Chemistry 2e
11: Solutions and Colloids
11.1: The Dissolution Process

1. How do solutions differ from compounds? From other mixtures?

Solution

A solution can vary in composition, while a compound cannot vary in composition. Solutions are homogeneous at the molecular level, while other mixtures are heterogeneous.

2. Which of the principal characteristics of solutions are evident in the solutions of $K_2Cr_2O_7$ shown in Figure 11.2?

Solution

The solutions are the same throughout (the color is constant throughout), and the composition of a solution of $K_2Cr_2O_7$ in water can vary.

3. When KNO_3 is dissolved in water, the resulting solution is significantly colder than the water was originally.

(a) Is the dissolution of KNO_3 an endothermic or an exothermic process?

(b) What conclusions can you draw about the intermolecular attractions involved in the process?

(c) Is the resulting solution an ideal solution?

Solution

(a) The process is endothermic as the solution is consuming heat. (b) Attraction between the K^+ and NO_3^- ions is stronger than between the ions and water molecules (the ion-ion interactions have a lower, more negative energy). Therefore, the dissolution process increases the energy of the molecular interactions, and it consumes the thermal energy of the solution to make up for the difference. (c) No, an ideal solution is formed with no appreciable heat release or consumption.

4. Give an example of each of the following types of solutions:

(a) a gas in a liquid

(b) a gas in a gas

(c) a solid in a solid

Solution

(a) CO_2 in water; (b) O_2 in N_2 (air); (c) bronze (solution of tin or other metals in copper)

5. Indicate the most important types of intermolecular attractions in each of the following solutions:

(a) The solution in Figure 11.2

(b) $NO(g)$ in $CO(l)$

(c) $Cl_2(g)$ in $Br_2(l)$

(d) $HCl(g)$ in benzene $C_6H_6(l)$

(e) Methanol $CH_3OH(l)$ in $H_2O(l)$

Solution

(a) ion-dipole forces; (b) dipole-dipole forces; (c) dispersion forces; (d) dispersion forces; (e) hydrogen bonding

6. Predict whether each of the following substances would be more soluble in water (polar solvent) or in a hydrocarbon such as heptane (C_7H_{16}, nonpolar solvent):

(a) vegetable oil (nonpolar)

(b) isopropyl alcohol (polar)

(c) potassium bromide (ionic)

Solution

(a) heptane; (b) water; (c) water

7. Heat is released when some solutions form; heat is absorbed when other solutions form. Provide a molecular explanation for the difference between these two types of spontaneous processes.

Solution

Heat is released when the total intermolecular forces (IMFs) between the solute and solvent molecules are stronger than the total IMFs in the pure solute and in the pure solvent: Breaking weaker IMFs and forming stronger IMFs releases heat. Heat is absorbed when the total IMFs in the solution are weaker than the total of those in the pure solute and in the pure solvent: Breaking stronger IMFs and forming weaker IMFs absorbs heat.

8. Solutions of hydrogen in palladium may be formed by exposing Pd metal to H_2 gas. The concentration of hydrogen in the palladium depends on the pressure of H_2 gas applied, but in a more complex fashion than can be described by Henry's law. Under certain conditions, 0.94 g of hydrogen gas is dissolved in 215 g of palladium metal (solution density = 10.8 g/cm^3).

(a) Determine the molarity of this solution.

(b) Determine the molality of this solution.

(c) Determine the percent by mass of hydrogen atoms in this solution.

Solution

(a) $\text{mol H} = \dfrac{0.94\text{ g}}{2.0158\text{ g mol}^{-1}} = 0.4633\text{ mol}$

$$\text{liters of solution} = \frac{(215 + 0.94)\text{ g}}{10.8\text{ g cm}^{-3}} = \frac{215.94}{10.8\text{ cm}^{-3}} = 20\text{ cm}^3$$

$$M = \frac{0.4633\text{ mol}}{0.020\text{ L}} = 23\ M$$

(b) $m = \dfrac{0.4633\text{ mol}}{0.215\text{ kg}} = 2.16\ m$

(c) $\%\text{H} = \dfrac{0.94\text{ g}}{(215 + 0.94)\text{ g}} \times 100 = 0.44\%$

Chemistry 2e
11: Solutions and Colloids
11.2: Electrolytes

9. Explain why the ions Na^+ and Cl^- are strongly solvated in water but not in hexane, a solvent composed of nonpolar molecules.

Solution

Crystals of NaCl dissolve in water, a polar liquid with a very large dipole moment, and the individual ions become strongly solvated. Hexane is a nonpolar liquid with a dipole moment of zero and, therefore, does not significantly interact with the ions of the NaCl crystals.

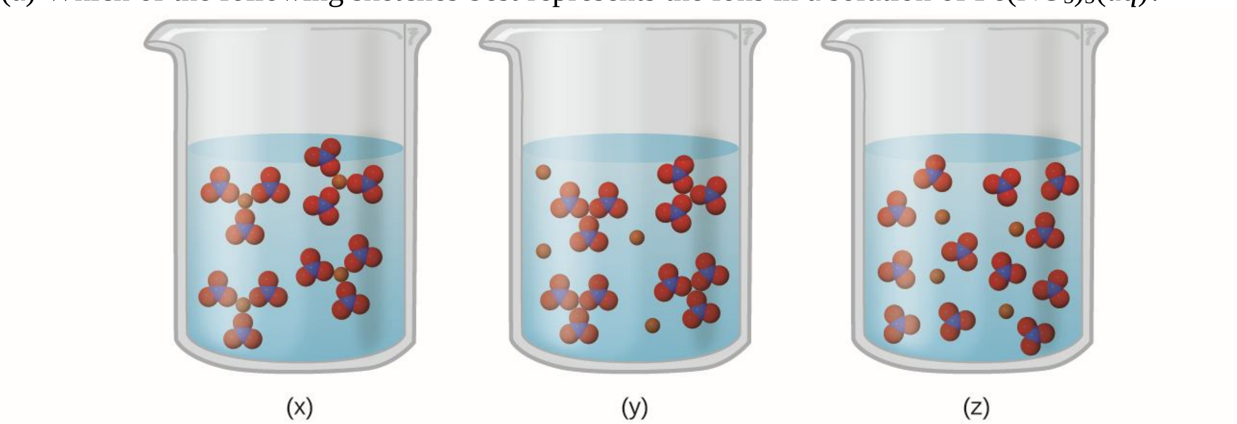
10. Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.

Solution

HBr is an acid and so its molecules react with water molecules to form H_3O^+ and Br^- ions that provide conductivity to the solution. Though HBr is soluble in benzene, it does not react chemically but remains dissolved as neutral HBr molecules. With no ions present in the benzene solution, it is electrically nonconductive.

11. Consider the solutions presented:

(a) Which of the following sketches best represents the ions in a solution of $Fe(NO_3)_3(aq)$?



(b) Write a balanced chemical equation showing the products of the dissolution of $Fe(NO_3)_3$.

Solution

(a) $Fe(NO_3)_3$ is a strong electrolyte, thus it should completely dissociate into Fe^{3+} and NO_3^- ions. Therefore, (z) best represents the solution. (b)

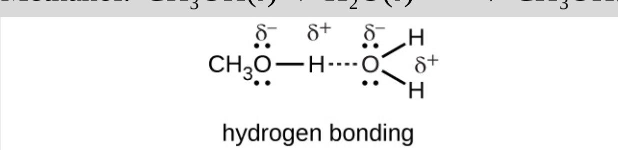
$$Fe(NO_3)_3(s) \longrightarrow Fe^{3+}(aq) + 3NO_3^-(aq)$$

12. Compare the processes that occur when methanol (CH_3OH), hydrogen chloride (HCl), and sodium hydroxide (NaOH) dissolve in water. Write equations and prepare sketches showing the form in which each of these compounds is present in its respective solution.

Solution

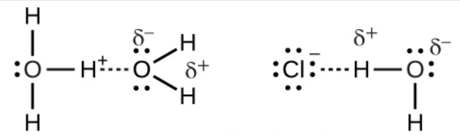
Methanol, CH_3OH, dissolves in water in all proportions, interacting via hydrogen bonding.

Methanol: $CH_3OH(l) + H_2O(l) \longrightarrow CH_3OH(aq)$



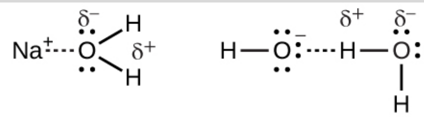
Hydrogen chloride, HCl, dissolves in and reacts with water to yield hydronium cations and chloride anions that are solvated by strong ion-dipole interactions.

Hydrogen chloride: $HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$



Sodium hydroxide, NaOH, dissolves in water and dissociates to yield sodium cations and hydroxide anions that are strongly solvated by ion-dipole interactions and hydrogen bonding, respectively.

Sodium hydroxide: $NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq)$



13. What is the expected electrical conductivity of the following solutions?
(a) NaOH(*aq*)
(b) HCl(*aq*)
(c) $C_6H_{12}O_6(aq)$ (glucose)
(d) $NH_3(aq)$

Solution
(a) high conductivity (solute is an ionic compound that will dissociate when dissolved); (b) high conductivity (solute is a strong acid and will ionize completely when dissolved); (c) nonconductive (solute is a covalent compound, neither acid nor base, unreactive towards water); (d) low conductivity (solute is a weak base and will partially ionize when dissolved)

14. Why are most *solid* ionic compounds electrically nonconductive, whereas aqueous solutions of ionic compounds are good conductors? Would you expect a *liquid* (molten) ionic compound to be electrically conductive or nonconductive? Explain.

Solution
A medium must contain freely mobile, charged entities to be electrically conductive. The ions present in a typical ionic solid are immobilized in a crystalline lattice and so the solid is not able to support an electrical current. When the ions are mobilized, either by melting the solid or dissolving it in water to dissociate the ions, current may flow and these forms of the ionic compound are conductive.

15. Indicate the most important type of intermolecular attraction responsible for solvation in each of the following solutions:
(a) the solutions in Figure 11.7
(b) methanol, CH_3OH, dissolved in ethanol, C_2H_5OH
(c) methane, CH_4, dissolved in benzene, C_6H_6
(d) the polar halocarbon CF_2Cl_2 dissolved in the polar halocarbon $CF_2ClCFCl_2$
(e) $O_2(l)$ in $N_2(l)$

Solution
(a) ion-dipole; (b) hydrogen bonds; (c) dispersion forces; (d) dipole-dipole attractions; (e) dispersion forces

Chemistry 2e
11: Solutions and Colloids
11.3: Solubility

16. Suppose you are presented with a clear solution of sodium thiosulfate, $Na_2S_2O_3$. How could you determine whether the solution is unsaturated, saturated, or supersaturated?

Solution

Add a small crystal of $Na_2S_2O_3$. It will dissolve in an unsaturated solution, remain apparently unchanged in a saturated solution, or initiate precipitation in a supersaturated solution.

17. Supersaturated solutions of most solids in water are prepared by cooling saturated solutions. Supersaturated solutions of most gases in water are prepared by heating saturated solutions. Explain the reasons for the difference in the two procedures.

Solution

The solubility of solids usually decreases upon cooling a solution, while the solubility of gases usually decreases upon heating.

18. Suggest an explanation for the observations that ethanol, C_2H_5OH, is completely miscible with water and that ethanethiol, C_2H_5SH, is soluble only to the extent of 1.5 g per 100 mL of water.

Solution

The hydrogen bonds between water and C_2H_5OH are much stronger than the intermolecular attractions between water and C_2H_5SH.

19. Calculate the percent by mass of KBr in a saturated solution of KBr in water at 10 °C. See Figure 11.16 for useful data, and report the computed percentage to one significant digit.

Solution

At 10 °C, the solubility of KBr in water is approximately 60 g per 100 g of water.

$$\%\ \text{KBr} = \frac{60\ \text{g KBr}}{(60 + 100)\ \text{g solution}} = 40\%$$

20. Which of the following gases is expected to be most soluble in water? Explain your reasoning.
(a) CH_4
(b) CCl_4
(c) $CHCl_3$

Solution

(c) $CHCl_3$ is expected to be most soluble in water. Of the three gases, only this one is polar and thus capable of experiencing relatively strong dipole-dipole attraction to water molecules.

21. At 0 °C and 1.00 atm, as much as 0.70 g of O_2 can dissolve in 1 L of water. At 0 °C and 4.00 atm, how many grams of O_2 dissolve in 1 L of water?

Solution

This problem requires the application of Henry's law. The governing equation is $C_g = kP_g$.

$$k = \frac{C_g}{P_g} = \frac{0.70\ \text{g}}{1.00\ \text{atm}} = 0.70\ \text{g atm}^{-1}$$

Under the new conditions, $C_g = 0.70\ \text{g atm}^{-1} \times 4.00\ \text{atm} = 2.80\ \text{g}$.

22. Refer to Figure 11.10.
(a) How did the concentration of dissolved CO_2 in the beverage change when the bottle was opened?

(b) What caused this change?
(c) Is the beverage unsaturated, saturated, or supersaturated with CO_2?

Solution

(a) It decreased as some of the CO_2 gas left the solution (evidenced by effervescence). (b) Opening the bottle released the high-pressure CO_2 gas above the beverage. The reduced CO_2 gas pressure, per Henry's law, lowers the solubility for CO_2. (c) The dissolved CO_2 concentration will continue to slowly decrease until equilibrium is reestablished between the beverage and the very low CO_2 gas pressure in the opened bottle. Immediately after opening, the beverage, therefore, contains dissolved CO_2 at a concentration greater than its solubility, a nonequilibrium condition, and is said to be supersaturated.

23. The Henry's law constant for CO_2 is 3.4×10^{-2} *M*/atm at 25 °C. Assuming ideal solution behavior, what pressure of carbon dioxide is needed to maintain a CO_2 concentration of 0.10 *M* in a can of lemon-lime soda?

Solution

$$P_g = \frac{C_g}{k} = \frac{0.10\ M}{3.4 \times 10^{-2}\ M/\text{atm}} = 2.9\ \text{atm}$$

24. The Henry's law constant for O_2 is 1.3×10^{-3} *M*/atm at 25 °C. Assuming ideal solution behavior, what mass of oxygen would be dissolved in a 40-L aquarium at 25 °C, assuming an atmospheric pressure of 1.00 atm, and that the partial pressure of O_2 is 0.21 atm?

Solution

$C_g = kP_g$. $C(O_2) = 1.3 \times 10^{-3}$ *M*/atm × 0.21 atm = 2.7×10^{-4} mol/L. The total amount is 2.7×10^{-4} mol/L × 40 L = 1.08×10^{-2} mol. The mass of oxygen is 1.08×10^{-2} mol × 32.0 g/mol = 0.346 g or, using two significant figures, 0.35 g.

25. Assuming ideal solution behavior, how many liters of HCl gas, measured at 30.0 °C and 745 torr, are required to prepare 1.25 L of a 3.20-*M* solution of hydrochloric acid?

Solution

First, calculate the moles of HCl needed. Then use the ideal gas law to find the volume required.

$M = \text{mol L}^{-1}$

$$3.20\ M = \frac{x\ \text{mol}}{1.25\ \text{L}}$$

x = 4.00 mol HCl

Before using the ideal gas law, change pressure to atmospheres and convert temperature from °C to kelvins.

$$\frac{1\ \text{atm}}{x} = \frac{760\ \text{torr}}{745\ \text{torr}}$$

x = 0.9803 atm

$$V = \frac{nRT}{P} = \frac{(4.000\ \cancel{\text{mol}}\ \text{HCl})(0.08206\ \text{L}\ \cancel{\text{atm}}\ \cancel{\text{K}^{-1}}\ \cancel{\text{mol}^{-1}})(303.15\ \cancel{\text{K}})}{0.9803\ \cancel{\text{atm}}} = 102\ \text{L HCl}$$

Chemistry 2e
11: Solutions and Colloids
11.4: Colligative Properties

26. Which is/are part of the macroscopic domain of solutions and which is/are part of the microscopic domain: boiling point elevation, Henry's law, hydrogen bond, ion-dipole attraction, molarity, nonelectrolyte, nonstoichiometric compound, osmosis, solvated ion?

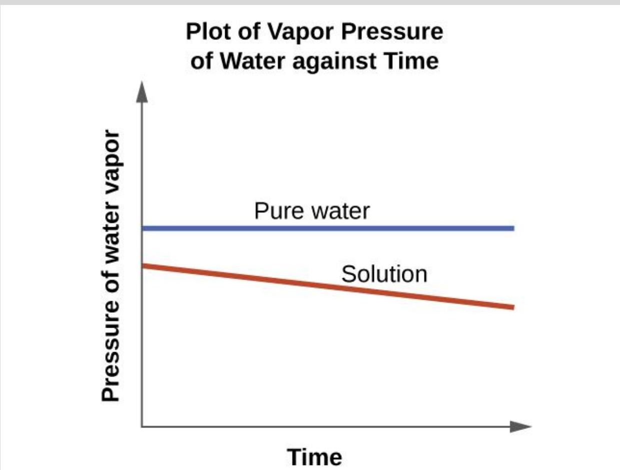
Solution
Macroscopic: boiling point elevation, Henry's law, molarity, nonelectrolyte, nonstoichiometric compound, osmosis. Microscopic: hydrogen bond, ion-dipole attraction, solvated ion

27. What is the microscopic explanation for the macroscopic behavior illustrated in Figure 11.14?

Solution
The strength of the bonds between like molecules is stronger than the strength between unlike molecules. Therefore, some regions will exist in which the water molecules will exclude oil molecules and other regions will exist in which oil molecules will exclude water molecules, forming a heterogeneous region.

28. Sketch a qualitative graph of the pressure versus time for water vapor above a sample of pure water and a sugar solution, as the liquids evaporate to half their original volume.

Solution



29. A solution of potassium nitrate, an electrolyte, and a solution of glycerin ($C_3H_5(OH)3$), a nonelectrolyte, both boil at 100.3 °C. What other physical properties of the two solutions are identical?

Solution
Both form homogeneous solutions; their boiling point elevations are the same, as are their lowering of vapor pressures. Osmotic pressure and the lowering of the freezing point are also the same for both solutions.

30. What are the mole fractions of H_3PO_4 and water in a solution of 14.5 g of H_3PO_4 in 125 g of water?
(a) Outline the steps necessary to answer the question.
(b) Answer the question.

Solution
(a) Determine the number of moles of each component. Then add the number of moles of components and divide that number into the moles of the component whose percentage is desired. (b)

$$\text{mol } H_3PO_4 = \frac{14.5\text{ g}}{97.9952\text{ g mol}^{-1}} = 0.148\text{ mol}$$

$$\text{mol } H_2O = \frac{125\text{ g}}{18.0153\text{ g mol}^{-1}} = 6.939\text{ mol}$$

Total number of moles = 0.148 mol + 6.939 mol = 7.087 mol

$$X_{H_3PO_4} = \frac{0.148\text{ mol}}{7.087\text{ mol}} = 0.0209$$

$$X_{H_2O} = \frac{6.939\text{ mol}}{7.087\text{ mol}} = 0.979$$

31. What are the mole fractions of HNO_3 and water in a concentrated solution of nitric acid (68.0% HNO_3 by mass)?
(a) Outline the steps necessary to answer the question.
(b) Answer the question.

Solution
(a) Find number of moles of HNO_3 and H_2O in 100 g of the solution. Find the mole fractions for the components.
(b) The number of moles of HNO_3 is $\frac{68\text{ g}}{63.01\text{ g/mol}} = 1.079\text{ mol}$. The number of moles of water is $\frac{32\text{ g}}{18.015\text{ g/mol}} = 1.776\text{ mol}$. The mole fraction of HNO_3 is $\frac{1.079}{(1.079 + 1.776)} = 0.378$. The mole fraction of H_2O is 1 − 0.378 = 0.622.

32. Calculate the mole fraction of each solute and solvent:
(a) 583 g of H_2SO_4 in 1.50 kg of water—the acid solution used in an automobile battery
(b) 0.86 g of NaCl in 1.00 × 10^2 g of water—a solution of sodium chloride for intravenous injection
(c) 46.85 g of codeine, $C_{18}H_{21}NO_3$, in 125.5 g of ethanol, C_2H_5OH
(d) 25 g of I_2 in 125 g of ethanol, C_2H_5OH

Solution
(a)

$$\text{mol } H_2SO_4 = \frac{583\text{ g}}{98.079\text{ g mol}^{-1}} = 5.94\text{ mol}$$

$$\text{mol } H_2O = \frac{1.5\times10^3\text{ g}}{18.0153\text{ g mol}^{-1}} = 83.3\text{ mol}$$

Total number of moles = 5.94 mol + 83.3 mol = 89.24 mol

$$X_{H_2SO_4} = \frac{83.3\text{ mol}}{89.24\text{ mol}} = 0.933$$

$$X_{H_2O} = \frac{5.94\text{ mol}}{89.24\text{ mol}} = 0.0666$$

(b)

$$\text{mol NaCl} = \frac{0.86\text{ g}}{58.44\text{ g/mol}} = 0.015\text{ mol}$$

$$\text{mol H}_2\text{O} = \frac{100\text{ g}}{18.0153\text{ g mol}^{-1}} = 5.551\text{ mol}$$

Total number of moles = 0.015 mol + 5.551 mol = 5.566 mol

$$X_{\text{NaCl}} = \frac{0.015\text{ mol}}{5.566\text{ mol}} = 0.0027$$

$$X_{\text{H}_2\text{O}} = \frac{5.551\text{ mol}}{5.566\text{ mol}} = 0.997$$

If the mole fraction of water is calculated by subtraction of the mole fraction of NaCl from 1, the answer is 0.9973. This procedure points out the importance in maintaining the correct number of significant figures and shows that slightly different answers may be obtained by different methods.

(c)

$$\text{mol C}_{18}\text{H}_{21}\text{NO}_3 = \frac{46.85\text{ g}}{299.364\text{ g mol}^{-1}} = 0.156\text{ mol}$$

$$\text{mol C}_2\text{H}_5\text{OH} = \frac{125\text{ g}}{46.068\text{ g mol}^{-1}} = 2.724\text{ mol}$$

Total number of moles = 0.156 mol + 2.724 mol = 2.880 mol

$$X_{\text{C}_{15}\text{H}_{21}\text{NO}_3} = \frac{0.156\text{ mol}}{2.880\text{ mol}} = 0.0542$$

$$X_{\text{C}_2\text{H}_5\text{OH}} = \frac{2.724\text{ mol}}{2.880\text{ mol}} = 0.9458$$

(d)

$$\text{mol I}_2 = \frac{25\text{ g}}{253.809\text{ g mol}^{-1}} = 0.0985\text{ mol}$$

$$\text{mol C}_2\text{H}_5\text{OH} = \frac{125\text{ g}}{46.068\text{ g mol}^{-1}} = 2.713\text{ mol}$$

Total number of moles = 0.0985 mol + 2.713 mol = 2.812 mol

$$X_{\text{I}_2} = \frac{0.0985\text{ mol}}{2.812\text{ mol}} = 0.035$$

$$X_{\text{C}_2\text{H}_5\text{OH}} = \frac{2.713\text{ mol}}{2.812\text{ mol}} = 0.965$$

33. Calculate the mole fraction of each solute and solvent:

(a) 0.710 kg of sodium carbonate (washing soda), Na_2CO_3, in 10.0 kg of water—a saturated solution at 0 °C

(b) 125 g of NH_4NO_3 in 275 g of water—a mixture used to make an instant ice pack

(c) 25 g of Cl_2 in 125 g of dichloromethane, CH_2Cl_2

(d) 0.372 g of tetrahydropyridine, C_5H_9N, in 125 g of chloroform, $CHCl_3$

Solution

(a)

$$\text{mol Na}_2\text{CO}_3 = 710\ \cancel{\text{g Na}_2\text{CO}_3} \times \frac{1\ \text{mol}}{105.9886\ \cancel{\text{g Na}_2\text{CO}_3}} = 6.70\ \text{mol}$$

$$\text{mol H}_2\text{O} = \frac{10{,}000\ \text{g}}{18.0153\ \text{g/mol}} = 555.08\ \text{mol}$$

Total number of moles = 555.08 mol + 6.70 mol = 561.78 mol

$$X_{\text{Na}_2\text{CO}_3} = \frac{6.70\ \text{mol}}{561.78\ \text{mol}} = 0.0119$$

$$X_{\text{H}_2\text{O}} = \frac{555.08\ \text{mol}}{561.78\ \text{mol}} = 0.988$$

(b)

$$\text{mol NH}_4\text{NO}_3 = 125\ \cancel{\text{g NH}_4\text{NO}_3} \times \frac{1\ \text{mol}}{80.0434\ \cancel{\text{g NH}_4\text{NO}_3}} = 1.56\ \text{mol}$$

$$\text{mol H}_2\text{O} = \frac{275\ \text{g}}{18.0153\ \text{g/mol}} = 15.26\ \text{mol}$$

Total number of moles = 15.26 mol + 1.56 mol = 16.82 mol

$$X_{\text{NH}_4\text{NO}_3} = \frac{1.56\ \text{mol}}{16.82\ \text{mol}} = 0.9927$$

$$X_{\text{H}_2\text{O}} = \frac{15.26\ \text{mol}}{16.82\ \text{mol}} = 0.907$$

(c)

$$\text{mol Cl}_2 = 25\ \cancel{\text{g Cl}_2} \times \frac{1\ \text{mol}}{70.9054\ \cancel{\text{g Cl}_2}} = 0.35\ \text{mol}$$

$$\text{mol CH}_2\text{Cl}_2 = \frac{125\ \text{g}}{84.93\ \text{g/mol}} = 1.47\ \text{mol}$$

Total number of moles = 1.47 mol + 0.35 mol = 1.82 mol

$$X_{\text{Cl}_2} = \frac{0.35\ \text{mol}}{1.82\ \text{mol}} = 0.192$$

$$X_{\text{CH}_2\text{Cl}_2} = \frac{1.47\ \text{mol}}{1.82\ \text{mol}} = 0.808$$

(d)

$$\text{mol C}_5\text{H}_9\text{N} = 0.372\ \cancel{\text{g C}_5\text{H}_9\text{N}} \times \frac{1\ \text{mol}}{83.1332\ \cancel{\text{g C}_5\text{H}_9\text{N}}} = 4.47 \times 10^{-3}\ \text{mol}$$

$$\text{mol CHCl}_3 = \frac{125\ \text{g}}{119.38\ \text{g/mol}} = 1.047\ \text{mol}$$

Total number of moles = 1.047 mol + 0.00447 mol = 1.05 mol

$$X_{\text{C}_5\text{H}_9\text{N}} = \frac{0.00447\ \text{mol}}{1.05\ \text{mol}} = 0.00426$$

$$X_{\text{CHCl}_3} = \frac{1.047\ \text{mol}}{1.05\ \text{mol}} = 0.997$$

34. Calculate the mole fractions of methanol, CH_3OH; ethanol, C_2H_5OH; and water in a solution that is 40% methanol, 40% ethanol, and 20% water by mass. (Assume the data are good to two significant figures.)

Solution

Assume that the total solution mass is 100 g. Calculate the mass and the number of moles of each component. Then calculate the respective mole fractions.

$$\text{mol } CH_3OH = 0.40 \times 100 \text{ g} \times \frac{1 \text{ mol}}{32.042 \text{ g}} = 1.24836 \text{ mol}$$

$$\text{mol } C_2H_5OH = 0.40 \times 100 \text{ g} \times \frac{1 \text{ mol}}{46.069 \text{ g}} = 0.86827 \text{ mol}$$

$$\text{mol } H_2O = 0.20 \times 100 \text{ g} \times \frac{1 \text{ mol}}{18.0152 \text{ g}} = 1.1101736 \text{ mol}$$

Total number of moles = 1.24836 mol + 0.84827 mol + 1.1101736 mol = 3.22680 mol

$$X_{CH_3OH} = \frac{1.24836}{3.22680} = 0.39$$

$$X_{C_2H_5OH} = \frac{0.86827}{3.22680} = 0.27$$

$$X_{H_2O} = \frac{1.1101736}{3.22680} = 0.34$$

35. What is the difference between a 1 *M* solution and a 1 *m* solution?

Solution

In a 1 *M* solution, the mole is contained in exactly 1 L of solution. In a 1 *m* solution, the mole is contained in exactly 1 kg of solvent.

36. What is the molality of phosphoric acid, H_3PO_4, in a solution of 14.5 g of H_3PO_4 in 125 g of water?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

Solution

(a) Determine the molar mass of H_3PO_4; determine the number of moles of acid in the solution; from the number of moles and the mass of solvent, determine the molality. (b) 1.18 *m*

37. What is the molality of nitric acid in a concentrated solution of nitric acid (68.0% HNO_3 by mass)?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

Solution

(a) Determine the molar mass of HNO_3. Determine the number of moles of acid in the solution. From the number of moles and the mass of solvent, determine the molality.

(b) Molar mass HNO_3 = 63.01288 g mol^{-1}

If we assume 100 g of solution, then 68.0 g is HNO_3 and 32.0 g is water.

$$\text{mol } HNO_3 = 68.0 \cancel{\text{ g } HNO_3} \times \frac{1 \text{ mol}}{63.02188 \cancel{\text{ g } HNO_3}} = 1.08 \text{ mol}$$

$$m \text{ } HNO_3 = \frac{1.08 \text{ mol}}{0.0320 \text{ g}} = 33.7\ m$$

38. Calculate the molality of each of the following solutions:
(a) 583 g of H_2SO_4 in 1.50 kg of water—the acid solution used in an automobile battery
(b) 0.86 g of NaCl in 1.00×10^2 g of water—a solution of sodium chloride for intravenous injection
(c) 46.85 g of codeine, $C_{18}H_{21}NO_3$, in 125.5 g of ethanol, C_2H_5OH
(d) 25 g of I_2 in 125 g of ethanol, C_2H_5OH

Solution
(a) 3.96 *m*; (b) 0.15 *m*; (c) 1.247 *m*; (d) 0.79 *m*

39. Calculate the molality of each of the following solutions:
(a) 0.710 kg of sodium carbonate (washing soda), Na_2CO_3, in 10.0 kg of water—a saturated solution at 0 °C
(b) 125 g of NH_4NO_3 in 275 g of water—a mixture used to make an instant ice pack
(c) 25 g of Cl_2 in 125 g of dichloromethane, CH_2Cl_2
(d) 0.372 g of tetrahydropyridine, C_5H_9N, in 125 g of chloroform, $CHCl_3$

Solution
(a)

$$\text{mol Na}_2\text{CO}_3 = 710\text{ g Na}_2\text{CO}_3 \times \frac{1\text{ mol}}{105.9886\text{ g Na}_2\text{CO}_3}$$

$$\text{molality of Na}_2\text{CO}_3 = \frac{6.70\text{ mol}}{10.0\text{ kg}} = 6.70 \times 10^{-1}\ m$$

(b)

$$\text{mol NH}_4\text{NO}_3 = 125\ \cancel{\text{g NH}_4\text{NO}_3} \times \frac{1\text{ mol}}{80.0434\ \cancel{\text{g NH}_4\text{NO}_3}} = 1.56\text{ mol}$$

$$\text{molality of NH}_4\text{NO}_3 = \frac{1.56\text{ mol}}{0.275\text{ kg}} = 5.67\ m$$

(c)

$$\text{mol Cl}_2 = 25\ \cancel{\text{g Cl}_2} \times \frac{1\text{ mol}}{70.9054\ \cancel{\text{g Cl}_2}} = 0.35\text{ mol}$$

$$m\ \text{Cl}_2 = \frac{0.35\text{ mol}}{0.125\text{ kg}} = 2.8\ m$$

(d)

$$\text{mol C}_5\text{H}_9\text{N} = 0.372\ \cancel{\text{g C}_5\text{H}_9\text{N}} \times \frac{1\text{ mol}}{83.1332\ \cancel{\text{g C}_5\text{H}_9\text{N}}} = 4.47 \times 10^{-3}\text{ mol}$$

$$\text{molality of C}_5\text{H}_9\text{N} = \frac{4.47 \times 10^{-3}\text{ mol}}{0.125\text{ kg}} = 0.0358\ m$$

40. The concentration of glucose, $C_6H_{12}O_6$, in normal spinal fluid is $\frac{75\text{ mg}}{100\text{ g}}$. What is the molality of the solution?

Solution
4.2×10^{-3} *m*

41. A 13.0% solution of K_2CO_3 by mass has a density of 1.09 g/cm^3. Calculate the molality of the solution.

Solution

Find the mass of K_2CO_3 and the mass of water in solution. Assume 100.0 mL of solution and that the density of water is 1.00 g cm^{-3}. Then find the moles of K_2CO_3 and the molality.

$$\text{Mass (solution)} = 100.0\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times 1.09\ \text{g}\ \cancel{\text{cm}^3} = 109.0\ \text{g}$$

$$\text{Mass }(K_2CO_3) = \frac{13.0\%}{100\%} \times 109\ \text{g} = 14.2\ \text{g}$$

Mass (H_2O) = 109.0 g − 14.2 g = 94.8 g

$$m\ (K_2CO_3) = \frac{0.1027\ \text{mol}}{0.0948\ \text{kg}} = 1.08\ m$$

42. Why does 1 mol of sodium chloride depress the freezing point of 1 kg of water almost twice as much as 1 mol of glycerin?

Solution

The presence of two ions for each NaCl molecule in a solution of NaCl compared with only one molecule for each glycerin in a solution of glycerin doubles the freezing point depression. This relationship is accounted for by the van't Hoff factor in the expression $\Delta T_f = K_f m$.

43. Assuming ideal solution behavior, what is the boiling point of a solution of 115.0 g of nonvolatile sucrose, $C_{12}H_{22}O_{11}$, in 350.0 g of water?

(a) Outline the steps necessary to answer the question

(b) Answer the question

Solution

(a) Determine the molar mass of sucrose; determine the number of moles of sucrose in the solution; convert the mass of solvent to units of kilograms; from the number of moles and the mass of solvent, determine the molality; determine the difference between the boiling point of water and the boiling point of the solution; determine the new boiling point.

(b) $$\text{mol sucrose} = \frac{115.0\ \text{g}}{342.300\ \text{g mol}^{-1}} = 0.3360\ \text{mol}$$

$$\text{molality} = \frac{0.3360\ \text{mol}\ C_{12}H_{22}O_{11}}{0.3500\ \text{kg}\ H_2O} = 0.9599\ m$$

$\Delta T_b = K_b m$ = (0.512 °C m^{-1})(0.9599 m) = 0.491 °C

The boiling point of pure water at 100.0 °C increases 0.491 °C to 100.491 °C, or 100.5 °C.

44. Assuming ideal solution behavior, what is the boiling point of a solution of 9.04 g of I_2 in 75.5 g of benzene, assuming the I_2 is nonvolatile?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

Solution

(a) Determine the molar mass of I_2. Determine the number of moles of I_2 in the solution. From the number of moles and the mass of solvent, determine the molality. Then determine the difference between the boiling point of benzene and the boiling point of the solution. Finally, determine the new boiling point.

(b)

$\Delta T_b = K_b m$ = (2.53 °C m^{-1})(0.472 m) = 1.19 °C

The boiling point of pure benzene is 80.1 °C. Since the temperature increase is 1.19 °C, the new temperature is (80.1 + 1.19) = 81.3 °C.

45. Assuming ideal solution behavior, what is the freezing temperature of a solution of 115.0 g of sucrose, $C_{12}H_{22}O_{11}$, in 350.0 g of water?
(a) Outline the steps necessary to answer the question.
(b) Answer the question.

Solution
(a) Determine the molar mass of sucrose; determine the number of moles of sucrose in the solution; convert the mass of solvent to units of kilograms; from the number of moles and the mass of solvent, determine the molality; determine the difference between the freezing temperature of water and the freezing temperature of the solution; determine the new freezing temperature.

(b) $$\text{mol sucrose} = \frac{115.0\text{ g}}{342.300\text{ g mol}^{-1}} = 0.336\text{ mol}$$

$$m\text{ sucrose} = \frac{0.336\text{ mol}}{0.350\text{ kg}} = 0.960\ m$$

$\Delta T_b = K_b m = (1.86\ °\text{C}\ m^{-1})(0.960\ m) = 1.78\ °\text{C}$
The freezing temperature is 0.0 °C − 1.78 °C = –1.8 °C.

46. Assuming ideal solution behavior, what is the freezing point of a solution of 9.04 g of I_2 in 75.5 g of benzene?
(a) Outline the steps necessary to answer the following question.
(b) Answer the question.

Solution
(a) Determine the molar mass of I_2. Determine the number of moles of I_2 in the solution. From the number of moles and the mass of solvent, determine the molality. Then determine the difference between the freezing temperature of pure benzene and the freezing temperature of the solution. Finally, determine the new freezing temperature.

(b) $$\text{mol } I_2 = \frac{9.04\text{ g}}{\left(253.80894\text{ g mol}^{-1}\right)} = 0.0356\text{ mol}$$

$$m\ I_2 = \frac{0.0356\text{ mol}}{0.0755\text{ kg}} = 0.472\ m$$

$\Delta T_b = K_b m = (5.12\ °\text{C}\ m^{-1})(0.472\ m) = 2.42\ °\text{C}$
Benzene freezes at 5.5 °C. The freezing temperature is 5.5 °C – 2.42 °C = 3.1 °C.

47. Assuming ideal solution behavior, what is the osmotic pressure of an aqueous solution of 1.64 g of $Ca(NO_3)_2$ in water at 25 °C? The volume of the solution is 275 mL.
(a) Outline the steps necessary to answer the question.
(b) Answer the question.

Solution
(a) Determine the molar mass of $Ca(NO_3)_2$; determine the number of moles of $Ca(NO_3)_2$ in the solution; determine the number of moles of ions in the solution; determine the molarity of ions, then the osmotic pressure.

(b) $$M\ \text{Ca}\left(\text{NO}_3\right)_2 = \frac{1.64\text{ g Ca}\left(\text{NO}_3\right)_2 \times 1\text{ mol}/164.088\text{ g Ca(NO}_3)_2}{0.275\text{ L}} = 0.363\ M$$

The molarity of the ions is three times the molarity of $Ca(NO_3)_2$. Therefore, multiply the molarity of $Ca(NO_3)_2$ by 3: $\Pi = MRT = 3 \times 0.0363\text{ mol L}^{-1} \times 0.08206\text{ L atm mol}^{-1}\text{ K}^{-1} \times 298.15\text{ K} = 2.67\text{ atm}$.

48. Assuming ideal solution behavior, what is osmotic pressure of a solution of bovine insulin (molar mass, 5700 g mol^{-1}) at 18 °C if 100.0 mL of the solution contains 0.103 g of the insulin?
(a) Outline the steps necessary to answer the question.
(b) Answer the question.

Solution
(a) From the molar mass of bovine insulin, determine the number of moles of insulin in the solution. Determine the molarity of insulin, and then determine the osmotic pressure.

(b) $$M \text{ insulin} = \frac{\frac{0.103\ \cancel{\text{g}}}{5700\ \cancel{\text{g}}\ \text{mol}^{-1}}}{0.100\ \text{L}} = 1.807 \times 10^{-4}\ M$$

$\Pi = MRT = 1.804 \times 10^{-4}\ \text{mol L}^{-1} \times 0.0826\ \text{L atm mol}^{-1}\ \text{K}^{-1} \times (273.15 + 18)\ \text{K} = 4.32 \times 10^{-4}\ \text{atm}$

49. Assuming ideal solution behavior, what is the molar mass of a solution of 5.00 g of a compound in 25.00 g of carbon tetrachloride (bp 76.8 °C; K_b = 5.02 °C/*m*) that boils at 81.5 °C at 1 atm?
(a) Outline the steps necessary to answer the question.
(b) Solve the problem.

Solution
(a) Determine the molal concentration from the change in boiling point and K_b; determine the moles of solute in the solution from the molal concentration and mass of solvent; determine the molar mass from the number of moles and the mass of solute. (b) ΔT_b = 81.5 − 76.8 = 4.7 °C,

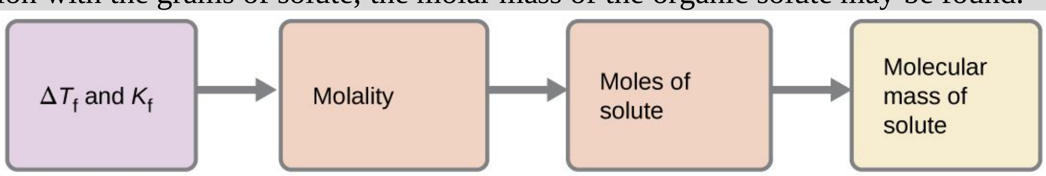
$\Delta T_b = K_b m$, so $m = \frac{\Delta T_b}{K_b} = \frac{4.7\ °\text{C}}{5.02\ °\text{C}/m} = 0.94\ m$. Moles of solute = molality × kg of solvent = 0.94 *m* × 0.02500 kg = 0.024 mol;

$$\text{Molar mass} = \frac{\text{mass}}{\text{moles}} = \frac{5.00\ \text{g}}{0.024\ \text{mol}} = 2.1 \times 10^2\ \text{g mol}^{-1}$$

Molecular mass = 2.1×10^2 amu

50. A sample of an organic compound (a nonelectrolyte) weighing 1.35 g lowered the freezing point of 10.0 g of benzene by 3.66 °C. Assuming ideal solution behavior, calculate the molar mass of the compound.

Solution
The freezing point lowering (ΔT_f) is known to be 3.66 °C. Also, K_f for benzene can be obtained from Table 11.2 (K_f = 5.12 °C/*m*). Therefore, the equation $\Delta T_f = K_f m$ can be used to find the molality of the solution. The definition of molality is moles of solute per kilogram of solvent. By using the mass of the solvent, benzene, the number of moles of solute may be found. Using this information with the grams of solute, the molar mass of the organic solute may be found.



$\Delta T_f = K_f m$
$3.66\ °\text{C} = (5.12\ °\text{C}/m) \times m$

$$m = \frac{\text{mol of solute}}{\text{kg solvent}} = 0.715$$

$$\text{molality} = \frac{0.715 \text{ mol solute}}{\text{kg solvent}} = 0.715\ m$$

$$0.715\ m = \frac{x}{0.010 \text{ kg benzene}}$$

$x$mol = 0.715 × 0.010 = 7.15 × 10^{-3}mol solute

$$\text{mol} = \frac{\text{g}}{\text{molar mass}}$$

$$7.15 \times 10^{-3} \text{ mol solute} = \frac{1.350 \text{ g}}{\text{molar mass}}$$

Molar mass = 189 g mol^{-1}

51. A 1.0 *m* solution of HCl in benzene has a freezing point of 0.4 °C. Is HCl an electrolyte in benzene? Explain.

Solution

No. Pure benzene freezes at 5.5 °C, and so the observed freezing point of this solution is depressed by ΔT_f = 5.5 – 0.4 = 5.1 °C. The value computed, assuming no ionization of HCl, is ΔT_f = (1.0 m)(5.14 °C/*m*) = 5.1 °C. Agreement of these values supports the assumption that HCl is not ionized.

52. A solution contains 5.00 g of urea, $CO(NH_2)_2$, a nonvolatile compound, dissolved in 0.100 kg of water. If the vapor pressure of pure water at 25 °C is 23.7 torr, what is the vapor pressure of the solution (assuming ideal solution behavior)?

Solution

The vapor pressure of the pure solvent is known; therefore, Raoult's law can be used along with the mole fraction to calculate the vapor pressure of the solution. First, calculate the molar amounts of urea and water, then compute the mole fraction of water, then use Raoult's law to compute the solution's vapor pressure:

$$P_{\text{solution}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}}$$

$$X_{H_2O} = \frac{5.55}{0.833 + 5.55} = 0.9852$$

$$P_{\text{solution}} = (0.9852)(23.7 \text{ torr}) = 23.3 \text{ torr}$$

53. A 12.0-g sample of a nonelectrolyte is dissolved in 80.0 g of water. The solution freezes at –1.94 °C. Assuming ideal solution behavior, calculate the molar mass of the substance.

Solution

ΔT_f = 1.94 °C

$$m = \frac{\Delta T_f}{K_f} = \frac{1.94 \text{ °C}}{1.86 \text{ °C}/m} = 1.04\ m$$

Moles of solute = 1.04 *m* × 0.0800 kg = 0.0834 mol

$$\text{Molar mass} = \frac{12.0 \text{ g}}{0.0834 \text{ mol}} = 144 \text{ g mol}^{-1}$$

Molecular mass = 144 amu

54. Arrange the following solutions in order by their decreasing freezing points: 0.1 *m* Na_3PO_4, 0.1 *m* C_2H_5OH, 0.01 *m* CO_2, 0.15 *m* NaCl, and 0.2 *m* $CaCl_2$.

Solution

The number of particles (ions or molecules) per mole is Na_3PO_4, 4; C_2H_5OH, 1; CO_2, 1; NaCl, 2; and $CaCl_2$, 3. For a 1.00-L solution, the number of moles of particles is:

Na_3PO_4: 0.1 mol × 4 = 0.4 mol
C_2H_5OH: 0.1 mol × 1 = 0.1 mol
CO_2: 0.01 mol × 1 = 0.01 mol
NaCl: 0.15 mol × 2 = 0.30 mol
$CaCl_2$: 0.2 mol × 3 = 0.6 mol

From the highest freezing point to lowest freezing point:
0.01 mol CO_2 > 0.1 mol C_2H_5OH > 0.15 mol NaCl > 0.1 mol Na_3PO_4 > 0.2 mol $CaCl_2$.

55. Calculate the boiling point elevation of 0.100 kg of water containing 0.010 mol of NaCl, 0.020 mol of Na_2SO_4, and 0.030 mol of $MgCl_2$, assuming complete dissociation of these electrolytes and ideal solution behavior.

Solution

0.010 mol NaCl contains 0.010 mol Na^+ + 0.010 mol $Cl^–$
0.020 mol Na_2SO_4 contains 0.040 mol Na^+ + 0.020 mol $SO_4^{2–}$
0.030 mol $MgCl_2$ contains 0.030 mol Mg^{2+} + 0.060 mol $Cl^–$
Total numbers of moles = 0.020 mol + 0.060 mol + 0.090 mol = 0.170 mol

$$\Delta T_\text{b} = K_\text{b}m = 0.512\ \text{°C}/m \times \frac{0.170\ \text{mol}}{0.100\ \text{kg}} = 0.870\ \text{°C}$$

56. How could you prepare a 3.08 *m* aqueous solution of glycerin, $C_3H_8O_3$? Assuming ideal solution behavior, what is the freezing point of this solution?

Solution

First, find the molar mass of glycerin. The molar mass of glycerin is 92.095 g $mol^{–1}$. A 3.08 *m* aqueous solution requires 3.08 mol × 92.095 g $mol^{–1}$ = 284 g

To prepare the solution, dissolve 284 g of glycerin in 1.00 kg of water. Glycerin is nonelectrolyte.

ΔT_f = 1.86 °C/*m* × 3.08 m = 5.73

The freezing point is –5.73 °C.

57. A sample of sulfur weighing 0.210 g was dissolved in 17.8 g of carbon disulfide, CS_2 (K_b = 2.43 °C/*m*). If the boiling point elevation was 0.107 °C, what is the formula of a sulfur molecule in carbon disulfide (assuming ideal solution behavior)?

Solution

The molality is

$$m = \frac{0.107\ \text{°C}}{2.34\ \text{°C}/m} = 0.00457\ m$$

mol S = 4.57 *m* × 0.0178 kg = 8.13 × $10^{–4}$ mol

$$\text{Molecular mass} = \frac{0.210\ \text{g}}{8.13 \times 10^{-4}\ \text{mol}} = 285\ \text{g mol}^{-1}$$

The atomic mass of sulfur is 32.066.

$$\frac{258}{32.066} = 8.05$$

The formula for the sulfur molecule is S_8.

58. In a significant experiment performed many years ago, 5.6977 g of cadmium iodide in 44.69 g of water raised the boiling point 0.181 °C. What does this suggest about the nature of a solution of CdI_2?

Solution

Determine the number of moles of CdI_2 and then the molarity.

From the elevation of the boiling point, determine the molarity expected to produce that increase. Compare that value with the calculated value. If a difference exists, the ratio of the actual value to the calculated value gives the number of ions.

$$\text{mol CdI}_2 = \frac{5.6977 \text{ g}}{366.220 \text{ g mol}^{-1}} = 0.015558 \text{ mol}$$

$$m \text{ CdI}_2 = \frac{0.015558 \text{ mol}}{0.04469 \text{ kg}} = 0.3481\ m$$

The boiling point elevation expected if no ionization occurs is:

$\Delta T_b = K_b m = 0.512\ °\text{C}\ m^{-1} \times 0.3481\ m = 0.178\ °\text{C}$.

The actual boiling point increase is 0.181°C. Therefore, since the ratio is almost 1, only a very slight amount of CdI_2 actually dissociated in solution.

59. Lysozyme is an enzyme that cleaves cell walls. A 0.100-L sample of a solution of lysozyme that contains 0.0750 g of the enzyme exhibits an osmotic pressure of 1.32×10^{-3} atm at 25 °C. Assuming ideal solution behavior, what is the molar mass of lysozyme?

Solution

The molarity of the solution is:

$$M = \frac{\Pi}{RT} = \frac{1.32 \times 10^{-3} \text{ atm}}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 5.40 \times 10^{-5} \text{ mol L}^{-1}$$

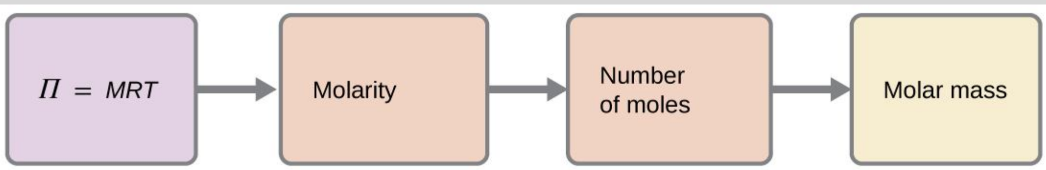
Number of moles = 5.40×10^{-5} mol $\text{L}^{-1} \times 0.100$ L = 5.40×10^{-6} mol

$$\text{Molar mass} = \frac{0.0750 \text{ g}}{5.40 \times 10^{-6} \text{ mol}} = 1.39 \times 10^{4} \text{ g mol}^{-1}$$

Molecular mass = 1.39×10^{4} amu.

60. The osmotic pressure of a solution containing 7.0 g of insulin per liter is 23 torr at 25 °C. Assuming ideal solution behavior, what is the molar mass of insulin?

Solution



First, since the ideal gas constant is in units of (L atm $\text{mol}^{-1} \cdot \text{K}^{-1}$), convert the osmotic pressure to atmospheres. For the same reasons, convert the temperature from °C to kelvin.

1 atm = 760 torr

$$23 \cancel{\text{torr}} \times \frac{1 \text{ atm}}{760 \cancel{\text{torr}}} = 0.030 \text{ atm}$$

$$M = \frac{\Pi}{RT} = \frac{0.030 \cancel{\text{atm}}}{(0.08206 \text{ L} \cancel{\text{atm}} \text{ mol}^{-1} \cancel{\text{K}^{-1}})(298.15 \text{ K})} = 1.23 \times 10^{-3}\ M$$

Number of moles = MV (in liters) = $(1.23 \times 10^{-3} M)(1 \text{ L}) = 1.23 \times 10^{-3} M$

$$\text{Number of moles} = \frac{\text{g}}{\text{molar mass}}$$

$$1.23 \times 10^{-3}\ \text{mol} = \frac{7.0\ \text{g insulin}}{\text{molar mass}}$$

Molar mass = 5.7×10^3 g mol^{-1}

61. The osmotic pressure of human blood is 7.6 atm at 37 °C. What mass of glucose, $C_6H_{12}O_6$, is required to make 1.00 L of aqueous solution for intravenous feeding if the solution must have the same osmotic pressure as blood at body temperature, 37 °C (assuming ideal solution behavior)?

Solution

The molarity of the solution is

$$M = \frac{\Pi}{RT} = \frac{7.6\ \text{atm}}{(0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1})(310\ \text{K})} = 0.30\ \text{mol/L}$$

Number of moles = 0.30 mol/L × 1.00 L = 0.30 mol

Mass (glucose) = 180.157 g mol^{-1} × 0.30 mol = 54 g

62. Assuming ideal solution behavior, what is the freezing point of a solution of dibromobenzene, $C_6H_4Br_2$, in 0.250 kg of benzene, if the solution boils at 83.5 °C?

Solution

Find the molality of the solution from the boiling point elevation. Using that value, determine the freezing point depression and then the freezing point. Find the values for the constants needed in Table 11.2.

ΔT = 83.5 °C – 80.1 °C = 3.4 °C = $K_b m$ = 2.53 °C m^{-1} × m

$$m\ \text{dibromobenzene} = \frac{3.4\ °\text{C}}{2.53\ °\text{C}\ m^{-1}} = 1.34\ m$$

$\Delta T = K_f m$ = 5.12 °C m^{-1} × 1.34 m = 6.86 °C

The freezing point of pure benzene is 5.5 °C. Subtraction gives 5.5 °C − 6.86 °C = –1.4 °C.

63. Assuming ideal solution behavior, what is the boiling point of a solution of NaCl in water if the solution freezes at –0.93 °C?

Solution

Find the molality of the solution from the freezing point depression. Using that value, determine the boiling point elevation and then the boiling point.

$$\Delta T_f = |0.0\ °\text{C} - 0.93\ °\text{C}| = 0.93\ °\text{C} = K_f m = 1.86\ °\text{C}\ m^{-1} \times m$$

$$m\ \text{NaCl} = \frac{0.93\ °\text{C}}{1.86\ °\text{C}\ m^{-1}} = 0.50\ m$$

$\Delta T_b = K_b m$ = 0.512 °C m^{-1} × 0.50 m = 0.256 °C

The boiling point of pure water is 100.00 °C. Addition gives 100.00 °C + 0.26 °C = 100.26 °C.

64. The sugar fructose contains 40.0% C, 6.7% H, and 53.3% O by mass. A solution of 11.7 g of fructose in 325 g of ethanol has a boiling point of 78.59 °C. The boiling point of ethanol is 78.35 °C, and K_b for ethanol is 1.20 °C/m. Assuming ideal solution behavior, what is the molecular formula of fructose?

Solution

Δbp = $K_b m$ and molecular mass = $(K_b)\dfrac{(\text{mass fructose})}{(\text{mass ethanol})}$

$$\text{molecular mass} = \frac{1.20\ °\text{C}/m \times 11.7\ \text{g}}{(0.24\ °\text{C})(0.325\ \text{kg})} = 180\ \text{g mol}^{-1}$$

Empirical formula:

$$\text{C:}\ \frac{40.0\ \text{g}}{12.011\ \text{g mol}^{-1}} = 3.330\ \text{mol} \qquad \frac{3.330}{3.33} = 1$$

$$\text{H:}\ \frac{6.7\ \text{g}}{1.00794\ \text{g mol}^{-1}} = 6.647\ \text{mol} \qquad \frac{6.647}{3.33} = 2$$

$$\text{O:}\ \frac{53.3\ \text{g}}{15.9994\ \text{g mol}^{-1}} = 3.331\ \text{mol} \qquad \frac{3.331}{3.33} = 1$$

The molar mass of CH_2O is 30 g/mol.

$$\text{C:}\ \frac{40.0\ \text{g}}{12.011\ \text{g mol}^{-1}} = 3.330\ \text{mol} \qquad \frac{3.330}{3.33} = 1$$

$$\text{H:}\ \frac{6.7\ \text{g}}{1.00794\ \text{g mol} - 1} = 6.647\ \text{mol} \qquad \frac{6.647}{3.33} = 2$$

$$\text{O:}\ \frac{53.3\ \text{g}}{15.9994\ \text{g mol}^{-1}} = 3.331\ \text{mol} \qquad \frac{3.331}{3.33} = 1$$

Given the molecular mass (180), the molecular formula is $C_6H_{12}O_6$.

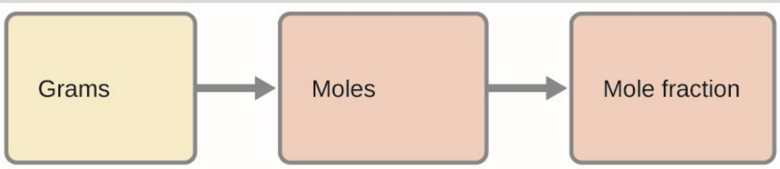
65. The vapor pressure of methanol, CH_3OH, is 94 torr at 20 °C. The vapor pressure of ethanol, C_2H_5OH, is 44 torr at the same temperature.

(a) Calculate the mole fraction of methanol and of ethanol in a solution of 50.0 g of methanol and 50.0 g of ethanol.

(b) Ethanol and methanol form a solution that behaves like an ideal solution. Calculate the vapor pressure of methanol and of ethanol above the solution at 20 °C.

(c) Calculate the mole fraction of methanol and of ethanol in the vapor above the solution.

Solution



(a)

$$X_A = \frac{X_A}{X_A + X_B}$$

$CH_3OH = 32.04246\ \text{g mol}^{-1}$

$C_2H_5OH = 46.063\ \text{g mol}^{-1}$

$$\text{mol CH}_3\text{OH} = \frac{50.0\ \cancel{\text{g}}}{32.04216\ \cancel{\text{g}}\ \text{mol}^{-1}} = 1.5604\ \text{mol}$$

$$\text{mol C}_2\text{H}_5\text{OH} = \frac{50.0\ \cancel{\text{g}}}{46.069\ \cancel{\text{g}}\ \text{mol}^{-1}} = 1.0853\ \text{mol}$$

$$X_{\text{CH}_3\text{OH}} = \frac{1.5604}{1.5604 + 1.0853} = 0.590$$

$$X_{\text{C}_2\text{H}_5\text{OH}} = \frac{1.0853}{1.5604 + 1.0853} = 0.410$$

(b) Vapor pressures are:
CH_3OH: 0.590 × 94 torr = 55 torr
C_2H_5OH: 0.410 × 44 torr = 18 torr
(c) The number of moles of each substance is proportional to the pressure, so the mole fraction of each component in the vapor can be calculated as follows:

$$\text{CH}_3\text{OH}: \ \frac{55}{(55 + 18)} = 0.75$$

$$\text{C}_2\text{H}_5\text{OH}: \ \frac{18}{(55 + 18)} = 0.25$$

66. The triple point of air-free water is defined as 273.16 K. Why is it important that the water be free of air?

Solution
As a solute, dissolved air will change the vapor pressure and, hence, the freezing point of the solution.

67. Meat can be classified as fresh (not frozen) even though it is stored at –1 °C. Why wouldn't meat freeze at this temperature?

Solution
The ions and compounds present in the water in the beef lower the freezing point of the beef below –1 °C.

68. An organic compound has a composition of 93.46% C and 6.54% H by mass. A solution of 0.090 g of this compound in 1.10 g of camphor melts at 158.4 °C. The melting point of pure camphor is 178.4 °C. K_f for camphor is 37.7 °C/*m*. Assuming ideal solution behavior, what is the molecular formula of the solute? Show your calculations.

Solution
Calculate the molecular mass of the substance. Then, using methods for calculating the empirical formula of the compound, compare the empirical formula mass to the calculated molecular mass to determine the formula of the compound.

$$\Delta\text{fp} = K_f m = K_f \left[\frac{\text{mass solute/molar mass}}{\text{mass camphor (kg)}}\right]$$

$$\text{Molar mass} = \frac{K_f\,(\text{mass solute})}{(\Delta\text{fp})(\text{kg camphor})} = \frac{(37.7\ {}^\circ\text{C}/m)(0.045\ \text{g})}{(20.0\ {}^\circ\text{C})(0.000550\ \text{kg})} = 154.2\ \text{g mol}^{-1}$$

Empirical formula:

$$\text{C:}\ \frac{93.46\ \text{g}}{12.011\ \text{g mol}^{-1}} = 7.781\ \text{mol} \qquad \frac{7.781}{6.4885} = 1.199\ \text{or}\ 1.2$$

H: $\frac{6.5 \text{ g}}{1.00794 \text{ g mol}^{-1}}$ = 6.4885 mol $\frac{6.4887}{6.4885} = 1$

Ratio: $C_{1.2}H_1$ or $C_{12}H_{10}$

The formula mass of $C_{12}H_{10}$ agrees with the calculated formula mass.

69. A sample of $HgCl_2$ weighing 9.41 g is dissolved in 32.75 g of ethanol, C_2H_5OH (K_b = 1.20 °C/*m*). The boiling point elevation of the solution is 1.27 °C. Is $HgCl_2$ an electrolyte in ethanol? Show your calculations.

Solution

$$\Delta\text{bp} = K_b m = (1.20\ °\text{C}/m)\left(\frac{9.41 \text{ g} \times \frac{1 \text{ mol HgCl}_2}{271.496 \text{ g}}}{0.03275 \text{ kg}}\right) = 1.27\ °\text{C}$$

The observed change equals the theoretical change; therefore, no dissociation occurs.

70. A salt is known to be an alkali metal fluoride. A quick approximate determination of freezing point indicates that 4 g of the salt dissolved in 100 g of water produces a solution that freezes at about –1.4 °C. Assuming ideal solution behavior, what is the formula of the salt? Show your calculations.

Solution

Assuming complete dissociation of the alkali metal fluoride, $i = 2$.

$$m = \frac{\Delta T_f}{2 \times K_f} = \frac{1.4\ °\text{C}}{2 \times 1.86\ °\text{C}/m} = 0.38\ m$$

Moles of fluoride = 0.100 kg × 0.38 *m* = 0.038 mol

$$\text{Molar mass} = \frac{4 \text{ g}}{0.038 \text{ mol}} = 105 \text{ g mol}^{-1}$$

The alkali metal fluoride with the nearest molecular mass is RbF (molar mass = 104.4662 g mol^{-1}).

Chemistry 2e
11: Solutions and Colloids
11.5: Colloids

71. Identify the dispersed phase and the dispersion medium in each of the following colloidal systems: starch dispersion, smoke, fog, pearl, whipped cream, floating soap, jelly, milk, and ruby.

Solution

Colloidal System	Dispersed Phase	Dispersion Medium
starch dispersion	starch	water
smoke	solid particles	air
fog	water	air
pearl	water	calcium carbonate ($CaCO_3$)
whipped cream	air	cream
floating soap	air	soap
jelly	fruit juice	pectin gel
milk	butterfat	water
ruby	chromium(III) oxide (Cr_2O_3)	aluminum oxide (Al_2O_3)

72. Distinguish between dispersion methods and condensation methods for preparing colloidal systems.

Solution

Dispersion methods use a grinding device or some other means to bring about the subdivision of larger particles. Condensation methods bring smaller units together to form a larger unit. For example, water molecules in the vapor state come together to form very small droplets that we see as clouds.

73. How do colloids differ from solutions with regard to dispersed particle size and homogeneity?

Solution

Colloidal dispersions consist of particles that are much bigger than the solutes of typical solutions. Colloidal particles are either very large molecules or aggregates of smaller species that usually are big enough to scatter light. Colloids are homogeneous on a macroscopic (visual) scale, while solutions are homogeneous on a microscopic (molecular) scale.

74. Explain the cleansing action of soap.

Solution

Soap molecules have both a hydrophobic and a hydrophilic end. The charged (hydrophilic) end, which is usually associated with an alkali metal ion, ensures water solubility. The hydrophobic end permits attraction to oil, grease, and other similar nonpolar substances that normally do not dissolve in water but are pulled into the solution by the soap molecules.

75. How can it be demonstrated that colloidal particles are electrically charged?

Solution

If they are placed in an electrolytic cell, dispersed particles will move toward the electrode that carries a charge opposite to their own charge. At this electrode, the charged particles will be neutralized and will coagulate as a precipitate.