution



increases while the amount of solute remains unchanged. Consequently, the solution becomes less concentrated. You can use your knowledge of solution concentrations, including mole fractions, to verify this concept. In the concentrated solution, there are 3 moles of solute per 7 moles of solution. The mole fraction of the solute is therefore 0.4. In the dilute solution, the additional moles of solvent bring the total moles of solution to 15. Therefore, the mole fraction of solute is 3 moles of solute divided by 15 moles of solution, which equals 0.2. You can see how the addition of solvent molecules changes the proportions of the solution.

Scene 32

Since the number of moles of solute does not change when making dilutions, a simple ratio can be used to determine how to prepare a dilute solution from a more concentrated solution. For example, suppose the all-purpose cleanser you use to clean the lab comes in a concentrated form called a stock solution. The label on the side of the bottle says to dilute the contents to a 0.60 molar solution before using. The label indicates that the cleaning solution is 3 molar. Therefore you know that there are 3 moles of solute per liter of solution. Since the moles of solute do not change when making dilutions, you can use the equation $(M_1)(V_1) = (M_2)(V_2)$ to relate the molarity and volume of a concentrated solution to the molarity and volume of a desired diluted solution. Recognize that canceling the units in the given equation leads to an equation stating that the number of moles in the concentrated solution equals the number of moles in the dilute solution.

Scene 33

The $(M_1)(V_1) = (M_2)(V_2)$ equation allows you to determine the volume of the concentrated solution required to make the dilute solution. You are asked to prepare 500.00 mL of a 0.60 molar solution from the 3.00 molar stock solution of lab cleanser. How do you prepare this solution? You know the desired volume of the dilute solution, which is designated as V_2 , is 500.00 mL, and the molarity of the dilute solution, which is designated as M_2 , is 0.60 moles per liter. M_1 equals the molarity of the concentrated solution, and in this example M_1 equals 3.0 moles per liter. Rearrange the equation to solve for V_1 , then substitute the numbers into their appropriate places in the equation. V_1 tells you the amount of the stock solution needed to make the dilution. The calculation gives you the answer 100.00 milliliters. Therefore, to make the desired solution, place 100.00 mL of the stock solution into a 500.00 mL container, and add enough water to bring the final volume to 500.00 mL. At this point, you have 500.00 mL of a 0.60 molar solution. If the volume of the solution must be precise, use a volumetric flask and pipette.

Scenes 34 – 48
Colligative Properties

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Scene 34.

Earlier in the program, you learned that properties of a solution can differ from the properties of the substances that comprise the solution. This was covered in the discussion of alloys. However, some properties of solutions depend on the number of solute particles in solution, and not the substances involved in the solution. The properties that depend on the number of solute particles are collectively called **colligative properties**. Colligative properties of solutions are **vapor pressure reduction**, boiling point elevation, freezing point depression, and osmotic pressure, each of which you will learn about in the next few scenes.

Scene 35.

The first colligative property you will investigate is vapor pressure reduction. Here you see a liquid solvent placed in a container, which is subsequently sealed. A short time later, some of the solvent molecules leave the liquid state, becoming gas molecules in a process called **evaporation**. Soon, some of these gas molecules return to the liquid state in a process called **condensation**. Eventually, a state of **equilibrium** occurs in which the rate of evaporation equals the rate of condensation. At equilibrium, there is no net change in the number of molecules in the gas and liquid phases although molecules continually evaporate and condense. The pressure of the gas molecules within the container at this time is known as **vapor pressure**.

Scene 36

Adding a nonvolatile solute, such as sugar or salt, to the solvent reduces the solvent's vapor pressure. The term **nonvolatile** means that the solute does not evaporate into a gas at the temperature of the solution. Although the nonvolatile solute mixes throughout the solution, it is the particles at the surface of the solution that have the greatest effect on vapor pressure. The solute particles at the surface displace some solvent particles, causing fewer solvent molecules to be located at the solution's surface. Consequently fewer solvent particles can evaporate since evaporation occurs at the surface of the solution. This in turn reduces vapor pressure in the container because fewer solvent molecules exist in the gas phase.

Scene 37.

The reduction in vapor pressure is directly proportional to the number of solute particles in the solvent once the solute dissolves. For example, adding the salt sodium chloride to a solvent reduces the vapor pressure twice as much as adding the same concentration of the sugar sucrose to a solvent. The reason the salt lowers the vapor pressure more than sugar is because sodium chloride separates into its two constituent ions, sodium ions and chloride ions, when it dissolves. Therefore, one mole of the solute sodium chloride yields two moles of particles, one of which is a mole of sodium ions and the other of which is a mole of chloride ions. In contrast, when sucrose dissolves, individual sucrose molecules separate from the solute. Therefore only one mole of sucrose particles

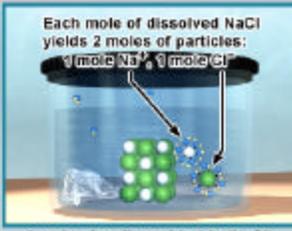
enters the solution for each mole of sucrose. Since two moles enter solution when sodium chloride dissolves, compared to one mole that enters solution when sucrose dissolves, the container with the sodium chloride solution has half the vapor pressure of the container with the sucrose solution. This is because two moles of ions from sodium chloride inhibit twice as many solvent particles at the solution's surface from evaporating as one mole of sucrose.

Vapor Pressure Reduction

The reduction in vapor pressure is directly proportional to the number of solute particles once the solute dissolves.

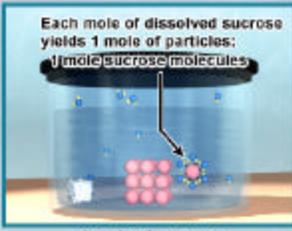
This container has 1/2 the vapor pressure reduction of the other because the two moles of particles inhibit twice as many solvent molecules from evaporating.

Each mole of dissolved NaCl yields 2 moles of particles: 1 mole Na^+ , 1 mole Cl^-



1 mole of sodium chloride (NaCl)

Each mole of dissolved sucrose yields 1 mole of particles: 1 mole sucrose molecules



1 mole of sucrose

Scene 38.

Another colligative property is boiling point elevation. **Boiling point** is defined as the temperature at which the vapor pressure of a liquid equals the pressure of the atmosphere. Water placed in a pot on the stove has a lower vapor pressure than the surrounding atmosphere; therefore the water does not boil. Heating the water causes water molecules to evaporate from the pot more quickly. Consequently, the vapor pressure of water increases until the water begins to boil. When the water boils, the vapor pressure within the pot equals the pressure of the atmosphere.

Scene 39

In previous scenes, you learned that the addition of a nonvolatile solute to a solvent decreases vapor pressure. The reduction of vapor pressure in turn affects boiling point because more heat is required in the presence of a nonvolatile solute to raise the vapor pressure enough to equal the pressure of the atmosphere. Consequently, a higher temperature is required to make the solvent boil, which can be seen in the example and on the graph showing the phase diagram of pure water. In this diagram, the solid blue line indicates the vapor pressure of pure water. The lower vapor pressure of an aqueous solution containing a nonvolatile solute is shown in red. Notice that the line representing the lower vapor pressure shows the solution has a higher boiling point than that of pure water.

Scene 40.

The change in boiling point temperature can be calculated using the equation ($\Delta T_b = K_b m$). ΔT_b is the number of degrees the boiling point increases. K_b is the molal boiling point elevation constant, which is expressed in terms of degrees Celsius per molal concentration. Each solvent has a specific K_b value regardless of the solutes added, and K_b values for various solvents can be found in many textbooks. In addition, m is the concentration of the solution measured in molality. Remember that molality measures a solution's concentration in terms of number of moles of solute per kilogram of solvent.

Scene 41.

Engine coolant, which is a solution of ethylene glycol and water, is an excellent example of a nonvolatile solute used to elevate the boiling point of water. Coolant is used in car radiators to prevent them from overheating on hot summer days. Let's calculate the boiling point for a solution containing 5.49 moles of ethylene glycol in 1400 grams of water. The first step is to determine the molality of the solution. Dividing 5.49 moles by 1400 grams gives 0.00392 moles of solute per gram of solvent. Since molality is expressed in terms of moles of solute per kilogram of solvent, this answer must be multiplied by the conversion factor 1000grams/ 1kg. The molality of this solution equals 3.92 moles per kilogram. Now you're ready to insert the known values into the boiling point elevation equation. According to the chart, the K_b value for water is 0.512°C per molal concentration. Multiplying 0.512°C per molal by 3.92 molal equals 2.01°C , which is the temperature difference between the boiling point of pure water and a 3.92 molal solution of ethylene glycol. Adding 2.01 degrees to water's boiling point, which at standard pressure is 100°C , gives an elevated boiling point of 102.01°C . Adding 5.49 moles of ethylene glycol to 1400 grams of water causes the solution to boil at 102.01°C .

Scene 42.

Freezing point depression is another colligative property. Freezing point is the temperature at which a liquid becomes a solid. During the freezing process, a substance changes from a loosely organized state to a highly organized state. As a substance changes from a liquid to a solid, intermolecular attractions between molecules become fixed. At zero degrees Celsius and one atmosphere of pressure, water freezes and turns to ice. The addition of a nonvolatile solute to a solvent inhibits the ability of the solvent to form an organized structure. Therefore, the solute must be moved out of the solvent particle's way in order for the solvent to freeze out of solution. The temperature of the solution must be lowered enough that the solvent molecules have a greater chance of forming attractions strong enough to hold the solvent particles together without the interference of the solute particles. When the appropriate temperature is reached, a solid made of only solvent particles forms.

Scene 43.

Engine coolant is not only used as a nonvolatile solute to prevent car radiators from boiling over. It is also called antifreeze, and as the name indicates, antifreeze prevents the water in a car's radiator from freezing during icy cold winters. The equation for freezing point depression is similar to the equation for boiling point elevation ($\Delta T_f = K_f m$). K_f is the molal freezing point depression constant, and you can find the K_f value for various solvents in many textbooks. Just as in the boiling point elevation equation, the concentration of the solution is measured in molality.

Scene 44

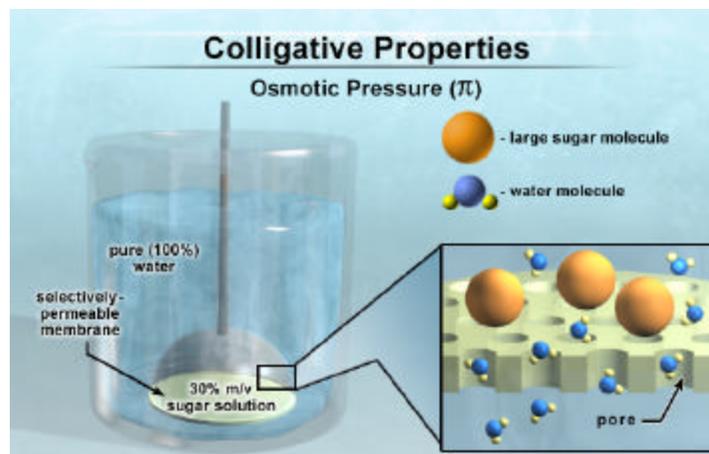
Calculate the freezing point of the solution containing 150.00 grams of ethylene glycol in 0.75 kilograms of water. Remember that you must first convert the 150 grams of antifreeze to moles in order to determine the solution's molality. The molecular formula of ethylene glycol is $C_2H_6O_2$, therefore its molar mass is 62.07 grams per mole. Dividing 150.00 grams by 62.07 grams per mole gives a value of 2.417 moles of ethylene glycol. The next step in solving this problem is to determine the concentration of this solution in units of molality. Because the amount of solvent is already given in kilograms, just divide the number of moles of ethylene glycol by the number of kilograms of water. Dividing 2.417 moles by 0.75 kilograms of water gives a molality of 3.2 moles of solute per kilogram of solvent. Now substitute the known values into the freezing point depression equation and multiply. From the chart, you determine the K_f value for water is 1.86°C per molal, and you just calculated the molality. The answer shows that 150.00 grams of antifreeze in 0.75 kilograms of water lowers the freezing point 6.0 degrees, to a temperature of -6.0°C .

Scene 45.

Adding a nonvolatile solute, like ethylene glycol, to a solvent increases the temperature range in which a solution remains a liquid. The temperature range increases because the nonvolatile solute elevates the boiling point and depresses the freezing point of a solvent. This relationship can be seen in this graph. The graph also shows the relationship between the amount of solute in a solution and the changes in the temperatures at which the solution boils and freezes. Since K_f and K_b values are constant for each solvent, the change in the temperature at which a solution boils and freezes is directly proportional to the molality of the solute. For example, doubling the amount of solute doubles the amount of change in freezing point depression and boiling point elevation, and a tripling of the amount of solute triples the amount of temperature change.

Scene 46

The last colligative property you'll examine is **osmotic pressure**, which is designated as π . This system will be used to illustrate the concept of osmosis and osmotic pressure. The beaker contains pure water and the tube contains a 30 percent mass by volume sugar-water solution. A



A selectively-permeable membrane is located at the bottom of the tube. This **selectively-permeable membrane** contains pores through which the water molecules can pass, but the sugar molecules, which are larger, cannot pass.

Scene 47.

Water tends to flow from an area of higher water concentration to an area of lower water concentration in a process called **osmosis**. Therefore in this example, water molecules will move across the selectively-permeable membrane into the sugar solution because the sugar solution contains a lower concentration of water, 70 percent water versus 100 percent water. Water continues to move across the selectively-permeable membrane into the sugar solution until equilibrium is reached. At equilibrium, water flows across the membrane in both directions at an equal rate. The movement of water into the tube is evidenced by the decrease in solvent height in the beaker and by the increase in solution height in the tube. At this point, the solution within the tube has a greater pressure exerted on it because the height of the solution is greater than the height of the solvent. The greater pressure on the solution, which is due to the pull of gravity, is called **osmotic pressure**. As with the other colligative properties, the amount of change due to the addition of the solute is directly proportional to the molality of the solution.

Scene 48.

If the osmotic pressure of a system is known, osmosis can be prevented by adding a pressure equal to the osmotic pressure on the solution. In fact, osmosis can even be reversed if the amount of pressure placed on the solution is greater than the osmotic pressure. The process of filtering water containing many dissolved solutes across a semi-permeable membrane is called **reverse osmosis**. Reverse osmosis filtration systems may be used for water treatment at home, in industry, or out at sea. The basic design of a reverse osmosis system places a thin selectively-permeable membrane between two compartments in a filter. Applying a pressure to the solution that is greater than the osmotic pressure forces water across the membrane into a different compartment, leaving the large solutes behind. The water in this compartment is purified and ready for use.

Scenes 49 – 52
Determining Molar Mass

	Scene Number
VII. Determining Molar Mass	49
A. Sample Problem	49
VIII. Conclusion	53

Scene 49.

Suppose your teacher gives you the molar masses of four substances and an unknown substance. During your lab, you must determine which of the four substances your teacher gave you. How can you determine which substance it is? You can use your knowledge of colligative properties to experimentally determine the molar mass of an unknown substance. There are a few things you must know and do before conducting the experiment. First, weigh the solute to determine the mass of unknown solute to be used in the experiment. Next, make sure the solvent chosen for the experiment has known K_b or K_f values that you can easily obtain. In addition, measure the quantity of solvent to be used in the experiment. Since the colligative property equations use molality as the concentration unit, it is easiest to measure the solvent in terms of molality—moles of solute per kilograms of solvent. Now perform an experiment to measure a colligative property that will allow you to determine the molar mass of the unknown substance.

Scene 50.

In this experiment 25.25 grams of the unknown solute is placed into 100.00 grams of water. Heating the solution reveals that the boiling point is 100.72 degrees Celsius. Subtracting the solution's elevated boiling point from the boiling point of pure water gives the change in boiling point. ΔT_b equals 0.72 degrees Celsius. Use the chart to determine the boiling point constant for water, $0.512^\circ\text{C}/\text{molal}$ concentration. Substitute the known values into the equation for molal boiling point then rearrange to solve for molality. Doing the calculation, the molality of this unknown substance is 1.4 moles per kilogram.

Scene 51

Now that molality is determined, calculating the molar mass is only steps away. The next step in solving for the molar mass is to break molality into its constituents, which are moles of solute divided by kilograms of solvent. Plug in the known values of this equation, which are molality and the mass of the solvent measured in kilograms, and you see that one variable remains. Rearrange the equation to solve for the number of moles of solute. At this point you probably recognize that the mass of solvent is still in grams. Convert the 100.00 grams of solvent to kilograms using the conversion factor 1 kilogram equals 1000 grams. Remember to show units and cross cancel them throughout the steps to ensure that you are on the right course. This becomes a little tricky at the end. Make sure you write the units of molality so you can see that kilograms cross cancel. Performing the calculation yields an answer of 0.14 moles of unknown solute.

Scene 52.

Remembering the components of molar mass helps derive the molar mass from the experimental data. The components of molar mass for the solute are mass of solute per mole of solute. In the previous scene, the number of moles of solute was determined to be 0.14, and the mass of solute was measured as 25.25 grams at the beginning of the experiment. Therefore both values for molar mass are known. Plugging in the values gives the answer that the molar mass equals 180.36 grams per mole. Comparing this value to the four values given by the teacher, you conclude that your unknown is substance X.

Scene 53.

This concludes the Solutions program. You have learned about various aspects of solutions, including how solutions form and the factors that affect the rate at which a solution forms. You now know there are several ways to express the concentration of a solution, and that reporting the units of concentration is important. In addition, you have seen the effects of a nonvolatile solute on the colligative properties of solutions. Now that you are familiar with the different types of solutions, you will begin recognizing how common solutions are in your daily life, both in and out of the chemistry lab.

Quiz 1
Solution Introduction and Formation

1. Which of the following is a solution?
 - A. coins in a jar
 - B. perfume in a bottle
 - C. paperclips in a drawer
 - D. mixed salad in a bowl
2. Liquid A mixes with liquid X, but not with liquid Z. What can be said about liquid A?
 - A. Liquid A is soluble in liquid Z.
 - B. Liquid A is miscible in liquid X.
 - C. Liquid A is immiscible in liquid Z.
 - D. Both B and C.
3. When two pure metals are blended, the resulting solution, or alloy, _____.
 - A. takes on the properties of the metal that has the greater atomic weight
 - B. takes on the properties of the metal that has the higher atomic number
 - C. has properties different from the metals that compose the solution
 - D. metals cannot make a solution because they are solids at room temperature
4. An aqueous solution is _____.
 - A. a heterogeneous solution
 - B. a type of liquid solution in which water is not the solvent
 - C. a type of solution where gas is dissolved in a solvent other than water
 - D. a type of liquid solution in which water is the solvent
5. Which of the following is not a solution?
 - A. tap water
 - B. a brass door knob
 - C. laughing gas anesthetic
 - D. a flower
6. Solutions are heterogeneous mixtures.
 - A. True
 - B. False
7. Which of the following is an example of an alloy?
 - A. pure gold
 - B. sterling silver
 - C. orange juice
 - D. carbon dioxide

8. The molecules in a gaseous solution are not evenly distributed.
- A. True
 - B. False
9. Sugar is often added to beverages to sweeten them. Sugar in this case is the _____ while the beverage is the _____.
- A. solute; solvent
 - B. solvent; solution
 - C. solvent; solute
 - D. solution; solvent
10. One way to distinguish between heterogeneous and homogeneous mixtures is to determine if _____.
- A. one mixture is solid and the other is liquid
 - B. one mixture is transparent and the other is opaque
 - C. one mixture has evenly distributed substances and the other does not
 - D. both A and C

Quiz 2

How Solutions Form and Factors That Affect Formation

1. Which of the following does not occur when a solid solute dissolves in a liquid solvent?
 - A. Attractions between the solute particles break.
 - B. Intermolecular attractions between the solvent particles break.
 - C. Attractions between the solute and solvent particles remain static.
 - D. Attractions between the solute and solvent particles form.
2. Forming attractions between solute and solvent particles is an _____.
 - A. endothermic process, which requires energy
 - B. endothermic process, which releases energy
 - C. exothermic process, which requires energy
 - D. exothermic process, which releases energy
3. In a particular reaction, more energy is required to break the attractions within the solute and solvent molecules than is released when the attractions between the solute and solvent particles form. Overall, this reaction is a(n) _____.
 - A. endothermic reaction
 - B. exothermic reaction
 - C. immiscible reaction
 - D. insoluble reaction
4. Three factors that affect the dissolving rate of a solid solute in a liquid solvent are _____.
 - A. stirring, pressure, and temperature
 - B. stirring, surface area, and pressure
 - C. stirring, surface area, and temperature
 - D. surface area, pressure, and temperature
5. Increasing temperature affects how quickly a solid solute dissolves because _____.
 - A. the particles move more quickly due to increased potential energy
 - B. the particles move around more quickly due to increased kinetic energy
 - C. solute particles are stripped from the solid more quickly
 - D. both B and C
6. Stirring an unsaturated solution causes solute particles to be drawn away from a solute and into the solution.
 - A. True
 - B. False

7. Which of the following provides the greatest surface area upon which a solvent can strip solute particles?
- A. one cubic cm of salt
 - B. one cubic cm of salt divided in half
 - C. one cubic cm of salt divided in thirds
 - D. one cubic cm of salt divided in quarters
8. Adding more solute to a saturated aqueous solution causes _____.
- A. an unsaturated solution because the solute will form a solid
 - B. a build up of undissolved solute
 - C. a supersaturated solution
 - D. both A and C
9. An equal amount of solute will dissolve most quickly at which temperature?
- A. 25 degrees Celsius
 - B. 15 degrees Celsius
 - C. 0 degrees Celsius
 - D. 32 degrees Celsius
10. A solute that is not soluble in a solvent at one temperature will always become soluble if you lower the temperature at which you try to dissolve it.
- A. True
 - B. False

Quiz 3

Solution Concentrations

1. A solution contains fifty-five grams of solute X dissolved in 100 milliliters of water. What can you determine about the solution?
 - A. Its molarity is 0.55.
 - B. The solution is saturated.
 - C. Solute X is soluble in water.
 - D. The solution is in equilibrium.

2. When determining the molarity of a solution, you must make sure that your answer is in terms of _____.
 - A. moles of solvent per liters of solvent
 - B. moles of solute per liter of solution
 - C. grams of solute per liter of solvent
 - D. moles of solvent per liter of solute

3. What is the molarity of an aqueous solution that has 8.0 moles of solute dissolved in 4.0 liters solution?
 - A. $M = 0.2$ moles/ liter
 - B. $M = 0.5$ moles/ liter
 - C. $M = 2.0$ moles/ liter
 - D. $M = 5.0$ moles/ liter

4. What is the molality of a solution that contains 0.50 moles of solute in one liter of water?
 - A. $m = 0.25$ moles / kg because one liter of water weighs two kilograms (0.50 moles / 2 kg).
 - B. $m = 0.50$ moles / kg because one liter of water weighs one kilogram (0.50 moles / 1 kg)
 - C. $m = 1.0$ mole / kg because one liter of water weighs two kilograms (0.50 moles X 2 kg)
 - D. cannot be determined because the molality units are in terms of kilograms of solvent and the information is given in terms of liters.

5. In determining a mole fraction, you calculate that the fraction of solute is 0.45. What is the fraction of solvent?
 - A. 0.10
 - B. 0.45
 - C. 0.55
 - D. 1.0

6. When diluting a solution, _____.
- the amount of solute increases, but the amount of solvent remains unchanged
 - the amount of solvent increases, but the amount of solute remains unchanged
 - the amounts of both solute and solvent are increased
 - the answer cannot be determined
7. For which variable do you solve to determine the molarity of the stock solution given the equation $M_1V_1 = M_2V_2$?
- V_1
 - M_1
 - V_2
 - M_2
8. You are asked to prepare 100 milliliters of a 0.50 molar solution from a stock solution that is 1.0 molar. Doing the math, you determine the volume of stock solution needed is 50 milliliters. How do you prepare the solution?
- Add 50 milliliters stock solution to 100 milliliters of water.
 - Add 50 milliliters of water to the bottle of stock solution.
 - Add enough water to 50 milliliters of stock solution to bring the final volume to 100.00 milliliters.
 - Cannot determine with this information.
9. A label on a bottle states it contains 53% sugar. What does this mean?
- There are 53 milliliters of sugar per every 100 milliliters of solution.
 - There are 53 grams of sugar per every 100 milliliters of solution.
 - There are no grams of fat in the solution.
 - Cannot be determined with this information because units are not given.
10. What is the percent solution, mass by volume, of 12.0 grams of solute in 100.0mL of solvent?
- 833 % because $\frac{100\text{mL of solvent}}{12.0 \text{ grams of solute}} (100\%) = 8.33 \%$
 - 8.33 % because $\frac{100\text{mL of solvent}}{12.0 \text{ grams of solute}} = 8.33 \%$
 - 12.0 % because $\frac{12.0 \text{ grams of solute}}{100\text{mL of solvent}} (100\%) = 12.0 \%$
 - 0.12% because $\frac{12.0 \text{ grams of solute}}{100\text{mL of solvent}} = 0.120 \%$

Quiz 4
Colligative Properties

1. Colligative properties of solutions depend on the number of solute particles in a solution.
 - A. True
 - B. False
2. Vapor pressure results from _____.
 - A. gases returning to a liquid causing the pressure of the gas to equal the pressure of the atmosphere
 - B. the pressure of a gas over its liquid at equilibrium
 - C. a liquid expanding as it solidifies, putting pressure on the surrounding container
 - D. a solid becoming a liquid
3. A reduction of vapor pressure occurs with the addition of a nonvolatile solute to a solvent because _____.
 - A. there are fewer solvent molecules at the surface that can evaporate
 - B. fewer solute molecules exist in the liquid
 - C. fewer solute molecules lie at the surface of the solution
 - D. fewer solvent molecules exist overall
4. You look outside and see that the thermometer reading is -1.00 degree Celsius.
You note the puddles are not frozen. Why not?
 - A. The osmotic pressure of the puddles is higher than the surrounding pressure, therefore ice does not form.
 - B. The presence of nonvolatile solutes in the puddles prevents ice formation.
 - C. The new thermometer has been dropped too many times to be accurate
 - D. Cannot determine with this information
5. Adding a nonvolatile solute to a solvent _____.
 - A. increases the boiling point of the solution
 - B. decreases the boiling point of the solution
 - C. increases the temperature range over which the solution remains a liquid
 - D. both A and C

6. Given the same solvent, which of the solutions would have the lowest freezing point?
- A. A 1 *m* solution containing a nonvolatile solute.
 - B. A 2 *m* solution containing a nonvolatile solute.
 - C. A 3 *m* solution containing a nonvolatile solute.
 - D. A 4 *m* solution containing a nonvolatile solute.
7. Doubling the amount of a nonvolatile solute in a solution _____.
A. doubles the elevation change of the boiling point
B. reduces the boiling point by one half
C. alters the amount of change in boiling point, but it cannot be determined by how much.
D. reduces the freezing point by one half
8. Osmosis is the movement of _____.
A. A solvent, which is often water, from a lower water concentration to a higher water concentration across a selectively-permeable membrane
B. A solvent, which is often water, from a higher concentration to a lower concentration across a selectively-permeable membrane
C. solutes from a lower water concentration to a higher water concentration across a selectively permeable-membrane
D. solutes from a higher water concentration to a lower water concentration across a selectively-permeable membrane
9. Osmotic pressure is caused by _____.
A. solvent molecules evaporating and becoming gas molecules
B. solute molecules evaporating and becoming gas molecules
C. solvent molecules moving across a selectively-permeable membrane into an area of lower solvent concentration
D. solutes molecules moving across a selectively-permeable membrane into an area of lower solvent concentration
10. Which of the following reduces vapor pressure the most?
A. a volatile solute
B. a molecular solute that dissolves to give off one mole of solute per liter solution
C. an ionic compound that dissolves to give off two moles of solute per liter solution
D. an element called tungsten

Quiz 5
Determining Molar Mass

- The molar mass of potassium chloride (KCl) is _____.
 - 70.55g/mol
 - 74.55g/mol
 - 78.25g/mol
 - 79.25g/mol
- When solving a chemistry problem, it is important to _____.
 - use the periodic table to find information you may need
 - make sure the units cross cancel so the final answer is in the units desired
 - determine which variables you know and which variables you do not know
 - all the above
- The information you need to determine the molar mass of an unknown solute is _____.
 - the amount of solute used in the experiment
 - the molal boiling or freezing constant value of the solvent
 - the amount of solvent used in the experiment
 - all the above
- Solvent X boils at 78.5 degrees Celsius and at 82.3 degrees Celsius when nonvolatile solute Y is added to it. What is ΔT_b ?
 - 3.8 degrees Celsius
 - 3.8 degrees Celsius
 - 21.5 degrees Celsius
 - 17.7 degrees Celsius
- Knowing that $\Delta T_b = K_b m$, you can determine the molality of unknown solute T once you determine the difference in boiling point. If $\Delta T_b = 5.0$ degrees Celsius, what is the molality of the unknown? $K_b = 0.512^\circ\text{C}/m$; $K_f = 1.86^\circ\text{C}/m$

A. $\frac{5.0^\circ\text{C}}{1.86^\circ\text{C}/m} = -2.69m$

B. $\frac{1.86^\circ\text{C}/m}{5.0^\circ\text{C}} = -3.7m$

C. $\frac{5.0^\circ\text{C}}{0.512^\circ\text{C}/m} = 9.8m$

D. $\frac{-0.512^\circ\text{C}/m}{5.0^\circ\text{C}} = 0.10m$

6. Once you determine the molality of unknown solute Q, you can determine the number of moles of that unknown using the units of molality, number of moles of solute per kilogram of solvent, and the amount of solvent used in the experiment. Calculate the number of moles of the unknown solute by using the unknown solute's recently calculated molality (3.4 m) when using 5.0 kilograms of solvent.

A. $5.0\text{ kg } (3.4\text{ m}) = 17\text{ moles}$

B. $\frac{5.0\text{ kg}}{3.4\text{ m}} = 1.5\text{ moles}$

C. $\frac{3.4\text{ m}}{5.0\text{ kg}} = 0.68\text{ moles}$

D. $\frac{3.4\text{ m}}{5.0\text{ kg}} = 0.68\text{ m/kg}$

7. In the final step of determining the molar mass of unknown solute A through experimentation, plug in the measured mass of the unknown and the number of moles calculated into the molar mass equation. If you weighed the mass as 20.05 grams and you calculated the number of moles to be 5.03, then the molar mass of the unknown is _____.

A. $\frac{5.03\text{ moles}}{20.05\text{ grams}} = 0.251\text{ grams/mol}$

B. $\frac{20.05\text{ grams}}{5.03\text{ moles}} = 3.99\text{ grams/mol}$

C. $\frac{5\text{ moles}}{20\text{ grams}} = 0.25\text{ grams/mol}$

D. cannot determine with just this information

8. You determined the molar mass for solute G is 95.2 g/mol. Which substance is your unknown? (Use a periodic table if necessary to determine the answer.)



9. Experiments using freezing point depression are more accurate than experiments using boiling point elevation.

A. True

B. False

10. Twenty moles of sodium chloride are placed in beaker A and 5 moles of sodium chloride are placed in beaker B. The solution in which beaker will show a greater increase in boiling point and decrease in freezing point?

A. Beaker A

B. Beaker B

C. Both A and B show an equal increase and decrease.

D. Neither A or B show a difference in boiling or freezing temperatures.

Multiple Choice Exam

1. A solution contains fifty-five grams of solute X dissolved in 100 milliliters of water. What can you determine about the solution?
 - A. Its molarity is 0.55.
 - B. The solution is saturated.
 - C. Solute X is soluble in water.
 - D. The solution is in equilibrium.
2. Which of the following is a solution?
 - A. coins in a jar
 - B. perfume in a bottle
 - C. paperclips in a drawer
 - D. mixed salad in a bowl
3. Liquid A mixes with liquid X, but not with liquid Z. What can be said about liquid A?
 - A. Liquid A is soluble in liquid Z.
 - B. Liquid A is miscible in liquid X.
 - C. Liquid A is immiscible in liquid Z.
 - D. Both B and C.
4. When two pure metals are blended, the resulting solution, or alloy, _____.
 - A. takes on the properties of the metal that has the greater atomic weight
 - B. takes on the properties of the metal that has the higher atomic number
 - C. has properties different from the metals that compose the solution
 - D. metals cannot make a solution because they are solids at room temperature
5. An aqueous solution is _____.
 - A. a heterogeneous solution
 - B. a type of liquid solution in which water is not the solvent
 - C. a type of solution where gas is dissolved in a solvent other than water
 - D. a type of liquid solution in which water is the solvent
6. Which of the following is not a solution?
 - A. tap water
 - B. a brass door knob
 - C. laughing gas anesthetic
 - D. a flower

7. Which of the following does not occur when a solid solute dissolves in a liquid solvent?

- A. Attractions between the solute particles break.
- B. Intermolecular attractions between the solvent particles break.
- C. Attractions between the solute and solvent particles remain static.
- D. Attractions between the solute and solvent particles form.

8. Forming attractions between solute and solvent particles is an _____.

- A. endothermic process, which requires energy
- B. endothermic process, which releases energy
- C. exothermic process, which requires energy
- D. exothermic process, which releases energy

9. In a particular reaction, more energy is required to break the attractions within the solute and solvent molecules than is released when the attractions between the solute and solvent particles form. Overall, this reaction is a(n) _____.

- A. endothermic reaction
- B. exothermic reaction
- C. immiscible reaction
- D. insoluble reaction

10. Three factors that affect the dissolving rate of a solid solute in a liquid solvent are _____.

- A. stirring, pressure, and temperature
- B. stirring, surface area, and pressure
- C. stirring, surface area, and temperature
- D. surface area, pressure, and temperature

11. Increasing temperature affects how quickly a solid solute dissolves because _____.

- A. the particles move more quickly due to increased potential energy
- B. the particles move around more quickly due to increased kinetic energy
- C. solute particles are stripped from the solid more quickly
- D. both B and C

12. Stirring an unsaturated solution causes solute particles to be drawn away from a solute and into the solution.

- A. True
- B. False

13. Which of the following provides the greatest surface area upon which a solvent can strip solute particles?
- A. one cubic cm of salt
 - B. one cubic cm of salt divided in half
 - C. one cubic cm of salt divided in thirds
 - D. one cubic cm of salt divided in quarters
14. Adding more solute to a saturated aqueous solution causes _____.
- A. an unsaturated solution because the solute will form a solid
 - B. a build up of undissolved solute
 - C. a supersaturated solution
 - D. both A and C
15. When determining the molarity of a solution, you must make sure that your answer is in terms of _____.
- A. moles of solvent per liters of solvent
 - B. moles of solute per liter of solution
 - C. grams of solute per liter of solvent
 - D. moles of solvent per liter of solute
16. What is the molarity of an aqueous solution that has 8.0 moles of solute dissolved in 4.0 liters solution?
- A. $M = 0.2$ moles/ liter
 - B. $M = 0.5$ moles/ liter
 - C. $M = 2.0$ moles/ liter
 - D. $M = 5.0$ moles/ liter
17. The molar mass of oxygen is approximately 16.0 grams/mole, and that of hydrogen is approximately 1.0 gram/mole. Knowing these values, you can determine the molar mass of hydrogen peroxide (H_2O_2), which is _____.
- A. 17.00 grams per mole because each elements' mass is counted only once when determining molar mass
 - B. 18.00 grams per mole because the lighter element's mass is counted twice
 - C. 32.00 grams per mole because only the heavier element's mass is counted
 - D. 34.00 grams per mole because the masses of all the atoms are counted

18. Carbon has a molar mass of 12.0 grams/mol. How many moles of carbon are in 24.0 grams of carbon?
- A. 0.50 moles because $12 \text{ grams/mole} / 24 \text{ grams} = 0.50 \text{ moles}$.
 - B. 2.0 moles because $24 \text{ grams} / 12 \text{ grams/mole} = 2.0 \text{ moles}$.
 - C. 3.0 moles because 0.50 must be multiplied by 6 since carbon is the sixth element.
 - D. 12.0 moles because 2.0 must be multiplied by 6 since carbon is the sixth element.
19. What is the molality of a solution that contains 0.50 moles of solute in one liter of water?
- A. $m = 0.25 \text{ moles} / \text{kg}$ because one liter of water weighs two kilograms (0.50 moles / 2 kg).
 - B. $m = 0.50 \text{ moles} / \text{kg}$ because one liter of water weighs one kilogram (0.50 moles / 1 kg)
 - C. $m = 1.0 \text{ mole} / \text{kg}$ because one liter of water weighs two kilograms (0.50 moles X 2 kg)
 - D. cannot be determined because the molality units are in terms of kilograms of solvent and the information is given in terms of liters.
20. In determining a mole fraction, you calculate that the fraction of solute is 0.45. What is the fraction of solvent?
- A. 0.10
 - B. 0.45
 - C. 0.55
 - D. 1.0
21. When diluting a solution, _____.
- A. the amount of solute increases, but the amount of solvent remains unchanged
 - B. the amount of solvent increases, but the amount of solute remains unchanged
 - C. the amounts of both solute and solvent are increased
 - D. the answer cannot be determined
22. For which variable do you solve to determine the molarity of the stock solution given the equation $M_1V_1 = M_2V_2$?
- A. V_1
 - B. M_1
 - C. V_2
 - D. M_2

23. You are asked to prepare 100 milliliters of a 0.50 molar solution from a stock solution that is 1.0 molar. Doing the math, you determine the volume of stock solution needed is 50 milliliters. How do you prepare the solution?
- A. Add 50 milliliters stock solution to 100 milliliters of water.
 - B. Add 50 milliliters of water to the bottle of stock solution.
 - C. Add enough water to 50 milliliters of stock solution to bring the final volume to 100.00 milliliters.
 - D. Cannot determine with this information.
24. Which of the following is not a colligative property?
- A. boiling point reduction
 - B. freezing point depression
 - C. osmotic pressure
 - D. vapor pressure reduction
25. Colligative properties of solutions depend on the number of solute particles in a solution.
- A. True
 - B. False
26. Vapor pressure results from _____.
- A. gases returning to a liquid causing the pressure of the gas to equal the pressure of the atmosphere
 - B. the pressure of a gas over its liquid at equilibrium
 - C. a liquid expanding as it solidifies, putting pressure on the surrounding container
 - D. a solid becoming a liquid
27. Which of the following is not an equilibrium event?
- A. The rate of evaporation equals the rate of condensation.
 - B. The rate of water flowing in opposite directions across a selectively permeable membrane is equal.
 - C. The forward and reverse reactions in a reversible reaction occurring at equal rates.
 - D. The boiling point of a solution is raised and the freezing point is lowered due to the addition of a nonvolatile solute into a pure solution.
28. Which of the colligative properties is not dependent on vapor pressure?
- A. boiling point elevation
 - B. vapor pressure reduction
 - C. osmotic pressure
 - D. A & B

29. A reduction of vapor pressure occurs with the addition of a nonvolatile solute to a solvent because _____.
- A. There are fewer solvent molecules at the surface that can evaporate
 - B. fewer solute molecules exist in the liquid
 - C. fewer solute molecules lie at the surface of the solution
 - D. fewer solvent molecules exist overall
30. You look outside and see that the thermometer reading is -1.00 degree Celsius. You note the puddles are not frozen. Why not?
- A. The osmotic pressure of the puddles is higher than the surrounding pressure, therefore ice does not form.
 - B. The presence of nonvolatile solutes in the puddles prevents ice formation.
 - C. The new thermometer has been dropped too many times to be accurate
 - D. Cannot determine with this information
31. The boiling point of water is 100.00 degrees Celsius. After a solute is added to water, the boiling point increases 0.75°C , which means _____.
- A. the new boiling point is 25.00°C
 - B. the new boiling point is 99.25°C
 - C. the new boiling point is 100.75°C
 - D. the new boiling point is 175.00°C
32. Adding a nonvolatile solute to a solvent _____.
- A. increases the boiling point of the solution
 - B. decreases the boiling point of the solution
 - C. increases the temperature range over which the solution remains a liquid
 - D. both A and C
33. Given the same solvent, which of the solutions would have the lowest freezing point?
- A. 1m solution containing a nonvolatile solute.
 - B. 2m solution containing a nonvolatile solute.
 - C. 3m solution containing a nonvolatile solute.
 - D. 4m solution containing a nonvolatile solute.
34. Doubling the amount of a nonvolatile solute in a solution _____.
- A. doubles the elevation of the boiling point
 - B. reduces the boiling point by one half
 - C. alters the amount of change in boiling point, but it cannot be determined by how much.
 - D. reduces the freezing point by one half

35. Osmosis is the movement of _____.
- A. a solvent, which is often water, from a lower water concentration to a higher water concentration across a selectively permeable membrane
 - B. a solvent, which is often water, from a higher water concentration to a lower water concentration across a selectively permeable membrane
 - C. solutes from a lower water concentration to a higher water concentration across a selectively permeable membrane
 - D. solutes from a higher water concentration to a lower water concentration across a selectively permeable membrane
36. Osmotic pressure is caused by _____.
- A. solvent molecules evaporating and becoming gas molecules
 - B. solute molecules evaporating and becoming gas molecules
 - C. solvent molecules moving across a selectively permeable membrane into an area of lower solvent concentration
 - D. solutes molecules moving across a selectively permeable membrane into an area of lower solvent concentration
37. Osmosis can be halted by _____.
- A. applying a pressure half as great as the osmotic pressure in the opposite direction osmosis is occurring
 - B. applying a pressure equal to the osmotic pressure, but in the opposite direction
 - C. applying a pressure greater than the osmotic pressure in the same direction osmosis is occurring
 - D. cannot be determined without additional information
38. Use the accompanying picture to determine the direction water will flow. Water flows _____.
- A. into the beaker
 - B. into the tube
 - C. neither direction because the system is at equilibrium
 - D. cannot determine with this information
- 
39. Reverse osmosis occurs when _____.
- A. a pressure half as great as the osmotic pressure is applied in the opposite direction of osmosis
 - B. a pressure equal to the osmotic pressure is applied in the opposite direction osmosis is occurring
 - C. a pressure greater than the osmotic pressure is applied in the opposite direction osmosis is occurring
 - D. cannot be determined with this information

40. Which of the following reduces vapor pressure the most?
- A. a volatile solute
 - B. a molecular solute that dissolves to give off one mole of solute per liter solution
 - C. an ionic compound that dissolves to give off two moles of solute per liter solution
 - D. an element called tungsten
41. Solutions are heterogeneous mixtures
- A. True
 - B. False
42. A label on a bottle states it contains 53% sugar. What does this mean?
- A. There are 53 milliliters of sugar per every 100 milliliters of solution.
 - B. There are 53 grams of sugar per every 100 milliliters of solution.
 - C. There are no grams of fat in the solution.
 - D. Cannot be determined with this information because units are not given.
43. Which of the following is an example of an alloy?
- A. pure gold
 - B. sterling silver
 - C. orange juice
 - D. carbon dioxide
44. A conversion factor between the mass and volume of water is one liter of water weighs one kilogram.
- A. True
 - B. False
45. When solving a chemistry problem, it is important to _____.
- A. use the periodic table to find information you may need
 - B. make sure the units cross cancel so the final answer is in the units desired
 - C. determine which variables you know and which variables you do not know
 - D. all the above
46. If you are given a quantity of solute in grams and need to convert it to moles, divide the grams by _____.
- A. the number of particles in a mole, 6.02×10^{23}
 - B. the number of elements in the solute, which varies according to solute
 - C. the molar mass of the solute, which is measured in grams/mole
 - D. by pi, 3.14

47. The information you need to determine the molar mass of an unknown solute is _____.
- A. the amount of solute used in the experiment
 - B. the molal boiling or freezing constant value of the solvent
 - C. the amount of solvent used in the experiment
 - D. all the above
48. The molecules in a gaseous solution are not evenly distributed.
- A. True
 - B. False
49. A supersaturated solution is most analogous to _____.
- A. a full airplane flight because each person has a seat
 - B. an oversold movie because there are too many people for the seats available
 - C. salt dissolving in water because salt will continue to dissolve until the solution is supersaturated
 - D. B and C
50. Which atom pulls the shared electrons more closely to itself in an O – H bond?
- A. oxygen
 - B. hydrogen

Keys to Quizzes

Quiz 1 Key	Quiz 2 Key	Quiz 3 Key	Quiz 4 Key	Quiz 5 Key
1. B	1. C	1. C	1. A	1. B
2. D	2. D	2. B	2. B	2. D
3. C	3. A	3. C	3. A	3. D
4. D	4. C	4. B	4. B	4. B
5. D	5. D	5. C	5. D	5. C
6. B	6. A	6. B	6. D	6. A
7. B	7. D	7. B	7. A	7. B
8. B	8. B	8. C	8. B	8. B
9. A	9. D	9. D	9. C	9. B
10. C	10. B	10. C	10. C	10. A

Multiple Choice Exam Key

1. C	11. D	21. B	31. C	41. B
2. B	12. A	22. B	32. D	42. D
3. D	13. D	23. C	33. D	43. B
4. C	14. B	24. A	34. A	44. A
5. D	15. B	25. A	35. B	45. D
6. D	16. C	26. B	36. C	46. C
7. C	17. D	27. D	37. B	47. D
8. D	18. B	28. C	38. A	48. B
9. A	19. B	29. A	39. C	49. B
10. C	20. C	30. B	40. C	50. A

Solutions Glossary

alloy: a solid solution that possesses properties different from the pure metals that compose the solution.

anesthesia: the partial or complete reduction of pain sensation.

aqueous solution: a solution in which water is the dissolving medium (solvent).

boiling point: the temperature at which the vapor pressure of a liquid equals atmospheric pressure.

boiling point elevation: a colligative property of solutions in which the boiling point of a pure liquid solvent increases due to the addition of a nonvolatile solute.

colligative properties: properties of a solution (vapor pressure reduction, boiling point elevation, freezing point depression, and osmotic pressure) that are affected by the number of solute particles and not the identity of the solutes themselves.

compound: a substance composed of elements bonded in a fixed ratio.

concentrated solution: an unsaturated solution that contains a relatively large amount of solute in relation to the amount of solvent. For example, a solution containing 30g of salt in 100mL of water is a concentrated solution compared to a solution containing 1g of salt in 100mL of water.

condensation: the transformation of vapor (gas) into liquid.

crystal lattice: a symmetrical arrangement of alternating cations and anions that forms a solid three dimensional structure.

dilute solution: a solution that contains a relatively small amount of solute in relation to the amount of solvent. For example, a solution containing 1g of salt in 100mL of water is a dilute solution compared to a solution containing 30 g of salt in 100mL of water.

element: a substance that cannot be separated into simpler substances; the most fundamental form of matter.

endothermic process: a process that absorbs heat energy.

equilibrium: a dynamic balance in which the rate of the forward process equals the rate of the reverse process.

evaporation: the transformation of liquid into vapor (gas).

exothermic process: a process that releases heat energy.

freezing point: the temperature at which a liquid becomes a solid.

freezing point depression: a colligative property of solutions in which the freezing point of a liquid solvent is decreased when a nonvolatile solute is added to the solvent.

heterogeneous mixture: a mixture in which substances do not occur uniformly throughout the mixture.

homogeneous mixture: a mixture in which all solutes and solvents occur uniformly throughout the mixture; a solution.

hydration: the process of water molecules surrounding and forming intermolecular attractions with dissolved solute particles.

ion: an atom or group of atoms that possesses a charge, either negative or positive, due to the gain or loss of electrons.

kinetic energy: the energy of motion.

immiscible: liquids that do not dissolve into one another.

insoluble: the inability of a substance to dissolve in a given solvent.

intermolecular attraction: attraction between molecules.

ionic compound: a compound composed of cations and anions, which occur in a fixed proportion.

matter: a term describing any substance that has mass and occupies space.

miscible: liquids capable of dissolving into one another.

mixture: a combination of two or more substances that do not react when mixed.

molality (m): a measurement of solution concentration expressed as the number of moles of solute per kilogram of solvent.

molarity (M): a measurement of solution concentration expressed as the number of moles of solute per liter of solution.

molar mass: the sum of all the molar masses of the atoms that compose a molecule, which is expressed as grams per mole.

mole: a quantity equal to 6.022×10^{23} .

mole fraction (X): a measurement of solution concentration expressed as the number of moles of one component (either solvent or solute) per the total moles of solution (solute plus solvent).

molecular substance: a collection of two or more covalently bonded atoms.

nonvolatile solute: a substance that does not evaporate when placed into a solution.

osmosis: the movement of water from an area of higher water concentration to an area of lower water concentration across a selectively permeable membrane.

osmotic pressure: a colligative property of solutions in which the pressure created by osmosis equals the pressure required to prevent osmosis.

percent solution (m/v): a measurement of solution concentration that is expressed as a percent. It is the mass of the solute in grams divided by the volume of the solvent in milliliters times 100%.

percent solution: (v/v): a measurement of solution concentration that is expressed as a percent. It is the volume of the solute divided by the volume of the solvent times 100%.

polar: a term describing a bond or molecule that has an uneven distribution of the electrons, giving one end a partial negative charge and the other end a partial positive charge.

reverse osmosis: the movement of solvent from an area of lower solvent concentration to an area of higher solvent concentration across a selectively permeable membrane due to a pressure greater than the osmotic pressure of the system.

saturated solution: a solution that contains the maximum amount of solute that dissolves at a given pressure and temperature.

Selectively-permeable membrane: a barrier that allows the passage of small ions, molecules, and compounds, while preventing the passage of large or hydrated ions and molecules.

solid solution: a solution that, when cooled, contains its component substances in a uniform and fixed position.

soluble: the ability to be dissolved in a given solvent.

solute: a substance that dissolves in a solvent.

solution: a homogeneous mixture of more than one substance that is uniformly mixed and in the same physical state.

solvation: the process of solvent molecules surrounding and forming intermolecular attractions with dissolved solute particles.

solvent: the medium in which a substance (a solute) is dissolved in a solution.

substance: any matter that has definite composition; for example, an element or a molecule.

supersaturated solution: a solution that contains more solute than can dissolve for a given pressure and temperature. Increasing temperature allows more solute to dissolve. Slow cooling of the solution sometimes keeps solutes in solution, therefore the solution becomes supersaturated.

unsaturated solution: a solution that contains less dissolved solute than the solvent can hold at a given pressure and temperature.

vapor pressure: the pressure due to vapor (gas) molecules above a liquid when the rate of condensation equals that of evaporation.

vapor pressure reduction: a colligative property of solutions in which vapor pressure is reduced when a nonvolatile solute is added to the solvent.