Drinking seawater can cause dehydration.

Seawater:
- Is a homogeneous mixture of salts with water
- Contains higher concentrations of salts than the salt content of your cells
- As seawater passes through your body, it pulls water out of your cells, due mainly to nature’s tendency toward spontaneous mixing.
- This reduces your cells’ water level and usually results in diarrhea as this extra liquid flows out with the seawater.

Seawater will dehydrate you and give you diarrhea.

The cell wall acts as a barrier to solute moving so the only way for the seawater and the cell solution to have uniform mixing is for water to flow out of the cells of your intestine and into your digestive tract.

Seawater is a mixture of two or more substances.

Composition may vary from one sample to another.

Appears to be one substance, though really contains multiple materials.

Most homogeneous materials we encounter are actually solutions.
- For example, air and seawater

Nature has a tendency toward spontaneous mixing.
- Generally, uniform mixing is more energetically favorable.
Solutions

• The majority component of a solution is called the **solvent**.
• The minority component is called the **solute**.
• Solutions form in part because of intermolecular forces.
  – The particles of the solute interact with the particles of the solvent through intermolecular forces.

Common Types of Solutions

• A solution may be composed of a solid and a liquid, a gas and a liquid, or other combinations.

<table>
<thead>
<tr>
<th>TABLE 12.1 Common Types of Solutions</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Solution Phase</strong></td>
</tr>
<tr>
<td>Gaseous solution</td>
</tr>
<tr>
<td>Liquid solution</td>
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<td></td>
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<tr>
<td></td>
</tr>
<tr>
<td>Solid solution</td>
</tr>
</tbody>
</table>

Solubility

• When one substance (solute) dissolves in another (solvent) it is said to be **soluble**.
  – Salt is soluble in water.
  – Bromine is soluble in methylene chloride.

• When one substance does not dissolve in another it is said to be **insoluble**.
  – Oil is insoluble in water

• The solubility of one substance in another depends on
  1. nature’s tendency toward mixing, and
  2. the types of intermolecular attractive forces.

Spontaneous Mixing

• When solutions with different solute concentrations come in contact, they spontaneously mix to result in a uniform distribution of solute throughout the solution.

Common Types of Solutions

• In aqueous solutions, water is the solvent.

Nature’s Tendency Toward Mixing: Entropy

• Many physical systems tend toward lower potential energy.

• But formation of a solution does not necessarily lower the potential energy of the system.

• When two ideal gases are put into the same container, they spontaneously mix, even though the difference in attractive forces is negligible.

• The gases mix because the energy of the system is lowered through the release of entropy.
Mixing and the Solution Process Entropy

• **Entropy** is the measure of energy dispersal throughout the system.

• Energy has a spontaneous drive to spread out over as large a volume as it is allowed.

• By each gas expanding to fill the container, it spreads its energy out and lowers its entropy.

Solutions: Effect of Intermolecular Forces

• Energy changes in the formation of most solutions also involve differences in attractive forces between the particles.

**Intermolecular Forces**

- Dispersion
- Dipole-dipole
- Hydrogen bond
- Ion-dipole

Solutions: Effect of Intermolecular Forces

• For the solvent and solute to mix you must overcome:
  1. all of the solute–solute attractive forces, or
  2. some of the solvent–solvent attractive forces.

  - Both processes are endothermic

  - At least some of the energy to do this comes from making new solute–solvent attractions, which is exothermic.

Relative Interactions and Solution Formation

<table>
<thead>
<tr>
<th>TABLE 12.2 Relative Interactions and Solution Formation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solvent-solute interactions</td>
</tr>
<tr>
<td>Solvent-solute interactions</td>
</tr>
<tr>
<td>Solvent-solute interactions</td>
</tr>
</tbody>
</table>

• When the solute-to-solvent attractions are weaker than the sum of the solute-to-solvent and solvent-to-solvent attractions, the solution will only form if the energy difference is small enough to be overcome by the increase in entropy from mixing.

Solution Interactions

• The maximum amount of solute that can be dissolved in a given amount of solvent is called the **solubility**.

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

    - Gases are **always** soluble in each other.
    - Two liquids that are mutually soluble are said to be **miscible**.
      - Alcohol and water are miscible.
      - Oil and water are immiscible.

  - The solubility of one substance in another varies with temperature and pressure.
Will It Dissolve?

- Chemist’s rule of thumb – *like dissolves like*

- A chemical will dissolve in a solvent if it has a similar structure to the solvent.

- Polar molecules and ionic compounds will be more soluble in polar solvents.

- Nonpolar molecules will be more soluble in nonpolar solvents.

Example 12.1 Solubility

Will It Dissolve?

Vitamins are often categorized as either fat soluble or water soluble. Water-soluble vitamins dissolve in body fluids and are easily eliminated in the urine, so there is little danger of overconsumption. Fat-soluble vitamins, on the other hand, can accumulate in the body’s fatty deposits. Overconsumption of a fat-soluble vitamin can be dangerous to your health. Examine the structure of each vitamin shown here and classify it as either fat soluble or water soluble.

Example 12.1 Solubility Continued

Solution

a. The four –OH bonds in vitamin C make it highly polar and allow it to hydrogen bond with water. Vitamin C is water soluble.

b. The C=O bonds in vitamin K₃ are nonpolar and the C—H bonds are nearly so. The C—O bonds are polar, but the bond dipoles oppose and largely cancel each other, so the molecule is dominated by the nonpolar bonds. Vitamin K₃ is fat soluble.

c. The C—C bonds in vitamin A are nonpolar and the C—H bonds are nearly so. The –OH bond may increase its water solubility slightly, but overall vitamin A is nonpolar and therefore fat soluble.

d. The three –OH bonds and one –NH bond in vitamin B₅ make it highly polar and allow it to hydrogen bond with water. Vitamin B₅ is water soluble.

For Practice 12.1

Determine whether each compound is soluble in hexane.

a. water (H₂O)

b. propane (CH₃CH₂CH₃)

c. ammonia (NH₃)

d. hydrogen chloride (HCl)

Heat of Solution

- When some compounds, such as NaOH, dissolve in water, a lot of heat is released. – The container gets hot.

- When other compounds, such as NH₄NO₃, dissolve in water, heat is absorbed from the surroundings. – The container gets cold.

- Why is this?

Energetics of Solution Formation: The Enthalpy of Solution

- To make a solution you must

  1. overcome all attractions between the solute particles; therefore, \( \Delta H_{\text{solute}} \) is **endothermic**.

  2. overcome some attractions between solvent molecules; therefore, \( \Delta H_{\text{solvent}} \) is **endothermic**.

  3. form new attractions between solute particles and solvent molecules; therefore, \( \Delta H_{\text{mix}} \) is **exothermic**.

  - The overall \( \Delta H \) for making a solution depends on the relative sizes of the \( \Delta H \) for these three processes:

    \[
    \Delta H_{\text{sol}} = \Delta H_{\text{solute}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}
    \]
Solution Process

Step 1: Separating the solute into its constituent particles

\[ \Delta H_{\text{solute}} > 0 \]

Solution Process

Step 2: Separating the solvent particles from each other to make room for the solute particles

\[ \Delta H_{\text{solvent}} > 0 \]

Solution Process

Step 3: Mixing the solute particles with the solvent particles

\[ \Delta H_{\text{mix}} < 0 \]

Energetics of Solution Formation

If the total energy cost for breaking attractions between particles in the pure solute and pure solvent is less than the energy released in making the new attractions between the solute and solvent, the overall process will be exothermic.

Heats of Hydration

- For aqueous solutions of ionic compounds, the energy added to overcome the attractions between water molecules and the energy released in forming attractions between the water molecules and ions is combined into a term called the heat of hydration.
  - Attractive forces between ions = lattice energy
    \[ \Delta H_{\text{lattice}} = -\Delta H_{\text{lattice energy}} \]
  - Attractive forces in water = H bonds
  - Attractive forces between ion and water = ion–dipole
  - \[ \Delta H_{\text{hydration}} = \text{heat released when 1 mole of gaseous ions dissolves in water} = \Delta H_{\text{solvent}} + \Delta H_{\text{mix}} \]
Ion–Dipole Interactions

- When ions dissolve in water they become hydrated.
  - Each ion is surrounded by water molecules.
- The formation of these ion–dipole attractions causes the heat of hydration to be very exothermic.

Heats of Solution for Ionic Compounds

- For an aqueous solution of an ionic compound, the $\Delta H_{\text{solution}}$ is the difference between the heat of hydration and the lattice energy.

$$
\Delta H_{\text{solution}} = \Delta H_{\text{solute}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}
$$

$$
\Delta H_{\text{solution}} = -\Delta H_{\text{lattice energy}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}
$$

$$
\Delta H_{\text{solution}} = \Delta H_{\text{hydration}} - \Delta H_{\text{lattice energy}}
$$

Heat of Hydration

$$
\Delta H_{\text{solution}} = \Delta H_{\text{hydration}} - \Delta H_{\text{lattice energy}}
$$

Comparing Heat of Solution to Heat of Hydration

- Because the lattice energy is always exothermic, the size and sign on the $\Delta H_{\text{sol}}$ tells us something about $\Delta H_{\text{hydration}}$.

- If the heat of solution is large and endothermic, then the amount of energy it costs to separate the ions is more than the energy released from hydrating the ions.

$$
\Delta H_{\text{hydration}} < \Delta H_{\text{lattice}} \text{ when } \Delta H_{\text{sol}} > 0
$$

- If the heat of solution is large and exothermic, then the amount of energy it costs to separate the ions is less than the energy released from hydrating the ions.

$$
\Delta H_{\text{hydration}} > \Delta H_{\text{lattice}} \text{ when } \Delta H_{\text{sol}} < 0
$$

Solution Equilibrium

- The dissolution of a solute in a solvent is an equilibrium process.
- Initially, when there is no dissolved solute, the only process possible is dissolution.
- Shortly after some solute is dissolved, solute particles can start to recombine to reform solute molecules, but the rate of dissolution $\gg$ the rate of deposition and the solute continues to dissolve.
- Eventually, the rate of dissolution $=$ the rate of deposition—the solution is saturated with solute and no more solute will dissolve.
A solution that has the solute and solvent in dynamic equilibrium is said to be **saturated**.
- If you add more solute it will not dissolve.
- The saturation concentration depends on the temperature and pressure of gases.

A solution that has less solute than saturation is said to be **unsaturated**.
- More solute will dissolve at this temperature.

A solution that has more solute than saturation is said to be **supersaturated**.

Solubility Limit

How Can You Make a Solvent Hold More Solute Than It Is Able To?

- Solutions can be made saturated at non-room conditions, and then can be allowed to come to room conditions slowly.
- For some solutes, instead of coming out of solution when the conditions change, they get stuck between the solvent molecules and the solution becomes supersaturated.
- Supersaturated solutions are unstable and lose all the solute above saturation when disturbed.
  - For example, shaking a carbonated beverage

Adding a Crystal of NaC₆H₃O₂ to a Supersaturated Solution

Temperature Dependence of Solubility of Solids in Water

- Solubility is generally given in grams of solute that will dissolve in 100 g of water.
- For most solids, the solubility of the solid increases as the temperature increases.
  - When $\Delta H_{\text{solution}}$ is endothermic
- Solubility curves can be used to predict whether a solution with a particular amount of solute dissolved in water is saturated (on the line), unsaturated (below the line), or supersaturated (above the line).

Solubility Curves

Purification by Recrystallization

- One of the common operations performed by a chemist is removing impurities from a solid compound.
- One method of purification involves dissolving a solid in a hot solvent until the solution is saturated.
- As the solution slowly cools, the solid crystallizes out, leaving impurities behind.
Temperature Dependence of Solubility of Gases in Water

- Gases generally have lower solubility in water than ionic or polar covalent solids because most are nonpolar molecules.
  - Gases with high solubility usually are actually reacting with water.
- For all gases, the solubility of the gas decreases as the temperature increases.
  - The ΔH_{solution} is exothermic because you do not need to overcome solute–solute attractions.

Pressure Dependence of Solubility of Gases in Water

- The larger the partial pressure of a gas in contact with a liquid, the more soluble the gas is in the liquid.

Henry’s Law

- The solubility of a gas (S_{gas}) is directly proportional to its partial pressure, (P_{gas}).
  \[ S_{gas} = k_H P_{gas} \]
- \( k_H \) is called the Henry’s law constant.

**Example 12.2 Henry’s Law**

What pressure of carbon dioxide is required to keep the carbon dioxide concentration in a bottle of club soda at 0.12 M at 25°C?

**Solve**

You are given the desired solubility of carbon dioxide and asked to find the pressure required to achieve this solubility.

**Given:** \( S_{CO_2} = 0.12 \text{ M} \)

**Find:** \( P_{CO_2} \)

**Strategize**

Use Henry’s law to find the required pressure from the solubility. You will need the Henry’s law constant for carbon dioxide, which is listed in Table 12.4.

**Conceptual Plan**

\[ S_{CO_2} = k_H P_{CO_2} \]

**Relationships Used**

\[ k_H = k_H(CO_2, \text{Henry’s law}) \]

**Example 12.2 Henry’s Law Continued**

**Solution**

\[ S_{CO_2} = k_H P_{CO_2} \]

\[ \frac{S_{CO_2}}{k_H} = P_{CO_2} \]

**Check**

The answer is in the correct units and seems reasonable. A small answer (for example, less than 1 atm) would be suspect because you know that the soda is under a pressure greater than atmospheric pressure when you open it. A very large answer (for example, over 100 atm) would be suspect because an ordinary can or bottle probably could not maintain such high pressures without bursting.

For Practice 12.2

Determine the solubility of oxygen in water at 25°C exposed to air at 1.0 atm. Assume a partial pressure for oxygen of 0.21 atm.
Concentrations

- Solutions have variable composition.
- To describe a solution, you need to describe the components and their relative amounts.
- The terms dilute and concentrated can be used as qualitative descriptions of the amount of solute in solution.
- Concentration = amount of solute in a given amount of solution.
  - Occasionally amount of solvent

Preparation of a Solution

- Need to know amount of solution and concentration of solution
- Calculate the mass of solute needed
  - Start with amount of solution
  - Use concentration as a conversion factor
    - 5% by mass ⇒ 5 g solute ≡ 100 g solution
  - “Dissolve the grams of solute in enough solvent to total the total amount of solution.”

Solution Concentration: Molarity

- Moles of solute per 1 liter of solution
- Describes how many molecules of solute in each liter of solution
- If a sugar solution concentration is 2.0 M,
  - 1 liter of solution contains 2.0 moles of sugar
  - 2 liters = 4.0 moles sugar
  - 0.5 liters = 1.0 mole sugar

Molarity, M = \frac{\text{moles of solute}}{\text{liters of solution}}

Solution Concentration: Moality, m

- Moles of solute per 1 kilogram of solvent
- Defined in terms of amount of solvent, not solution
  - Like the others
- Does not vary with temperature
  - Because based on masses, not volumes

\text{molality, m} = \frac{\text{moles of solute}}{\text{kg of solvent}}

Concentrations

| TABLE 12.5 Solution Concentration Terms |
|-------------------------------|------------------|
| Unit                     | Definition | Units |
| Molarity (M)              | amount solute (in mol) | mol |
| Molarity (m)              | amount solute (in mol) | l |
| Molarity (mm)             | mass solute (in kg)  | kg |
| Molar fraction (x)        | total amount of solute and solvent (in mol) | mg |
| Molar percent (mol %)     | amount solute (in mol) | % |
| Molar percent (mole %)    | total amount of solute and solvent (in mol) | % |
| Parts by mass             | mass solute | mg solute | multiplication factor |
| Parts by weight            | mass solution | % |
| Parts per million by mass (ppm) | Multiplication factor | ppm |
| Parts per billion by mass (ppb) | Multiplication factor | ppm |
| Parts by volume (%)        | volume solute | volume solution | multiplication factor |

"Multiplication factors for parts by volume are identical to those for parts by mass."
Parts Solute in Parts Solution

- Parts can be measured by mass or volume.
- Parts are generally measured in the same units.
  - By mass in grams, kilogram, lbs, etc.
  - By volume in mL, L, gallons, etc.
  - Mass and volume combined in grams and mL

Parts Solute in Parts Solution

- Percentage = parts of solute in every 100 parts solution
  - If a solution is 0.9% by mass, then there are 0.9 grams of solute in every 100 grams of solution (or 0.9 kg solute in every 100 kg solution).
- Parts per million = parts of solute in every 1 million parts solution
  - If a solution is 36 ppm by volume, then there are 36 mL of solute in 1 million mL of solution.

Parts Solute in Parts Solution

- Grams of solute per 1,000,000 g of solution
- mg of solute per 1 kg of solution
- 1 liter of water = 1 kg of water
  - For aqueous solutions we often approximate the kg of the solution as the kg or L of water.
  - For dilute solutions, the difference in density between the solution and pure water is usually negligible.

Solution Concentrations: Mole Fraction, $X_A$

- The mole fraction is the fraction of the moles of one component in the total moles of all the components of the solution.
- Total of all the mole fractions in a solution = 1.
- Unitless
- The mole percentage is the percentage of the moles of one component in the total moles of all the components of the solution.
  - $= \text{mole fraction} \times 100\%$

PPM

- Grams of solute per 1,000,000 g of solution
- mg of solute per 1 kg of solution
- 1 liter of water = 1 kg of water
  - For aqueous solutions we often approximate the kg of the solution as the kg or L of water.
  - For dilute solutions, the difference in density between the solution and pure water is usually negligible.

Converting Concentration Units

1. Write the given concentration as a ratio.
2. Separate the numerator and denominator.
   - Separate into the solute part and solution part
3. Convert the solute part into the required unit.
4. Convert the solution part into the required unit.
5. Use the definitions to calculate the new concentration units.

Calculating Concentrations

Problem:
62.1 g (1.00 mol) of ethylene glycol is dissolved in 250. g of water.
Calculate mol fraction, molality, and weight % of the solution.
Calculating Concentrations

**Problem:**
62.1 g (1.00 mol) of ethylene glycol is dissolved in 250. g of water.

Mole Fraction:

\[
\begin{align*}
62.1 \text{ g C}_2\text{H}_6\text{O}_2 &\times \frac{1 \text{ mol C}_2\text{H}_6\text{O}_2}{62.1 \text{ g}} = 1.00 \text{ mol C}_2\text{H}_6\text{O}_2 \\
250. \text{ g H}_2\text{O} &\times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} = 13.9 \text{ mol H}_2\text{O}
\end{align*}
\]

Molality:

\[
m_{\text{C}_2\text{H}_6\text{O}_2} = \frac{62.1 \text{ g C}_2\text{H}_6\text{O}_2}{250. \text{ g H}_2\text{O} \times \frac{1 \text{ kg}}{10^3 \text{ g}}} = 4.00 \text{ mol/kg}
\]

Calculating Concentrations

**Problem:**
62.1 g (1.00 mol) of ethylene glycol is dissolved in 250. g of water.

Mole Fraction:

\[
\begin{align*}
62.1 \text{ g C}_2\text{H}_6\text{O}_2 &\times \frac{1 \text{ mol C}_2\text{H}_6\text{O}_2}{62.1 \text{ g}} = 1.00 \text{ mol C}_2\text{H}_6\text{O}_2 \\
250. \text{ g H}_2\text{O} &\times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} = 13.9 \text{ mol H}_2\text{O}
\end{align*}
\]

\[
X_{\text{C}_2\text{H}_6\text{O}_2} = \frac{1.00 \text{ mol}}{1.00 \text{ mol} + 13.9 \text{ mol}} = 0.0671
\]

Calculating Concentrations

**Problem:**
62.1 g (1.00 mol) of ethylene glycol is dissolved in 250. g of water.

Molality:

\[
m_{\text{C}_2\text{H}_6\text{O}_2} = \frac{62.1 \text{ g C}_2\text{H}_6\text{O}_2}{250. \text{ g H}_2\text{O} \times \frac{1 \text{ kg}}{10^3 \text{ g}}} = 4.00 \text{ mol/kg}
\]

Calculating Concentrations

**Problem:**
62.1 g (1.00 mol) of ethylene glycol is dissolved in 250. g of water.

Wt. %:

\[
\%\text{(C}_2\text{H}_6\text{O}_2) = \frac{62.1 \text{ g C}_2\text{H}_6\text{O}_2}{62.1 \text{ g C}_2\text{H}_6\text{O}_2 + 250. \text{ g H}_2\text{O}} \times 100 = 19.9\\
\]

Colligative Properties

**Colligative properties** are properties whose value depends only on the number of solute particles, and not on what they are.

- Value of the property depends on the concentration of the solution.
- The difference in the value of the property between the solution and the pure substance is generally related to the different attractive forces and solute particles occupying solvent molecules positions.
Vapor Pressure of Solutions

- The vapor pressure of a solvent above a solution is lower than the vapor pressure of the pure solvent.
  - The solute particles replace some of the solvent molecules at the surface.
  - The pure solvent establishes a liquid vapor equilibrium.

Addition of a nonvolatile solute reduces the rate of vaporization, decreasing the amount of vapor.

Eventually, equilibrium is re-established, but with a smaller number of vapor molecules; therefore, the vapor pressure will be lower.

- A concentrated solution will draw solvent molecules toward it due to the natural drive for materials in nature to mix.
- Similarly, a concentrated solution will draw pure solvent vapor into it due to this tendency to mix.
- The result is reduction in vapor pressure.

Thirsty Solutions Revisited

- Beakers with equal liquid levels of pure solvent and a solution are placed in a bell jar. Solvent molecules evaporate from each one and fill the bell jar, establishing an equilibrium with the liquids in the beakers.
- When equilibrium is established, the liquid level in the solution beaker is higher than the solution level in the pure solvent beaker; the thirsty solution grabs and holds solvent vapor more effectively.
Raoult’s Law
• The vapor pressure of a volatile solvent above a solution is equal to its normal vapor pressure, $P^o$, multiplied by its mole fraction in the solution.
  \[ P_{\text{solvent in solution}} = \chi_{\text{solvent}} \cdot P^o \]
  – Because the mole fraction is always less than 1, the vapor pressure of the solvent in solution will always be less than the vapor pressure of the pure solvent.

Vapor Pressure Lowering
• The vapor pressure of a solvent in a solution is always lower than the vapor pressure of the pure solvent.
• The vapor pressure of the solution is directly proportional to the amount of the solvent in the solution.
• The difference between the vapor pressure of the pure solvent and the vapor pressure of the solvent in solution is called the vapor pressure lowering.
  \[ \Delta P = P^o_{\text{solvent}} - P_{\text{solution}} = \chi_{\text{solute}} \cdot P^o_{\text{solvent}} \]

Raoult’s Law for Volatile Solute
• When both the solvent and the solute can evaporate, both molecules will be found in the vapor phase.
• The total vapor pressure above the solution will be the sum of the vapor pressures of the solute and solvent.
  – For an ideal solution
    \[ P_{\text{total}} = P_{\text{solute}} + P_{\text{solvent}} \]
• The solvent decreases the solute vapor pressure in the same way the solute decreased the solvent’s.
  \[ P_{\text{solute}} = \chi_{\text{solute}} \cdot P^o_{\text{solute}} \text{ and } P_{\text{solvent}} = \chi_{\text{solvent}} \cdot P^o_{\text{solvent}} \]

Ideal versus Nonideal Solution
• In ideal solutions, the made solute–solvent interactions are equal to the sum of the broken solute–solute and solvent–solvent interactions.
  – Ideal solutions follow Raoult’s law
• Effectively, the solute is diluting the solvent.
• If the solute–solute interactions are stronger or weaker than the broken interactions the solution is nonideal.

Vapor Pressure of a Nonideal Solution
• When the solute–solvent interactions are stronger than the solute–solute + solvent–solvent, the total vapor pressure of the solution will be less than predicted by Raoult’s law, because the vapor pressures of the solute and solvent are lower than ideal.
• When the solute–solvent interactions are weaker than the solute–solute + solvent–solvent, the total vapor pressure of the solution will be more than predicted by Raoult’s Law.

Deviations from Raoult’s Law
Example 12.7 Calculating the Vapor Pressure of a Two-Component Solution

A solution contains 3.95 g of carbon disulfide (CS$_2$) and 2.43 g of acetone (CH$_3$COCH$_3$). The vapor pressures at 35°C of pure carbon disulfide and pure acetone are 515 torr and 332 torr, respectively. Assuming ideal behavior, calculate the vapor pressures of each of the components and the total vapor pressure above the solution. The experimentally measured total vapor pressure of the solution at 35°C is 645 torr. Is the solution ideal? If not, what can you say about the relative strength of carbon disulfide–acetone interactions compared to the acetone–acetone and carbon disulfide–carbon disulfide interactions?

Sort

You are given the masses and vapor pressures of carbon disulfide and acetone and are asked to find the vapor pressures of each component in the mixture and the total pressure assuming ideal behavior.

GIVEN

- Mass of CS$_2$: 3.95 g
- Mass of CH$_3$COCH$_3$: 2.43 g
- Vapor pressure of CS$_2$ at 35°C: 515 torr
- Vapor pressure of CH$_3$COCH$_3$ at 35°C: 332 torr

FIND

- Partial pressures of CS$_2$ and CH$_3$COCH$_3$
- Total pressure (ideal)

Strategy

This problem requires the use of Raoult's law to calculate the partial pressures of each component. In order to use Raoult's law, you first calculate the mole fractions of the two components. Convert the masses of each component to moles and then use the definition of mole fraction to calculate the mole fraction of carbon disulfide.

You can then find the mole fraction of acetone because the mole fractions of the two components add up to 1.

Relationships Used

\[ P = xP^\text{pure} \]

Solve

Begin by converting the mass of each component to the amounts in moles.

Then calculate the mole fraction of carbon disulfide.

Calculate the partial pressures of carbon disulfide and acetone by using Raoult's law and the given values of the vapor pressures of the pure substances.

Calculate the total pressure by summing the partial pressures.

Check

The units of the answer (torr) are correct. The magnitude seems reasonable given the partial pressures of the pure substances.

Solution

\[ P_{\text{tot}}(\text{ideal}) = 285 \text{ torr} + 148 \text{ torr} \]
\[ = 433 \text{ torr} \]

\[ P_{\text{tot}}(\text{exp}) = 645 \text{ torr} \]

\[ P_{\text{tot}}(\text{exp}) > P_{\text{tot}}(\text{ideal}) \]

The solution is not ideal and shows positive deviations from Raoult's law. Therefore, carbon disulfide–acetone interactions must be weaker than acetone–acetone and carbon disulfide–carbon disulfide interactions.

Other Colligative Properties Related to Vapor Pressure Lowering

- Vapor pressure lowering occurs at all temperatures.
- This results in the temperature required to boil the solution being higher than the boiling point of the pure solvent.
- This also results in the temperature required to freeze the solution being lower than the freezing point of the pure solvent.
Freezing Point Depression and Boiling Point Elevation

- Pure water freezes at 0 °C. At this temperature, ice and liquid water are in dynamic equilibrium.
- Adding salt disrupts the equilibrium. The salt particles dissolve in the water, but do not attach easily to the solid ice.
- When an aqueous solution containing a dissolved solid solute freezes slowly, the ice that forms does not normally contain much of the solute.
- To return the system to equilibrium, the temperature must be lowered sufficiently to make the water molecules slow down enough so that more can attach themselves to the ice.

**Freezing Salt Water**

- The freezing point of a solution is lower than the freezing point of the pure solvent.
  - Therefore, the melting point of the solid solution is lower.
- The difference between the freezing point of the solution and freezing point of the pure solvent is directly proportional to the molal concentration of solute particles.
  \[(FP_{\text{solvent}} - FP_{\text{solution}}) = \Delta T_f = m \cdot K_f\]
- The proportionality constant is called the freezing point depression constant, \(K_f\).
  - The value of \(K_f\) depends on the solvent.
  - The units of \(K_f\) are °C/m.

**Example 12.8**

**Given:** 1.7 m C₂H₆O₂ \((aq)\)

**Find:** \(T_f\) °C

**Conceptual Plan:**

\[\Delta T_f = m \cdot K_f\]

**Relationships:**

\[\Delta T_f = m \cdot K_f\] for H₂O = 1.86 °C/m, FP\(_{\text{H₂O}}\) = 0.00

**Solve:**

\[\Delta T_f = m \cdot K_f \cdot \text{H₂O} = 1.7 \cdot 1.86 \cdot \text{°C/m} \cdot \text{c/mol}\]

\[= 3.2^\circ \text{C}\]

**Check:** the unit is correct, the freezing point lower than the normal freezing point makes sense

**Boiling Point Elevation**

- The boiling point of a solution is higher than the boiling point of the pure solvent.
  - For a nonvolatile solute
- The difference between the boiling point of the solution and boiling point of the pure solvent is directly proportional to the molal concentration of solute particles.
  \[(BP_{\text{solution}} - BP_{\text{solvent}}) = \Delta T_b = m \cdot K_b\]
- The proportionality constant is called the boiling point elevation constant, \(K_b\).
  - The value of \(K_b\) depends on the solvent.
  - The units of \(K_b\) are °C/m.

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- The proportionality constant is called the boiling point elevation constant, \(K_b\).
  - The value of \(K_b\) depends on the solvent.
  - The units of \(K_b\) are °C/m.
Example 12.9: How many grams of ethylene glycol, $C_2H_6O_2$, must be added to 1.0 kg $H_2O$ to give a solution that boils at 105 °C?

**Given:** 1.0 kg $H_2O$, $T_b = 105$ °C

**Find:** mass $C_2H_6O_2$, g

**Conceptual Plan:**

**Relationships:**

$\Delta T_b = m \times K_b$

$K_b = 0.512 °C/m$

$BP_{H2O} = 100.0 °C$

$MM_{C_2H_6O_2} = 62.07$ g/mol

$1$ kg = $1000$ g

**Solve:**

$m = \frac{\Delta T_b}{K_b} = \frac{105.0 °C - 100.0 °C}{0.512 °C/m} = 9.77 m$

$1.0$ kg $H_2O \times \frac{1.0 kg H_2O}{1 mol C_2H_6O_2} \times \frac{62.07 g C_2H_6O_2}{1 mol C_2H_6O_2} = 61.0 \times 10^2$ g $C_2H_6O_2$

Molar Mass By Boiling Point Elevation

**Problem:**

Benzyl acetate is one of the active components of oil of jasmine. If 0.125 g of the compound is added to 25.0 g of chloroform (CHCl₃), the boiling point of the solution is 61.82 °C. What is the molar mass of benzyl acetate?

$\Delta T_b = K_b \times m_{solute}$

$m_{solute} = \frac{\Delta T_b}{K_b}$

$0.125$ g benzyl acetate

$25.0$ g CHCl₃

$\frac{8.3 \times 10^{-4}}{mol CHCl₃}$

$0.033$ mol benzyl acetate

$150$ g/mol

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$\frac{8.3 \times 10^{-4}}{mol CHCl₃}$

$0.033$ mol benzyl acetate

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Osmosis

- Osmosis is the flow of solvent from a solution of low concentration into a solution of high concentration.
- The solutions may be separated by a semipermeable membrane.
- A semipermeable membrane allows solvent to flow through it, but not solute.

Osmotic Pressure

- The amount of pressure needed to keep osmotic flow from taking place is called the osmotic pressure.
- The osmotic pressure, $\Pi$, is directly proportional to the molarity of the solute particles.
  - $R = 0.08206$ (atm-L)/(mol-K)
  - $\Pi = MRT$
Example 12.10: What is the molar mass of a protein if 5.87 mg per 10 mL gives an osmotic pressure of 2.45 torr at 25 °C?

Given: 5.87 mg/10 mL, Π = 2.45 torr, T = 25 °C
Find: molar mass, g/mol

Conceptual Plan:

\[ \Pi = \frac{RT}{M} \]
\[ M = \frac{RT}{\Pi} \]

Relationships:

\[ M = \text{mol/L}, \quad 1 \text{ mL} = 0.001 \text{ L}, \quad \text{MM} = \text{g/mol}, \quad 1 \text{ atm} = 760 \text{ torr} \]

Solve:

\[ M = \text{mole/L}, \quad 1 \text{ mL} = 0.001 \text{ L}, \quad \text{MM} = \text{g/mol}, \quad 1 \text{ atm} = 760 \text{ torr} \]
\[ T(K) = 25 + 273.15 = 298 \text{ K} \]

\[ M = \frac{0.08206 \text{ atm} \cdot \text{L/mol} \cdot \text{K}}{2.45 \text{ torr}} = 3.18 \times 10^{-4} \text{ mol/L} \]

\[ \text{MM} = \frac{5.87 \times 10^{-3} \text{ g}}{1.318 \times 10^{-4} \text{ mol}} = 4.5 \times 10^3 \text{ g/mol} \]

Van’t Hoff Factors

- Ionic compounds produce multiple solute particles for each formula unit.
- The theoretical van’t Hoff factor, \( i \), is the ratio of moles of solute particles to moles of formula units dissolved.
- The measured van’t Hoff factors are generally less than the theoretical due to ion pairing in solution.
  - Therefore, the measured van’t Hoff factors often cause the \( \Delta T \) to be lower than you might expect.
Example 12.11: What is the measured van’t Hoff factor if 0.050 m CaCl\(_2\)\(_{\text{aq}}\) has freezing point of −0.27 °C?

\[
\text{Given: } 0.050 \text{ m CaCl}_2(\text{aq}), T_f = -0.27 ^\circ \text{C} \\
\text{Find: } i
\]

Conceptual Plan:

\[
\text{Relationships: } \Delta T_f = i \cdot m \cdot K_f \\
\text{Solve: } \frac{\Delta T_f}{m \cdot K_f} = i
\]

\[
\Delta T_f = 0.050 \text{ m} \cdot 1.86 ^\circ \text{C/m} \\
\Rightarrow i = \frac{0.27^\circ \text{C}}{0.050 \text{ m} \cdot 1.86 ^\circ \text{C/m}} = 2.9
\]

A hyperosmotic solution has a higher osmotic pressure than the solution inside the cell; as a result there is a net flow of water out of the cell, causing it to shrivel.

A hyposmotic solution has a lower osmotic pressure than the solution inside the cell; as a result there is a net flow of water into the cell, causing it to swell.

Mixtures

- Solutions = homogeneous (solute size 0.1nm)
- Suspensions = heterogeneous, separate on standing (solute size Larger than a micron)
- Colloids = heterogeneous, do not separate on standing (solute size 1-1000 nm)
  - Particles can coagulate
  - Cannot pass through semipermeable membrane
  - Hydrophilic
    - Stabilized by attraction for solvent (water)
  - Hydrophobic
    - Stabilized by charged surface repulsions
  - Show the Tyndall effect and Brownian motion.

Brownian Motion
Colloids

- Colloidal dispersions scatter light, a phenomenon known as the "The Tyndall effect."

Types of Colloidal Suspensions

Soaps

- Triglycerides can be broken down into fatty acid salts and glycerol by treatment with a strong hydroxide solution.
- Fatty acid salts have a very polar "head" because it is ionic and a very nonpolar "tail" because it is all C and H.
  - Hydrophilic head and hydrophobic tail
- This unique structure allows the fatty acid salts, called soaps, to help oily substances be attracted to water.
  - Micelle formation
  - Emulsification

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