

## Sample Exercise 13.1 Predicting Solubility Patterns

Predict whether each of the following substances is more likely to dissolve in the nonpolar solvent carbon tetrachloride ( $\text{CCl}_4$ ) or in water:  $\text{C}_7\text{H}_{16}$ ,  $\text{Na}_2\text{SO}_4$ ,  $\text{HCl}$ , and  $\text{I}_2$ .

### Solution

**Analyze** We are given two solvents, one that is nonpolar ( $\text{CCl}_4$ ) and the other that is polar ( $\text{H}_2\text{O}$ ), and asked to determine which will be the better solvent for each solute listed.

**Plan** By examining the formulas of the solutes, we can predict whether they are ionic or molecular. For those that are molecular, we can predict whether they are polar or nonpolar. We can then apply the idea that the nonpolar solvent will be better for the nonpolar solutes, whereas the polar solvent will be better for the ionic and polar solutes.

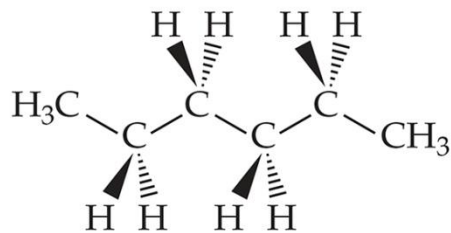
**Solve**  $\text{C}_7\text{H}_{16}$  is a hydrocarbon, so it is molecular and nonpolar.  $\text{Na}_2\text{SO}_4$ , a compound containing a metal and nonmetals, is ionic.  $\text{HCl}$ , a diatomic molecule containing two nonmetals that differ in electronegativity, is polar.  $\text{I}_2$ , a diatomic molecule with atoms of equal electronegativity, is nonpolar. We would therefore predict that  $\text{C}_7\text{H}_{16}$  and  $\text{I}_2$  (the nonpolar solutes) would be more soluble in the nonpolar  $\text{CCl}_4$  than in polar  $\text{H}_2\text{O}$ , whereas water would be the better solvent for  $\text{Na}_2\text{SO}_4$  and  $\text{HCl}$  (the ionic and polar covalent solutes).

## Sample Exercise 13.1 Predicting Solubility Patterns

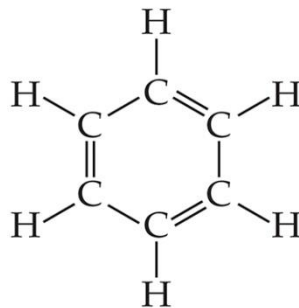
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### Practice Exercise 1

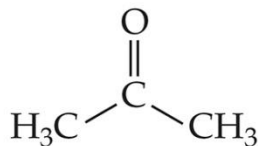
Which of the following solvents will best dissolve wax, which is a complex mixture of compounds that mostly are  $\text{CH}_3\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-}$ ?



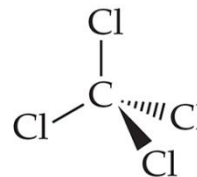
(a) Hexane



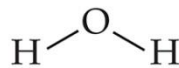
(b) Benzene



(c) Acetone



(d) Carbon tetrachloride



(e) Water



## Sample Exercise 13.2 A Henry's Law Calculation

Calculate the concentration of  $\text{CO}_2$  in a soft drink that is bottled with a partial pressure of  $\text{CO}_2$  of 4.0 atm over the liquid at 25 °C. The Henry's law constant for  $\text{CO}_2$  in water at this temperature is  $3.4 \times 10^{-2}$  mol/L-atm.

### Solution

**Analyze** We are given the partial pressure of  $\text{CO}_2$ ,  $P_{\text{CO}_2}$ , and the Henry's law constant,  $k$ , and asked to calculate the concentration of  $\text{CO}_2$  in the solution.

**Plan** With the information given, we can use Henry's law, Equation 13.4, to calculate the solubility,  $S_{\text{CO}_2}$ .

**Solve** 
$$S_{\text{CO}_2} = kP_{\text{CO}_2} = (3.4 \times 10^{-2} \text{ mol/L-atm})(4.0 \text{ atm})$$
$$= 0.14 \text{ mol/L} = 0.14 \text{ M}$$

**Check** The units are correct for solubility, and the answer has two significant figures consistent with both the partial pressure of  $\text{CO}_2$  and the value of the Henry's law constant.

### Practice Exercise 1

You double the partial pressure of a gas over a liquid at constant temperature. Which of these statements is then true?

- (a) The Henry's law constant is doubled.
- (b) The Henry's law constant is decreased by half.
- (c) There are half as many gas molecules in the liquid.
- (d) There are twice as many gas molecules in the liquid.
- (e) There is no change in the number of gas molecules in the liquid.

## Sample Exercise 13.2 A Henry's Law Calculation

Continued

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### Practice Exercise 2

Calculate the concentration of  $\text{CO}_2$  in a soft drink after the bottle is opened and the solution equilibrates at  $25\text{ }^\circ\text{C}$  under a  $\text{CO}_2$  partial pressure of  $3.0 \times 10^{-4}\text{ atm}$ .

## Sample Exercise 13.3 Calculation of Mass-Related Concentrations

- (a) A solution is made by dissolving 13.5 g of glucose ( $C_6H_{12}O_6$ ) in 0.100 kg of water. What is the mass percentage of solute in this solution?
- (b) A 2.5-g sample of groundwater was found to contain  $5.4 \mu\text{g}$  of  $Zn^{2+}$ . What is the concentration of  $Zn^{2+}$  in parts per million?

### Solution

**(a) Analyze** We are given the number of grams of solute (13.5 g) and the number of grams of solvent (0.100 kg = 100 g). From this, we must calculate the mass percentage of solute.

**Plan** We can calculate the mass percentage by using Equation 13.5. The mass of the solution is the sum of the mass of solute (glucose) and the mass of solvent (water).

**Solve** Mass % of glucose =  $\frac{\text{mass glucose}}{\text{mass soln}} \times 100$

$$= \frac{13.5 \text{ g}}{(13.5 + 100) \text{ g}} \times 100 = 11.9\%$$

**Comment** The mass percentage of water in this solution is  $(100 - 11.9)\% = 88.1\%$ .

**(b) Analyze** In this case we are given the number of micrograms of solute. Because  $1 \mu\text{g}$  is  $1 \times 10^{-6} \text{ g}$ ,  $5.4 \mu\text{g} = 5.4 \times 10^{-6} \text{ g}$ .

**Plan** We calculate the parts per million using Equation 13.6.

## Sample Exercise 13.3 Calculation of Mass-Related Concentrations

Continued

**Solve**

$$\begin{aligned}\text{ppm} &= \frac{\text{mass of solute}}{\text{mass of soln}} \times 10^6 \\ &= \frac{5.4 \times 10^{-6} \text{ g}}{2.5 \text{ g}} \times 10^6 = 2.2 \text{ ppm}\end{aligned}$$

### Practice Exercise 1

Calculate the mass percentage of NaCl in a solution containing 1.50 g of NaCl in 50.0 g of water.

(a) 0.0291% (b) 0.0300% (c) 0.0513% (d) 2.91% (e) 3.00%

### Practice Exercise 2

A commercial bleaching solution contains 3.62% by mass of sodium hypochlorite, NaOCl. What is the mass of NaOCl in a bottle containing 2.50 kg of bleaching solution?

## Sample Exercise 13.4 Calculation of Molality

A solution is made by dissolving 4.35 g of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) in 25.0 mL of water at 25 °C. Calculate the molality of glucose in the solution. Water has a density of 1.00 g/mL.

### Solution

**Analyze** We are asked to calculate a solution concentration in units of molality. To do this, we must determine the number of moles of solute (glucose) and the number of kilograms of solvent (water).

**Plan** We use the molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$  to convert grams of glucose to moles. We use the density of water to convert milliliters of water to kilograms. The molality equals the number of moles of solute (glucose) divided by the number of kilograms of solvent (water).

### Solve

Use the molar mass of glucose, 180.2 g/mol, to convert grams to moles:

$$\text{Mol C}_6\text{H}_{12}\text{O}_6 = (4.35 \text{ g C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.2 \text{ g C}_6\text{H}_{12}\text{O}_6} \right) = 0.0241 \text{ mol C}_6\text{H}_{12}\text{O}_6$$

Because water has a density of 1.00 g/mL, the mass of the solvent is:

$$(25.0 \text{ mL})(1.00 \text{ g/mL}) = 25.0 \text{ g} = 0.0250 \text{ kg}$$



## Sample Exercise 13.4 Calculation of Molality

Continued

Finally, use Equation 13.9 to obtain the molality:

$$\text{Molality of C}_6\text{H}_{12}\text{O}_6 = \frac{0.0241 \text{ mol C}_6\text{H}_{12}\text{O}_6}{0.0250 \text{ kg H}_2\text{O}} = 0.964 m$$

### Practice Exercise 1

Suppose you take a solution and add more solvent, so that the original mass of solvent is doubled. You take this new solution and add more solute, so that the original mass of the solute is doubled. What happens to the molality of the final solution, compared to the original molality?

- (a) It is doubled.
- (b) It is decreased by half.
- (c) It is unchanged.
- (d) It will increase or decrease depending on the molar mass of the solute.
- (e) There is no way to tell without knowing the molar mass of the solute.

### Practice Exercise 2

What is the molality of a solution made by dissolving 36.5 g of naphthalene ( $\text{C}_{10}\text{H}_8$ ) in 425 g of toluene ( $\text{C}_7\text{H}_8$ )?

## Sample Exercise 13.5 Calculation of Mole Fraction and Molality

An aqueous solution of hydrochloric acid contains 36% HCl by mass. (a) Calculate the mole fraction of HCl in the solution. (b) Calculate the molality of HCl in the solution.

### Solution

**Analyze** We are asked to calculate the concentration of the solute, HCl, in two related concentration units, given only the percentage by mass of the solute in the solution.

**Plan** In converting concentration units based on the mass or moles of solute and solvent (mass percentage, mole fraction, and molality), it is useful to assume a certain total mass of solution. Let's assume that there is exactly 100 g of solution. Because the solution is 36% HCl, it contains 36 g of HCl and  $(100 - 36) \text{ g} = 64 \text{ g}$  of  $\text{H}_2\text{O}$ . We must convert grams of solute (HCl) to moles to calculate either mole fraction or molality. We must convert grams of solvent ( $\text{H}_2\text{O}$ ) to moles to calculate mole fractions and to kilograms to calculate molality.

### Solve

(a) To calculate the mole fraction of HCl, we convert the masses of HCl and  $\text{H}_2\text{O}$  to moles and then use Equation 13.7:

$$\text{Moles HCl} = (36 \text{ g HCl}) \left( \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} \right) = 0.99 \text{ mol HCl}$$

$$\text{Moles H}_2\text{O} = (64 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g H}_2\text{O}} \right) = 3.6 \text{ mol H}_2\text{O}$$

$$X_{\text{HCl}} = \frac{\text{moles HCl}}{\text{moles H}_2\text{O} + \text{moles HCl}} = \frac{0.99}{3.6 + 0.99} = \frac{0.99}{4.6} = 0.22$$

## Sample Exercise 13.5 Calculation of Mole Fraction and Molality

Continued

(b) To calculate the molality of HCl in the solution, we use Equation 13.9. We calculated the number of moles of HCl in part (a), and the mass of solvent is 64 g = 0.064 kg:

$$\text{Molality of HCl} = \frac{0.99 \text{ mol HCl}}{0.064 \text{ kg H}_2\text{O}} = 15 \text{ } m$$

Notice that we can't readily calculate the molarity of the solution because we don't know the volume of the 100 g of solution.

### Practice Exercise 1

The solubility of oxygen gas in water at 40 °C is 1.0 mmol per liter of solution. What is this concentration in units of mole fraction?

- (a)  $1.00 \times 10^{-6}$  (b)  $1.80 \times 10^{-5}$  (c)  $1.00 \times 10^{-2}$   
(d)  $1.80 \times 10^{-2}$  (e)  $5.55 \times 10^{-2}$

### Practice Exercise 2

A commercial bleach solution contains 3.62% by mass of NaOCl in water. Calculate (a) the mole fraction and (b) the molality of NaOCl in the solution.

## Sample Exercise 13.6 Calculation of Molarity Using the Density of the Solution

A solution with a density of 0.876 g/mL contains 5.0 g of toluene ( $C_7H_8$ ) and 225 g of benzene. Calculate the molarity of the solution.

### Solution

**Analyze** Our goal is to calculate the molarity of a solution, given the masses of solute (5.0 g) and solvent (225 g) and the density of the solution (0.876 g/mL).

**Plan** The molarity of a solution is the number of moles of solute divided by the number of liters of solution (Equation 13.8). The number of moles of solute ( $C_7H_8$ ) is calculated from the number of grams of solute and its molar mass. The volume of the solution is obtained from the mass of the solution (mass of solution = mass of solute + mass of solvent = 5.0 g + 225 g = 230 g) and its density.

### Solve

The number of moles of solute is:

$$\text{Moles } C_7H_8 = (5.0 \text{ g } C_7H_8) \left( \frac{1 \text{ mol } C_7H_8}{92 \text{ g } C_7H_8} \right) = 0.054 \text{ mol}$$

The density of the solution is used to convert the mass of the solution to its volume:

$$\text{Milliliters soln} = (230 \text{ g}) \left( \frac{1 \text{ mL}}{0.876 \text{ g}} \right) = 263 \text{ mL}$$

## Sample Exercise 13.6 Calculation of Molarity Using the Density of the Solution

Continued

Molarity is moles of solute per liter of solution:

$$\text{Molarity} = \left( \frac{\text{moles C}_7\text{H}_8}{\text{liter soln}} \right) = \left( \frac{0.054 \text{ mol C}_7\text{H}_8}{263 \text{ mL soln}} \right) \left( \frac{1000 \text{ mL soln}}{1 \text{ L soln}} \right) = 0.21 \text{ M}$$

**Check** The magnitude of our answer is reasonable. Rounding moles to 0.05 and liters to 0.25 gives a molarity of  $(0.05 \text{ mol})/(0.25 \text{ L}) = 0.2 \text{ M}$ .

The units for our answer (mol/L) are correct, and the answer, 0.21, has two significant figures, corresponding to the number of significant figures in the mass of solute (2).

**Comment** Because the mass of the solvent (0.225 kg) and the volume of the solution (0.263) are similar in magnitude, the molarity and molality are also similar in magnitude:  $(0.054 \text{ mol C}_7\text{H}_8)/(0.225 \text{ kg solvent}) = 0.24 \text{ m}$ .

### Practice Exercise 1

Maple syrup has a density of 1.325 g/mL, and 100.00 g of maple syrup contains 67 mg of calcium in the form of  $\text{Ca}^{2+}$  ions. What is the molarity of calcium in maple syrup?

(a) 0.017 M (b) 0.022 M (c) 0.89 M (d) 12.6 M (e) 45.4 M

## Sample Exercise 13.6 Calculation of Molarity Using the Density of the Solution

Continued

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### Practice Exercise 2

A solution containing equal masses of glycerol ( $\text{C}_3\text{H}_8\text{O}_3$ ) and water has a density of 1.10 g/mL. Calculate (a) the molality of glycerol, (b) the mole fraction of glycerol, (c) the molarity of glycerol in the solution.


## Sample Integrative Exercise Putting Concepts Together

A 0.100-L solution is made by dissolving 0.441 g of  $\text{CaCl}_2(s)$  in water. (a) Calculate the osmotic pressure of this solution at  $27^\circ\text{C}$ , assuming that it is completely dissociated into its component ions. (b) The measured osmotic pressure of this solution is 2.56 atm at  $27^\circ\text{C}$ . Explain why it is less than the value calculated in (a), and calculate the van't Hoff factor,  $i$ , for the solute in this solution. (c) The enthalpy of solution for  $\text{CaCl}_2$  is  $\Delta H = -81.3 \text{ kJ/mol}$ . If the final temperature of the solution is  $27^\circ\text{C}$ , what was its initial temperature? (Assume that the density of the solution is 1.00 g/mL, that its specific heat is 4.18 J/g-K, and that the solution loses no heat to its surroundings.)

### Solution

(a) The osmotic pressure is given by Equation 13.14,  $\Pi = iMRT$ . We know the temperature,  $T = 27^\circ\text{C} = 300 \text{ K}$ , and the gas constant,  $R = 0.0821 \text{ L-atm/mol-K}$ . We can calculate the molarity of the solution from the mass of  $\text{CaCl}_2$  and the volume of the solution:

$$\begin{aligned}\text{Molarity} &= \left( \frac{0.441 \text{ g } \cancel{\text{CaCl}_2}}{0.100 \text{ L}} \right) \left( \frac{1 \text{ mol } \text{CaCl}_2}{110 \text{ g } \cancel{\text{CaCl}_2}} \right) \\ &= 0.0397 \text{ mol } \text{CaCl}_2/\text{L}\end{aligned}$$

Soluble ionic compounds are strong electrolytes.  (Sections 4.1 and 4.3) Thus,  $\text{CaCl}_2$  consists of metal cations ( $\text{Ca}^{2+}$ ) and nonmetal anions ( $\text{Cl}^-$ ). When completely dissociated, each  $\text{CaCl}_2$  unit forms three ions (one  $\text{Ca}^{2+}$  and two  $\text{Cl}^-$ ). Hence, the calculated osmotic pressure is:

$$\begin{aligned}\Pi &= iMRT = (3)(0.0397 \text{ mol/L})(0.0821 \text{ L-atm/mol-K})(300 \text{ K}) \\ &= 2.93 \text{ atm}\end{aligned}$$

## Sample Integrative Exercise Putting Concepts Together

Continued

- (b) The actual values of colligative properties of electrolytes are less than those calculated because the electrostatic interactions between ions limit their independent movements. In this case, the van't Hoff factor, which measures the extent to which electrolytes actually dissociate into ions, is given by:

$$\begin{aligned}i &= \frac{\Pi(\text{measured})}{\Pi(\text{calculated for nonelectrolyte})} \\ &= \frac{2.56 \text{ atm}}{(0.0397 \text{ mol/L})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(300 \text{ K})} = 2.62\end{aligned}$$

Thus, the solution behaves as if the  $\text{CaCl}_2$  has dissociated into 2.62 particles instead of the ideal 3.

- (c) If the solution is 0.0397 M in  $\text{CaCl}_2$  and has a total volume of 0.100 L, the number of moles of solute is:

$$(0.100 \text{ L})(0.0397 \text{ mol/L}) = 0.00397 \text{ mol}$$

Hence, the quantity of heat generated in forming the solution is:

$$(0.00397 \text{ mol})(-81.3 \text{ kJ/mol}) = -0.323 \text{ kJ}$$

The solution absorbs this heat, causing its temperature to increase. The relationship between temperature change and heat is given by Equation 5.22:

$$q = (\text{specific heat})(\text{grams})(\Delta T)$$



## Sample Integrative Exercise Putting Concepts Together

Continued

The heat absorbed by the solution is  $q = +0.323 \text{ kJ} = 323 \text{ J}$ . The mass of the 0.100 L of solution is  $(100 \text{ mL})(1.00 \text{ g/mL}) = 100 \text{ g}$  (to three significant figures). Thus, the temperature change is:

$$\begin{aligned}\Delta T &= \frac{q}{(\text{specific heat of solution})(\text{grams of solution})} \\ &= \frac{323 \text{ J}}{(4.18 \text{ J/g}\cdot\text{K})(100 \text{ g})} = 0.773 \text{ K}\end{aligned}$$

A kelvin has the same size as a degree Celsius.  (Section 1.4) Because the solution temperature increases by  $0.773 \text{ }^\circ\text{C}$ , the initial temperature was:

$$27.0 \text{ }^\circ\text{C} - 0.773 \text{ }^\circ\text{C} = 26.2 \text{ }^\circ\text{C}$$