Chapter 12

Preview

- Objectives
- Solutions
- Suspensions
- Colloids
- Solutes: Electrolytes Versus Nonelectrolytes
Objectives

• Distinguish between electrolytes and nonelectrolytes.

• List three different solute-solvent combinations.

• Compare the properties of suspensions, colloids, and solutions.

• Distinguish between electrolytes and nonelectrolytes.
Solutions

• You know from experience that sugar dissolves in water. Sugar is described as “soluble in water.” By **soluble** we mean capable of being dissolved.

• When sugar dissolves, all its molecules become uniformly distributed among the water molecules. The solid sugar is no longer visible.

• Such a mixture is called a solution. A **solution** is a homogeneous mixture of two or more substances in a single phase.
Chapter 12

Section 1 Types of Mixtures

Solutions

Click below to watch the Visual Concept.

Visual Concept
Solutions, continued

• The dissolving medium in a solution is called the **solvent**, and the substance dissolved in a solution is called the **solute**.

• Solutions may exist as gases, liquids, or solids. There are many possible solute-solvent combinations between gases, liquids, and solids.

  • example: Alloys are solid solutions in which the atoms of two or more metals are uniformly mixed.
    • Brass is made from zinc and copper.
    • Sterling silver is made from silver and copper.
Solutes, Solvents, and Solutions

Click below to watch the Visual Concept.
Chapter 12

Types of Solutions

- Ethanol molecule, $C_2H_5OH$
- Water molecule, $H_2O$
- Liquid solute in liquid solution
- Solid solute in solid solution
- 14 karat
- Silver
- Gold

Visual Concepts
24-karat gold is pure gold.

14-karat gold is an alloy of gold with silver and copper. 14-karat gold is 14/24, or 58.3%, gold.
Suspensions

• If the particles in a solvent are so large that they settle out unless the mixture is constantly stirred or agitated, the mixture is called a suspension.

• For example, a jar of muddy water consists of soil particles suspended in water. The soil particles will eventually all collect on the bottom of the jar, because the soil particles are denser than the solvent, water.

• Particles over 1000 nm in diameter—1000 times as large as atoms, molecules or ions—form suspensions.
Suspensions

Click below to watch the Visual Concept.

Visual Concept
Colloids

- Particles that are intermediate in size between those in solutions and suspensions form mixtures known as colloidal dispersions, or simply colloids.
  - The particles in a colloid are small enough to be suspended throughout the solvent by the constant movement of the surrounding molecules.
  - Colloidal particles make up the dispersed phase, and water is the dispersing medium.
    - example: Mayonnaise is a colloid.
      - It is an emulsion of oil droplets in water.
Colloids, continued

Tyndall Effect

- Many colloids look similar to solutions because their particles cannot be seen.

- The Tyndall effect occurs when light is scattered by colloidal particles dispersed in a transparent medium.
  - example: a headlight beam is visible from the side on a foggy night.

- The Tyndall effect can be used to distinguish between a solution and a colloid.
Colloids

Not a colloid

Colloid
Emulsions

Visual Concepts

Chapter 12

Emulsion

Emulsion Stabilized with an Emulsifier
## Properties of Solutions, Colloids, and Suspensions

<table>
<thead>
<tr>
<th>Solutions</th>
<th>Colloids</th>
<th>Suspensions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Homogeneous</td>
<td>Heterogeneous</td>
<td>Heterogeneous</td>
</tr>
<tr>
<td>Particle size: 0.01–1 nm; can be atoms, ions, molecules</td>
<td>Particle size: 1–1000 nm, dispersed; can be aggregates or large molecules</td>
<td>Particle size: over 1000 nm, suspended; can be large particles or aggregates</td>
</tr>
<tr>
<td>Do not separate on standing</td>
<td>Do not separate on standing</td>
<td>Particles settle out</td>
</tr>
<tr>
<td>Cannot be separated by filtration</td>
<td>Cannot be separated by filtration</td>
<td>Can be separated by filtration</td>
</tr>
<tr>
<td>Do not scatter light</td>
<td>Scatter light (Tyndall effect)</td>
<td>May scatter light, but are not transparent</td>
</tr>
</tbody>
</table>
Solute: Electrolytes Versus Nonelectrolytes

• A substance that dissolves in water to give a solution that conducts electric current is called an electrolyte.

• Any soluble ionic compound, such as sodium chloride, NaCl, is an electrolyte.

• The positive and negative ions separate from each other in solution and are free to move, making it possible for an electric current to pass through the solution.
A substance that dissolves in water to give a solution that does not conduct electric current is called a **nonelectrolyte**.

Sugar is an example of a nonelectrolyte.

Neutral solute molecules do not contain mobile charged particles, so a solution of a nonelectrolyte cannot conduct electric current.
Chapter 12

Section 1  Types of Mixtures

Electrical Conductivity of Solutions

Sucrose solution

Acetic acid solution

Sodium chloride solution

Hydrochloric acid solution

Sucrose molecule, C₁₂H₂₂O₁₁

Acetic acid, CH₃COOH

Acetate ion, CH₃COO⁻

Hydronium ion, H₃O⁺

Chloride ion, Cl⁻

Sodium ion, Na⁺

Hydronium ion, H₃O⁺

Chloride ion, Cl⁻
Chapter 12

Section 2 The Solution Process

Preview

- Objectives
- Factors Affecting the Rate of Dissolution
- Solubility
- Solute-Solvent Interactions
- Enthalpies of Solution
Objectives

- **List** and explain three factors that affect the rate at which a solid solute dissolves in a liquid solvent.

- **Explain** solution equilibrium, and distinguish among saturated, unsaturated, and supersaturated solutions.

- **Explain** the meaning of “like dissolves like” in terms of polar and nonpolar substances.
Objectives, *continued*

- **List** the three interactions that contribute to the enthalpy of a solution, and explain how they combine to cause dissolution to be exothermic or endothermic.

- **Compare** the effects of temperature and pressure on solubility.
Dissolving Process

Click below to watch the Visual Concept.
Factors Affecting the Rate of Dissolution

- Because the dissolution process occurs at the surface of the solute, it can be speeded up if the surface area of the solute is increased.

- Stirring or shaking helps to disperse solute particles and increase contact between the solvent and solute surface. This speeds up the dissolving process.

- At higher temperatures, collisions between solvent molecules and solvent are more frequent and of higher energy. This helps to disperse solute molecules among the solvent molecules, and speed up the dissolving process.
Factors Affecting the Rate of Dissolution

Click below to watch the Visual Concept.
Solubility

- If you add spoonful after spoonful of sugar to tea, eventually no more sugar will dissolve.

- This illustrates the fact that for every combination of solvent with a solid solute at a given temperature, there is a limit to the amount of solid that can be dissolved.

- The point at which this limit is reached for any solute-solvent combination depends on the nature of the solute, the nature of the solvent, and the temperature.
Particle Model for Soluble and Insoluble Substances

Chapter 12

Section 2 The Solution Process

Water and lithium chloride

Soluble

Water
Chloride ion, Cl^-
Lithium ion, Li^+

Toluene and lithium chloride

Insoluble

Toluene
Lithium chloride
Particle Model for Soluble and Insoluble Substances

Water and toluene

Insoluble

Toluene

Water
Solubility, *continued*

- When a solute is first added to a solvent, solute molecules leave the solid surface and move about at random in the solvent.

- As more solute is added, more collisions occur between dissolved solute particles. Some of the solute molecules return to the crystal.

- When maximum solubility is reached, molecules are returning to the solid form at the same rate at which they are going into solution.
Solubility, *continued*

- *Solution equilibrium* is the physical state in which the opposing processes of dissolution and crystallization of a solute occur at the same rates.
Solution Equilibrium

Click below to watch the Visual Concept.
Solubility, *continued*

Saturated Versus Unsaturated Solutions

- A solution that contains the maximum amount of dissolved solute is described as a **saturated solution**.
  - If more solute is added to a saturated solution, it falls to the bottom of the container and does not dissolve.
  - This is because an equilibrium has been established between ions leaving and entering the solid phase.

- A solution that contains less solute than a saturated solution under the same conditions is an **unsaturated solution**.
Mass of Solute Added Versus Mass of Solute Dissolved

**A. Unsaturated**
If a solution is unsaturated, more solute can dissolve. No undissolved solute remains.

**B. Saturated**
If the amount of solute added exceeds the solubility, some solute remains undissolved.

**Solubility = 46.4 g/100 g**
Solubility, *continued*

**Supersaturated Solutions**

- When a saturated solution is cooled, the excess solute usually comes out of solution, leaving the solution saturated at the lower temperature.

- But sometimes the excess solute does not separate, and a *supersaturated solution* is produced, which is a solution that contains more dissolved solute than a saturated solution contains under the same conditions.

- A supersaturated solution will form crystals of solute if disturbed or more solute is added.
Solubility, *continued*

**Solubility Values**

- The **solubility** of a substance is the amount of that substance required to form a saturated solution with a specific amount of solvent at a specified temperature.
  - **example:** The solubility of sugar is 204 g per 100 g of water at 20°C.

- Solubilities vary widely, and must be determined experimentally.
  - They can be found in chemical handbooks and are usually given as grams of solute per 100 g of solvent at a given temperature.
### Solubility of Common Compounds

Compounds containing these ions are soluble in water . . . unless they also contain these ions, which make them insoluble.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Ion</th>
<th>Insoluble Ions</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonium</td>
<td>NH$_4^+$</td>
<td></td>
</tr>
<tr>
<td>potassium</td>
<td>K$^+$</td>
<td></td>
</tr>
<tr>
<td>sodium</td>
<td>Na$^+$</td>
<td></td>
</tr>
<tr>
<td>acetate</td>
<td>C$_2$H$_3$O$_2^-$</td>
<td>Fe$^{3+}$, Al$^{3+}$, Hg$_2^{2+}$</td>
</tr>
<tr>
<td>chlorate</td>
<td>ClO$_3^-$</td>
<td></td>
</tr>
<tr>
<td>chloride</td>
<td>Cl$^-$</td>
<td>Ag$^+$, Hg$_2^{2+}$, Pb$^{2+}$</td>
</tr>
<tr>
<td>nitrate</td>
<td>NO$_3^-$</td>
<td></td>
</tr>
<tr>
<td>sulfate</td>
<td>SO$_4^{2-}$</td>
<td>Ca$^{2+}$, Ba$^{2+}$, Pb$^{2+}$, Sr$^{2+}$, Hg$_2^{2+}$</td>
</tr>
</tbody>
</table>
### Solubility of Common Compounds

<table>
<thead>
<tr>
<th>Compounds containing these ions are insoluble in water . . .</th>
<th>. . . unless they also contain these ions, which make them soluble.</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbonate</td>
<td>CO$_3^{2-}$</td>
</tr>
<tr>
<td>hydroxide</td>
<td>OH$^-$</td>
</tr>
<tr>
<td>oxide</td>
<td>O$_{2-}$</td>
</tr>
<tr>
<td>phosphate</td>
<td>PO$_4^{3-}$</td>
</tr>
<tr>
<td>silicate</td>
<td>SiO$_3^{2-}$</td>
</tr>
<tr>
<td>sulfide</td>
<td>S$^{2-}$</td>
</tr>
<tr>
<td>sulfite</td>
<td>SO$_3^{2-}$</td>
</tr>
</tbody>
</table>
Chapter 12

Solubility of a Solid in a Liquid

Click below to watch the Visual Concept.
Solute-Solvent Interactions

- Solubility varies greatly with the type of compounds involved.

- "Like dissolves like" is a rough but useful rule for predicting whether one substance will dissolve in another.

- What makes substances similar depends on:
  - type of bonding
  - polarity or nonpolarity of molecules
  - intermolecular forces between the solute and solvent
Like Dissolves Like

Click below to watch the Visual Concept.
Solute-Solvent Interactions, *continued*

Dissolving Ionic Compounds in Aqueous Solution

- The polarity of water molecules plays an important role in the formation of solutions of ionic compounds in water.

- The slightly charged parts of water molecules attract the ions in the ionic compounds and surround them, separating them from the crystal surface and drawing them into the solution.

- This solution process with water as the solvent is referred to as *hydration*. The ions are said to be *hydrated*. 
Solute-Solvent Interactions, continued

Dissolving Ionic Compounds in Aqueous Solution

The hydration of the ionic solute lithium chloride is shown below.
Solute-Solvent Interactions, continued

Nonpolar Solvents

- Ionic compounds are generally not soluble in nonpolar solvents such as carbon tetrachloride, $\text{CCl}_4$, and toluene, $\text{C}_6\text{H}_5\text{CH}_3$.

- The nonpolar solvent molecules do not attract the ions of the crystal strongly enough to overcome the forces holding the crystal together.

- Ionic and nonpolar substances differ widely in bonding type, polarity, and intermolecular forces, so their particles cannot intermingle very much.
Solute-Solvent Interactions, continued

Liquid Solutes and Solvents

• Oil and water do not mix because oil is nonpolar whereas water is polar. The hydrogen bonding between water molecules squeezes out whatever oil molecules may come between them.

• Two polar substances, or two nonpolar substances, on the other hand, form solutions together easily because their intermolecular forces match.

• Liquids that are not soluble in each other are **immiscible**. Liquids that dissolve freely in one another in any proportion are **miscible**.
Chapter 12

Section 2  The Solution Process

Comparing Miscible and Immiscible Liquids

Click below to watch the Visual Concept.
Solute-Solvent Interactions, *continued*

Effects of Pressure on Solubility

- Changes in pressure have very little effect on the solubilities of liquids or solids in liquid solvents. However, increases in pressure increase gas solubilities in liquids.

- An equilibrium is established between a gas above a liquid solvent and the gas dissolved in a liquid.

- As long as this equilibrium is undisturbed, the solubility of the gas in the liquid is unchanged at a given pressure:

  \[
  \text{gas} + \text{solvent} \rightleftharpoons \text{solution}
  \]
Solute-Solvent Interactions, *continued*

Effects of Pressure on Solubility, *continued*  

- Increasing the pressure of the solute gas above the solution causes gas particles to collide with the liquid surface more often. This causes more gas particles to dissolve in the liquid.

\[ \uparrow \text{gas} + \text{solvent} \quad \leftrightarrow \quad \text{solution} \]

- Decreasing the pressure of the solute gas above the solution allows more dissolved gas particles to escape from solution.

\[ \downarrow \text{gas} + \text{solvent} \quad \leftrightarrow \quad \text{solution} \]
Pressure, Temperature, and Solubility of Gases

Click below to watch the Visual Concept.
Solute-Solvent Interactions, *continued*

**Henry’s Law**

- **Henry’s law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of that gas on the surface of the liquid.

  - In carbonated beverages, the solubility of carbon dioxide is increased by increasing the pressure. The sealed containers contain CO$_2$ at high pressure, which keeps the CO$_2$ dissolved in the beverage, above the liquid.

  - When the beverage container is opened, the pressure above the solution is reduced, and CO$_2$ begins to escape from the solution.

- The rapid escape of a gas from a liquid in which it is dissolved is known as *effervescence*. 
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Section 2  The Solution Process

Henry’s Law

Click below to watch the Visual Concept.

Visual Concept
Chapter 12

Section 2 The Solution Process

Effervescence

Click below to watch the Visual Concept.

Visual Concept
Solute-Solvent Interactions, continued

Effects of Temperature on Solubility

• Increasing the temperature usually decreases gas solubility.

  • As temperature increases, the average kinetic energy of molecules increases.

  • A greater number of solute molecules are therefore able to escape from the attraction of solvent molecules and return to the gas phase.

  • At higher temperatures, therefore, equilibrium is reached with fewer gas molecules in solution, and gases are generally less soluble.
Solute-Solvent Interactions, continued

Effects of Temperature on Solubility

- Increasing the temperature usually increases solubility of solids in liquids, as mentioned previously.

- The effect of temperature on solubility for a given solute is difficult to predict.

- The solubilities of some solutes vary greatly over different temperatures, and those for other solutes hardly change at all.

- A few solid solutes are actually less soluble at higher temperatures.
Solubility vs. Temperature

Temperature-Solubility Relationships

- CsCl
- NaNO₃
- NaC₂H₃O₂
- RbCl
- LiCl
- NH₄Cl
- KCl
- NaCl
- Li₂SO₄
- SO₂

Solubility in grams per 100 g of water vs. Temperature (°C)
Enthalpies of Solution

- The formation of a solution is accompanied by an energy change.

- If you dissolve some potassium iodide, KI, in water, you will find that the outside of the container feels cold to the touch.

- But if you dissolve some sodium hydroxide, NaOH, in the same way, the outside of the container feels hot.

- The formation of a solid-liquid solution can either absorb energy (KI in water) or release energy as heat (NaOH in water)
Enthalpies of Solution, continued

• Before dissolving begins, solute particles are held together by intermolecular forces. Solvent particles are also held together by intermolecular forces.

• Energy changes occur during solution formation because energy is required to separate solute molecules and solvent molecules from their neighbors.

• A solute particle that is surrounded by solvent molecules is said to be solvated.
The diagram above shows the enthalpy changes that occur during the formation of a solution.
Enthalpies of Solution, continued

• The net amount of energy absorbed as heat by the solution when a specific amount of solute dissolves in a solvent is the **enthalpy of solution**.

• The enthalpy of solution is negative (energy is released) when the sum of attractions from Steps 1 and 2 is less than Step 3, from the diagram on the previous slide.

• The enthalpy of solution is positive (energy is absorbed) when the sum of attractions from Steps 1 and 2 is greater than Step 3.
Chapter 12

Section 3 Concentration of Solutions

Preview

- Objectives
- Concentration
- Molarity
- Molality
Objectives

• **Given** the mass of solute and volume of solvent, calculate the concentration of solution.

• **Given** the concentration of a solution, determine the amount of solute in a given amount of solution.

• **Given** the concentration of a solution, determine the amount of solution that contains a given amount of solute.
Concentration

- The concentration of a solution is a measure of the amount of solute in a given amount of solvent or solution.

- Concentration is a ratio: any amount of a given solution has the same concentration.

- The opposite of concentrated is *dilute*.

- These terms are unrelated to the degree to which a solution is saturated: a saturated solution of a solute that is not very soluble might be very dilute.
## Concentration Units

<table>
<thead>
<tr>
<th>Name</th>
<th>Abbr.</th>
<th>Units</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>grams/100. g</td>
<td>g/100.g</td>
<td>(\frac{\text{g solute}}{100. \text{ g solvent}})</td>
<td>solubility descriptions, medical products</td>
</tr>
<tr>
<td>mass percent or “weight percent”</td>
<td>%</td>
<td>(\frac{\text{g solute}}{100. \text{ g solution}})</td>
<td>biological research</td>
</tr>
<tr>
<td>parts per million</td>
<td>ppm</td>
<td>(\frac{\text{g solute}}{1 \times 10^6 \text{ g solution}})*</td>
<td>small concentrations</td>
</tr>
<tr>
<td>parts per billion</td>
<td>ppb</td>
<td>(\frac{\text{g solute}}{1 \times 10^9 \text{ g solution}})*</td>
<td>very small concentrations, as in pollutants or contaminants</td>
</tr>
<tr>
<td>parts per trillion</td>
<td>ppt</td>
<td>(\frac{\text{g solute}}{1 \times 10^{12} \text{ g solution}})*</td>
<td>extremely small concentrations, as in isotopes used as tracers in medicine</td>
</tr>
<tr>
<td>molarity</td>
<td>M</td>
<td>(\frac{\text{mol solute}}{\text{L solution}})</td>
<td>laboratory chemistry, where the solute may undergo a chemical change according to a mole ratio</td>
</tr>
<tr>
<td>molality</td>
<td>m</td>
<td>(\frac{\text{mol solute}}{\text{kg solvent}})</td>
<td>calculation of special properties such as boiling-point elevation and freezing-point depression</td>
</tr>
</tbody>
</table>

*volume for gases
Concentration

Click below to watch the Visual Concept.
Molarity •

Molarity is the number of moles of solute in one liter of solution.

• For example, a “one molar” solution of sodium hydroxide contains one mole of NaOH in every liter of solution.

• The symbol for molarity is M. The concentration of a one molar NaOH solution is written 1 M NaOH.
Molarity, *continued*

- To calculate molarity, you must know the amount of solute in moles and the volume of solution in liters.

- When weighing out the solute, this means you will need to know the molar mass of the solute in order to convert mass to moles.

  - *example:* One mole of NaOH has a mass of 40.0 g. If this quantity of NaOH is dissolved in enough water to make 1.00 L of solution, it is a 1.00 M solution.
Molarity, continued

- The molarity of any solution can be calculated by dividing the number of moles of solute by the number of liters of solution:

$$\text{molarity (M)} = \frac{\text{amount of solute (mol)}}{\text{volume of solution (L)}}$$

- Note that a 1 M solution is not made by adding 1 mol of solute to 1 L of solvent. In such a case, the final total volume of the solution might not be 1 L.

- Solvent must be added carefully while dissolving to ensure a final volume of 1 L.
Chapter 12

Preparation of a Solution Based on Molarity

Click below to watch the Visual Concept.

Visual Concept
Molarity, *continued*

**Sample Problem A**

You have 3.50 L of solution that contains 90.0 g of sodium chloride, NaCl. What is the molarity of that solution?
Molarity, continued

Sample Problem A Solution

**Given:** solute mass = 90.0 g NaCl
solution volume = 3.50 L

**Unknown:** molarity of NaCl solution

**Solution:**

\[
\text{90.0 g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 1.54 \text{ mol NaCl}
\]

\[
\frac{1.54 \text{ mol NaCl}}{3.50 \text{ L of solution}} = 0.440 \text{ M NaCl}
\]
Molarity, continued

Sample Problem B

You have 0.8 L of a 0.5 M HCl solution. How many moles of HCl does this solution contain?
Molarity, *continued*

**Sample Problem B Solution**

**Given:** volume of solution = 0.8 L  
concentration of solution = 0.5 M HCl  

**Unknown:** moles of HCl in a given volume  

**Solution:**  

\[
\frac{0.5 \text{ mol HCl}}{1.0 \text{ L of solution}} \times 0.8 \text{ L of solution} = 0.4 \text{ mol HCl}
\]
Molarity, continued

Sample Problem C

To produce 40.0 g of silver chromate, you will need at least 23.4 g of potassium chromate in solution as a reactant. All you have on hand is 5 L of a 6.0 M $\text{K}_2\text{CrO}_4$ solution. What volume of the solution is needed to give you the 23.4 g $\text{K}_2\text{CrO}_4$ needed for the reaction?
Molarity, *continued*

**Sample Problem C Solution**

**Given:**
- volume of solution = 5 L
- concentration of solution = 6.0 M $K_2CrO_4$
- mass of solute = 23.4 $K_2CrO_4$
- mass of product = 40.0 g $Ag_2CrO_4$

**Unknown:** volume of $K_2CrO_4$ solution in L
Molarity, *continued*

Sample Problem C Solution, *continued*  ▼

**Solution:**  ▼

\[ 23.4 \text{ g } K_2CrO_4 \times \frac{1 \text{ mol } K_2CrO_4}{194.2 \text{ g } K_2CrO_4} = 0.120 \text{ mol } K_2CrO_4 \]

\[ 6.0 \text{ M } K_2CrO_4 = \frac{0.120 \text{ mol } K_2CrO_4}{x \text{ L } K_2CrO_4 \text{ solution}} \]

\[ x = 0.020 \text{ L } K_2CrO_4 \text{ solution} \]
Molality

• **Molality** is the concentration of a solution expressed in moles of solute per kilogram of solvent.

• A solution that contains 1 mol of solute dissolved in 1 kg of solvent is a “one molal” solution.

• The symbol for molality is $m$, and the concentration of this solution is written as $1 \, m \, \text{NaOH}$. 
Molality, *continued*

- The molality of any solution can be calculated by dividing the number of moles of solute by the number of kilograms of solvent:

\[
m = \frac{\text{amount of solute (mol)}}{\text{mass of solvent (kg)}}
\]

- Unlike molarity, which is a ratio of which the denominator is liters of solution, molality is per kilograms of solvent.

- Molality is used when studying properties of solutions related to vapor pressure and temperature changes, because molality does not change with temperature.
Comparing Molarity and Molality

Click below to watch the Visual Concept.
Making a Molal Solution

Calculate the mass of CuSO$_4$$\cdot$5H$_2$O needed. To make this solution, each kilogram of solvent (1000 g) will require 0.5000 mol of CuSO$_4$$\cdot$5H$_2$O. This mass is calculated to be 124.8 g.

Add exactly 1 kg of solvent to the solute in the beaker. Because the solvent is water, 1 kg will equal 1000 mL.

Mix thoroughly.

The resulting solution has 0.5000 mol of solute dissolved in 1 kg of solvent.
Molality, continued

Sample Problem D

A solution was prepared by dissolving 17.1 g of sucrose (table sugar, C_{12}H_{22}O_{11}) in 125 g of water. Find the molal concentration of this solution.
Molality, continued

Sample Problem D Solution

**Given:**
- solute mass = 17.1 $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
- solvent mass = 125 g $\text{H}_2\text{O}$

**Unknown:** molal concentration

**Solution:** First, convert grams of solute to moles and grams of solvent to kilograms.

$$17.1 \text{ g }\text{C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{1 \text{ mol }\text{C}_{12}\text{H}_{22}\text{O}_{11}}{342.34 \text{ g }\text{C}_{12}\text{H}_{22}\text{O}_{11}} = 0.0500 \text{ mol }\text{C}_{12}\text{H}_{22}\text{O}_{11}$$

$$\frac{125 \text{ g }\text{H}_2\text{O}}{1000 \text{ g/kg}} = 0.125 \text{ kg }\text{H}_2\text{O}$$
Molality, *continued*

*Sample Problem D Solution, continued*

Then, divide moles of solute by kilograms of solvent.

\[
\frac{0.0500 \text{ mol} \ C_{12}H_{22}O_{11}}{0.125 \text{ kg} \ H_2O} = 0.400 \ m \ C_{12}H_{22}O_{11}
\]
Molality, *continued*

**Sample Problem E**  
A solution of iodine, $I_2$, in carbon tetrachloride, $CCl_4$, is used when iodine is needed for certain chemical tests. How much iodine must be added to prepare a 0.480 $m$ solution of iodine in $CCl_4$ if 100.0 g of $CCl_4$ is used?
Molality, *continued*

**Sample Problem E Solution**

**Given:** molality of solution = 0.480 $m$ $I_2$

mass of solvent = 100.0 g CCl$_4$

**Unknown:** mass of solute

**Solution:** First, convert grams of solvent to kilograms.

\[
\frac{100.0 \text{ g CCl}_4}{1000 \text{ g/kg}} = 0.100 \text{ kg CCl}_4
\]
Molality, *continued*

Sample Problem E Solution, *continued*  

**Solution, *continued***: Then, use the equation for molality to solve for moles of solute.

\[
0.480 \, m = \frac{x \, \text{mol} \, l_2}{0.1 \, \text{kg} \, \text{H}_2\text{O}} \quad x = 0.0480 \, \text{mol} \, l_2
\]

Finally, convert moles of solute to grams of solute.

\[
0.480 \, \text{mol} \, l_2 \times \frac{253.8 \, \text{g} \, l_2}{\text{mol} \, l_2} = 12.2 \, \text{g} \, l_2
\]
End of Chapter 12 Show