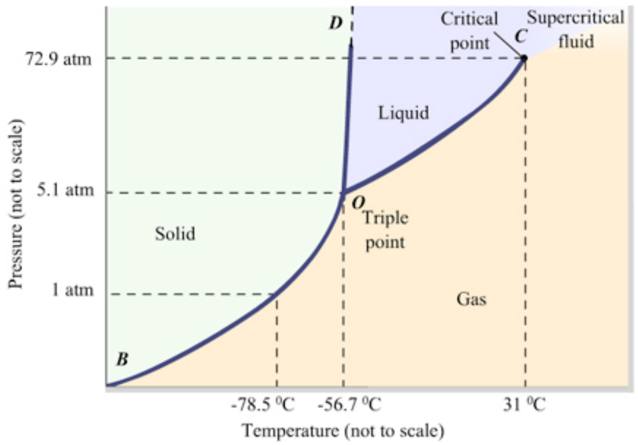
# Problems Chapter 13 (Solutions)

1. In flushing and cleaning columns using in liquid chromatography to remove adsorbed contaminants, a series of solvents is used. Hexane ($C_6H_{14}$), chloroform ($CHCl_3$), methanol ($CH_3OH$) and water are passed through the column in that order. Rationalize the order in terms of intermolecular forces and the mutual solubility (miscibility) of the solvents.

$C_6H_{14}$, a nonpolar liquid, is used to remove relatively nonpolar organic compounds (that are routinely separated on the instrument)
$CHCl_3$, a compound that is more polar than $C_6H_{14}$ (but still not very polar in general) is used to remove $C_6H_{14}$.
$CH_3OH$, a relatively polar organic compound (it can hydrogen bond, but it also has a nonpolar $CH_3$ part), is used to remove $CHCl_3$.
$H_2O$, which can hydrogen bond to $CH_3OH$, is used to remove $CH_3OH$.

2. Explain the following on the basis of the behavior of atoms and/or ions.
   a. Cooking with water is faster in a pressure cooker than in an open pan.
   b. Salt is used on icy roads.
   c. Melted sea ice from the Arctic Ocean produces fresh water.
   d. $CO_2$ (s) (dry ice) does not have a normal boiling point under normal atmospheric conditions, even though $CO_2$ is a liquid in fire extinguishers.
   e. Adding a solute to a solvent extends the liquid phase over a larger temperature range.

a. Recall that liquids will not boil until their vapor pressure reaches surrounding pressure. In a pressure cooker, surrounding pressure will be higher than 1 atm, allowing water to remain liquid over a larger temperature range (and thus allowing faster cooking, as food will cook faster when surrounded by hotter liquid water). On the open pan, the highest temperature of cooking in liquid water that is allowed will be ~100 ˚C (the temperature at which the vapor pressure of water will equal ~1 atm surrounding pressure).
b. Adding salt to water will cause freezing point depression, allowing ice to melt at lower temperatures and preventing falls (which can happen when people/cars move on solid, slippery ice).
c. When a solution begins to freeze, the solid phase is formed from solvent molecules only – the solute particles (here, ions found in seawater) remain dissolved in the liquid phase, causing freezing point depression.
d. See phase diagram of $CO_2$. At P = 1 atm, the liquid-gas line is not crossed – it does not begin until 5.1 atm. The solid-gas line is crossed instead, which is why $CO_2$ sublimes under normal pressure. The pressure in fire extinguishers is therefore at least 5.1 atm, allowing $CO_2$ to stay liquid.



e. Adding a solute to a solvent causes boiling point elevation and freezing point depression. For example, adding 1 mol of solute to 1 kg of water will elevate the boiling point to 100.51 ˚C and lower the freezing point to –1.86 ˚C. So while pure water is liquid over a range of 100 ˚C, water with the concentration of 1 *m* solute is liquid over a range of 102.37 ˚C.

3. The term "proof" is defined as twice the percent by volume of pure ethanol in solution. Thus, a solution that is 95% (by volume) ethanol is 190 proof. What is the molarity of ethanol in a 92-proof ethanol/water solution? Assume the density of ethanol, $C_2H_5OH$, is 0.79 g/cm$^3$ and the density of water is 1.0 g/cm$^3$.

$$\frac{92}{200} \times 100 = 46\%\ by\ volume\ ethanol\ (EtOH)$$

$$\frac{46\ L\ EtOH}{100\ L\ solution} \times \frac{1000\ mL}{1\ L} \times \frac{0.79\ g\ EtOH}{1\ mL\ EtOH} \times \frac{1\ mol\ EtOH}{46.07\ g\ EtOH} = \mathbf{7.9\ M\ EtOH}$$

4. An unknown compound contains only carbon, hydrogen and oxygen. Combustion analysis of the compound gives mass percents of 31.57% C and 5.30% H. The molar mass is determined by measuring the freezing-point depression of an aqueous solution. A freezing point of -5.20 ˚C is recorded for a solution made by dissolving 10.56 g of the compound in 25.0 g water. Determine the empirical formula, molar mass and molecular formula of the compound. Assume that the compound is a nonelectrolyte.

100 – 31.57 – 5.30 = 63.13% O by mass

$$31.57\ g\ C \times \frac{1\ mol\ C}{12.011\ g\ C} = 2.628\ mol\ C$$

$$5.30\ g\ H \times \frac{1\ mol\ H}{1.008\ g\ H} = 5.26\ mol\ H$$

$$63.13\ g\ O \times \frac{1\ mol\ O}{15.999\ g\ O} = 3.946\ mol\ O$$

Divide by mol C → empirical formula $CH_2O_{1.5}$ = **$C_2H_4O_3$**

Molar mass of empirical formula = 2(12.011 g/mol) + 4(1.008 g/mol) + 3(15.999 g/mol) =

76.05 g/mol

$\Delta T_{fp} = K_{fp}m$

5.20˚C = (1.86 ˚C/$m$)$m$ → $m$ = 2.80 mol/kg

$$25.0\ g\ water \times \frac{1\ kg}{1000\ g} \times \frac{2.80\ mol\ compound}{1\ kg\ water} = 0.0699\ mol\ compound$$

$$\frac{10.56\ g\ compound}{0.0699\ mol\ compound} = \mathbf{151\ g/mol}$$

151 g/mol ≈ 2(76.05 g)

So the molecular formula is **$C_4H_8O_6$**.

5. The Henry's law constant for $N_2O$ is 2.4 x $10^{-2}$ mol $kg^{-1}$ $bar^{-1}$. (Nitrous oxide, $N_2O$, laughing gas, is used as an anesthetic.) Determine the mass of $N_2O$ that will dissolve in 500. mL of water, under a $N_2O$ pressure of 1.00 bar. What is the concentration of $N_2O$ in this solution, expressed in ppm (density of $H_2O$ = 1.00 g/mL)?

$S_{N2O} = k_HP_{N2O}$ = (2.4 x $10^{-2}$ mol $kg^{-1}$ $bar^{-1}$)(1.00 bar) = 0.024 mol/kg

$$500.\ mL\ H_2O \times \frac{1.00\ g\ H_2O}{1\ mL\ H_2O} \times \frac{1\ kg}{1000\ g} \times \frac{0.024\ mol\ N_2O}{1\ kg\ H_2O} \times \frac{44.013\ g\ N_2O}{1\ mol\ N_2O} = \mathbf{0.53\ g\ N_2O}$$

$$\frac{0.53\ g\ N_2O}{500.528\ g\ solution} \times 1 \times 10^6\ g\ solution = \mathbf{1.1 \times 10^3\ ppm}$$

6. A compound is known to be a potassium halide, KX. If 4.00 g of the salt is dissolved in exactly 100 g of water, the solution freezes at –1.28 ˚C. Identify the halide ion in this formula.

$\Delta T_{fp} = K_{fp}mi$ → 1.28 ˚C = (1.86 ˚C/$m$)$m$(2) → $m$ = 0.344 mol/kg

$$\frac{0.344\ mol\ KX}{1000\ g\ H_2O} \times 100\ g\ H_2O = 0.0344\ mol\ KX$$

$$\frac{4.00\ g\ KX}{0.0344\ mol\ KX} = 116\ g/mol$$

116 g/mol – 39.1 g/mol K = 77.1 g/mol X

**X = Br**

7. Arrange the following aqueous solutions in order of (i) increasing vapor pressure of water and (ii) increasing boiling point.

a. 0.35 *m* $HOCH_2CH_2OH$ (a nonvolatile solute)
b. 0.50 *m* sugar
c. 0.20 *m* KBr (a strong electrolyte)
d. 0.20 *m* $Na_2SO_4$ (a strong electrolyte)

$$P_{H2O} = x_{H2O}P^{\circ}_{H2O}$$

$$1000\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.015\ g\ H_2O} = 55.51\ mol\ H_2O$$

$$(a)\ x_{H_2O} = \frac{55.51\ mol\ H_2O}{55.51\ mol\ H_2O + 0.35\ mol\ solute} = 0.9937$$

$$(b)\ x_{H_2O} = \frac{55.51\ mol\ H_2O}{55.51\ mol\ H_2O + 0.50\ mol\ sugar} = 0.9911$$

$$(c)\ x_{H_2O} = \frac{55.51\ mol\ H_2O}{55.51\ mol\ H_2O + 0.40\ mol\ ions} = 0.9928$$

$$(d)\ x_{H_2O} = \frac{55.51\ mol\ H_2O}{55.51\ mol\ H_2O + 0.60\ mol\ ions} = 0.9893$$

(Highest VP) (a) > (c) > (b) > (d) (lowest VP)

$$\Delta T_{bp} = K_{bp}mi$$

(a) $mi$ = 0.35 mol/kg

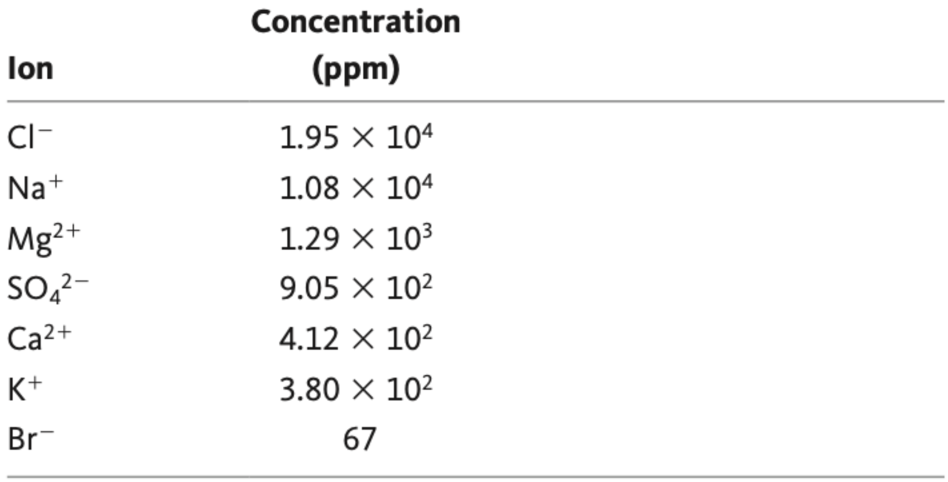
(b) $mi$ = 0.50 mol/kg

(c) $mi$ = 0.40 mol/kg

(d) $mi$ = 0.60 mol/kg

(highest BP) (d) > (b) > (c) > (a)

8. The following table lists the concentrations of the principal ions in seawater.



| Ion | Concentration (ppm) |
|---|---|
| $Cl^-$ | $1.95 \times 10^4$ |
| $Na^+$ | $1.08 \times 10^4$ |
| $Mg^{2+}$ | $1.29 \times 10^3$ |
| $SO_4^{2-}$ | $9.05 \times 10^2$ |
| $Ca^{2+}$ | $4.12 \times 10^2$ |
| $K^+$ | $3.80 \times 10^2$ |
| $Br^-$ | 67 |

a. Calculate the freezing point of seawater.
b. Calculate the osmotic pressure of seawater at 25 ˚C. What is the minimum pressure needed to purify seawater by reverse osmosis? The density of seawater is 1.05 g/mL.

In 1,000,000 g seawater we have:

$$1.95 \times 10^4\ g\ Cl^- \times \frac{1\ mol\ Cl}{35.453\ g\ Cl} = 550.\ mol\ Cl^-$$

$$1.08 \times 10^4\ g\ Na^+ \times \frac{1\ mol\ Na}{22.99\ g\ Na} = 470.\ mol\ Na^+$$

$$1.29 \times 10^3\ g\ Mg^{2+} \times \frac{1\ mol\ Mg}{24.305\ g\ Mg} = 53.1\ mol\ Mg^{2+}$$

$$9.05 \times 10^2\ g\ SO_4{}^{2-} \times \frac{1\ mol\ SO_4{}^{2-}}{96.06\ g\ SO_4{}^{2-}} = 9.42\ mol\ SO_4{}^{2-}$$

$$4.12 \times 10^2\ g\ Ca^{2+} \times \frac{1\ mol\ Ca}{40.078\ g\ Ca} = 10.3\ mol\ Ca^{2+}$$

$$3.80 \times 10^2\ g\ K^+ \times \frac{1\ mol\ K}{30.098\ g\ K} = 9.72\ mol\ K^+$$

$$67\ g\ Br^- \times \frac{1\ mol\ Br}{79.904\ g\ Br} = 0.84\ mol\ Br^-$$

1,000,000 g seawater – ($1.95 \times 10^4$ g $Cl^-$) – ($1.08 \times 10^4$ g $Na^+$) – ($1.29 \times 10^3$ g $Mg^{2+}$) – ($9.05 \times 10^2$ g $SO_4{}^{2-}$) – ($4.12 \times 10^2$ g $Ca^{2+}$) – ($3.80 \times 10^2$ g $K^+$) – (67 g $Br^-$) = $9.666 \times 10^5$ g $H_2O$

$$9.666 \times 10^5 g\ H_2O \times \frac{1\ kg}{1000\ g} = 966\ kg\ H_2O$$

$$m = \frac{(550.mol\ Cl^-)+(470.mol\ Na^+)+(53.1\ mol\ Mg^{2+})+(9.42\ mol\ SO_4{}^{2-})+(10.3\ mol\ Ca^{2+})+(9.72\ mol\ Ca^{2+})+(0.84\ mol\ Br^-)}{966\ kg\ H_2O} =$$

1.14 mol/kg

$\Delta T_{fp} = K_{fp}m$ = (1.86 ˚C/$m$)(1.14 $m$) = 2.12 ˚C

$T_{fp}$ = **–2.12 ˚C**

$$1{,}000{,}000\ g\ seawater\ \times \frac{1\ mL\ seawater}{1.05\ g\ seawater} \times \frac{1\ L}{1000\ mL} = 952\ L\ seawater$$

$$\frac{1.10 \times 10^3\ mol\ ions}{952\ L\ seawater} = 1.16\ M$$

Π = cRT = (1.16 M)(0.08206 L atm $mol^{-1}$ $K^{-1}$)(298.15 K) = **28.3 atm**
One needs to apply pressure greater than 28.3 atm to purify seawater.

9. A typical root beet contains 0.13% of a 75% $H_3PO_4$ solution by mass. How many milligrams of phosphorus are contained in a 12 oz can of this root beer? Assume a solution density of 1.00 g/mL; also, 1 oz = 29.6 mL.

$$12\ oz\ soda \times \frac{29.6\ mL}{1\ oz} \times \frac{1.00\ g\ soda}{1\ mL\ soda} \times \frac{0.13\ g\ H_3PO_4\ solution}{100\ g\ soda} = 0.46\ g\ H_3PO_4\ solution$$

$$0.46\ g\ H_3PO_4\ solution \times \frac{75\ g\ \ H_3PO_4}{100\ g\ H_3PO_4\ solution} \times \frac{30.97\ g\ P}{97.994\ g\ \ H_3PO_4} \times \frac{1000\ mg}{1\ g}$$
$$= \mathbf{110\ mg\ P}$$

10. Suppose you have available 2.50 L of a solution (d = 0.9767 g/mL) that is 13.8% ethanol ($CH_3CH_2OH$) by mass. From this solution you would like to make the maximum quantity of ethanol-water antifreeze solution that will offer protection to –2.0 ˚C. Would you add more ethanol or more water to the solution? What mass of liquid would you add?

$\Delta T_{fp} = K_{fp}m \rightarrow 2.0\ ˚C = (1.86\ ˚C/m)m \rightarrow m = 1.1$ mol/kg

$$2.50\ L\ solution \times \frac{1000\ mL}{1\ L} \times \frac{0.9767\ g}{1\ mL} \times \frac{1\ kg}{1000\ g} = 2.44\ kg\ solution$$

$$2.44\ kg\ solution \times \frac{13.8\ g\ EtOH}{0.100\ kg\ solution} \times \frac{1\ mol\ EtOH}{46.07\ g\ EtOH} = 7.31\ mol\ EtOH$$

$$2.44\ kg\ solution \times \frac{0.862\ kg\ H_2O}{1\ kg\ solution} = 2.10\ kg\ H_2O$$

$$\frac{7.31\ mol\ EtOH}{2.10\ kg\ H_2O} = 3.47 mol/kg$$

Need to add **more $H_2O$** to reduce the molality to 1.1 mol/kg

$$\frac{7.31\ mol\ EtOH}{1.1\ mol\ EtOH/(kg\ H_2O)} = 6.6\ kg\ H_2O$$

Need to add 6.6 kg – 2.10 kg = **4.5 kg $H_2O$**

11. Nitrobenzene, $C_6H_5NO_2$, and benzene, $C_6H_6$, are completely miscible in each other. Other properties of the two liquids are:

| **Substance** | **Freezing point (˚C)** | **$K_f$ (˚C $m^{-1}$)** |
|---|---|---|
| Benzene | 5.5 | 5.12 |
| Nitrobenzene | 5.7 | 8.1 |

It is possible to prepare two different solutions with these two liquids having a freezing point of 0.0 ˚C. What are the compositions of these two solutions, expressed as mass percent nitrobenzene?

We can make one solution where benzene is the solvent (more abundant component) and one – where nitrobenzene is the solvent.

$\Delta T_f = K_f m$

Benzene: 5.5 ˚C = (5.12 ˚C/$m$)($m$) → $m$ = 1.1 mol/kg

$$1.1\ mol\ nitrobenzene \times \frac{123.11\ g}{1\ mol\ nitrobenzene} = 130\ g\ nitrobenzene$$

$$\frac{130\ g\ nitrobenzene}{130\ g\ nitrobenzene + 1000\ g\ benzene} \times 100 = \mathbf{12\%\ nitrobenzene\ by\ mass}$$

Nitrobenzene: 5.7 ˚C = (8.1 ˚C/$m$)($m$) → $m$ = 0.70 mol/kg

$$0.70\ mol\ benzene \times \frac{78.11\ g}{1\ mol\ benzene} = 55\ g\ benzene$$

$$\frac{1000\ g\ nitrobenzene}{55\ g\ benzene + 1000\ g\ nitrobenzene} \times 100 = \mathbf{95\%\ nitrobenzene\ by\ mass}$$

12. A solution contains 750 g of ethanol and 85.0 g of glucose (180 g/mol). The volume of the solution is 810.0 mL. Determine
    a. the density of the solution
    b. the mass percent of sucrose in the solution
    c. the mole fraction of sucrose in the solution
    d. the molality of the solution
    e. the molarity of the solution

$$(a)\frac{750\ g\ EtOH + 85.0\ g\ sucrose}{810.0\ mL} = \mathbf{1.0\ g/mL}$$

$$(b)\frac{85.0\ g\ sucrose}{750\ g\ EtOH + 85.0\ g\ sucrose} \times 100 = \mathbf{10.\%\ sucrose\ by\ mass}$$

$$(c) 750\ g\ EtOH \times \frac{1\ mol}{46.07\ g\ EtOH} = 16\ mol\ EtOH$$

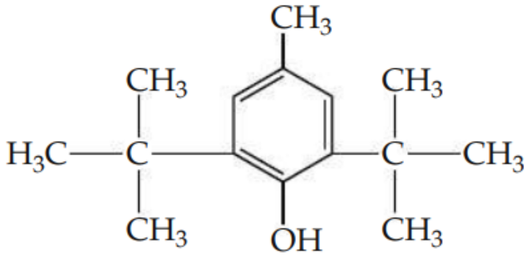
$$85.0\ g\ sucrose \times \frac{1\ mol}{180\ g\ sucrose} = 0.47\ mol\ sucrose$$

$$\frac{0.47\ mol\ sucrose}{16\ mol\ EtOH + 0.47\ mol\ sucrose} = \mathbf{0.028}$$

$$(d) \frac{0.47\ mol\ sucrose}{0.75\ kg\ EtOH} = \mathbf{0.63\ mol/kg}$$

$$(e) \frac{0.47\ mol\ sucrose}{0.8100\ L\ solution} = \mathbf{0.58\ M}$$

13. Butylated hydroxytoluene (BHT) has the following molecular structure:



BHT

It is widely used as a preservative in a variety of foods, including dried cereals. Based on its structure, would you expect BHT to be more soluble in water or hexane ($C_6H_{14}$)? Explain.

Although BHT can form hydrogen bonds through the –OH group, the majority of the molecule is nonpolar (C–H bonds). So BHT is expected to be more soluble in nonpolar hexane.

14. The presence of the radioactive gas radon (Rn) in well water presents a possible health hazard in parts of the United States.
    a. Assuming that the solubility of radon in water with 1 atm pressure of the gas over the water at 30 ˚C is $7.27 \times 10^{-3}$ M, what is the Henry's law constant for radon in water at this temperature?
    b. A sample consisting of various gases contains $3.5 \times 10^{-6}$ mole fraction of radon. This gas at a total pressure of 32 atm is shaken with water at 30 ˚C. Calculate the molar concentration of radon in the water.

(a) $S_g = k_H P_g \rightarrow 7.27 \times 10^{-3}$ M = $k_H$(1 atm) → $k_H$ = **$7 \times 10^{-3}$ M atm$^{-1}$**

(b) $P_{Ar} = P_{total}x_{Ar} \rightarrow P_{Ar} = (32\text{ atm})(3.5 \times 10^{-6}) = 1.1 \times 10^{-4}$ atm
$S_g = (7 \times 10^{-3}\text{ M atm}^{-1})(1.1 \times 10^{-4}\text{ atm})$ = **$8 \times 10^{-7}$ M**

15. The solubility of carbon tetrachloride ($CCl_4$) in water at 25 ˚C is 1.2 g/L. The solubility of chloroform ($CHCl_3$) at the same temperature is 10.1 g/L. Why is chloroform almost ten times more soluble in water than carbon tetrachloride?

    $CCl_4$ is a very symmetrical molecule (tetrahedral geometry, all vertices of the tetrahedron occupied by the same kind of atoms) – the symmetry makes it quite nonpolar. $CHCl_3$ is less symmetrical (one vertex, H, is not like the others, Cl) which maks it more polar and thus more soluble in polar water.

16. Potassium perchlorate ($KClO_4$) has a lattice energy of –599 kJ/mol and a heat of hydration of –548 kJ/mol. Find the heat of solution for $KClO_4$ and determine the temperature change that occurs when 10.0 g of $KClO_4$ is dissolved with enough water to make 100.0 mL of solution. Assume a heat capacity of 4.05 J g$^{-1}$ ˚C$^{-1}$) for the solution and a density of 1.05 g/mL.

    Lattice energy: $K^+ (g) + ClO_4^- (g) \rightarrow KClO_4 (s)$ $\Delta H^\circ = -599$ kJ/mol
    Heat of hydration: $K^+ (g) + ClO_4^- (g) \rightarrow KClO_4 (aq)$ $\Delta H^\circ = -548$ kJ/mol
    Heat of solution: $KClO_4 (s) \rightarrow KClO_4 (aq)$
    $\Delta H^\circ_{soln}$ = 599 kJ/mol – 548 kJ/mol = **+51 kJ/mol**

$$10.0\ g\ KClO_4 \times \frac{1\ mol}{138.55\ g} \times \frac{51\ kJ}{1\ mol} = 3.7\ kJ$$

$$100.0\ mL \times \frac{1.05\ g}{1\ mL} = 105\ g$$

$q = mC\Delta T \rightarrow 3700\text{ J} = (105\text{ g})(4.05\text{ J g}^{-1}\ ^\circ\text{C}^{-1})\Delta T \rightarrow \Delta T = 8.7\ ^\circ\text{C}$

Since $KClO_4$ absorbs energy when it dissolves, the thermometer will register a drop by 8.7 ˚C.

**$\Delta T = -8.7$ ˚C**

17. The Safe Drinking Water Act (SDWA) sets a limit for mercury – a toxin to the central nervous system – at 0.0020 ppm by mass. Water suppliers must periodically test their water to ensure that mercury levels do not exceed this limit. Suppose water becomes contaminated with mercury at twice the legal limit (0.0040 ppm). How much of this water would a person have to consume to ingest 50.0 mg of mercury?

$$50.0\ mg\ Hg \times \frac{1\ g}{1000\ mg} \times \frac{1{,}000{,}000\ g\ water}{0.0040\ g\ Hg} \times \frac{1\ kg}{1000\ g}$$
$$= 1.2 \times 10^4\ kg\ water\ (\sim 1.2 \times 10^4\ L)$$

18. Magnesium citrate, $Mg_3(C_6H_5O_7)_2$, belongs to a class of laxatives called *hyperosmotics*, which cause rapid emptying of the bowel. When a concentrated solution of magnesium citrate is consumed, it passes through the intestines, drawing water and promoting diarrhea, usually within 6 hours. Calculate the osmotic pressure of a magnesium citrate laxative solution containing 28.5 g of magnesium citrate in 235 mL of solution at 37 ˚C (approximate body temperature). Assume complete dissociation of the ionic compound.

$$28.5\ g\ Mg_3(C_6H_5O_7)_2 \times \frac{1\ mol\ Mg_3(C_6H_5O_7)_2}{451.113\ g} \times \frac{5\ mol\ ions}{1\ mol\ Mg_3(C_6H_5O_7)_2} = 0.316\ mol\ ions$$
$$\frac{0.665\ mol\ ions}{0.235\ L} = 1.34\ M$$

$\Pi = cRT = (1.34\text{ M})(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(310.15\text{ K})$ **= 34.2 atm**

19. A solution of a nonvolatile solute in water has a boiling point of 375.3 K. Calculate the vapor pressure of water above this solution at 338 K. The vapor pressure of pure water at this temperature is 0.2467 atm.

$\Delta T_b = K_b m \rightarrow (375.3\text{ K} - 373.15\text{ K}) = (0.51\text{ K}/m)m \rightarrow m = 4.2\text{ mol/kg}$

$$1000\ g\ H_2O \times \frac{1\ mol}{18.015\ g} = 55.5\ mol\ H_2O$$

$$x_{H_2O} = \frac{55.5\ mol}{55.5\ mol + 4.2\ mol} = 0.929$$

$P_{H2O} = x_{H2O}P^{\circ}_{H2O} = (0.929)(0.2467\text{ atm})$ = **0.229 atm**

20. The vapor pressure of carbon tetrachloride, $CCl_4$, is 0.354 atm, and the vapor pressure of chloroform, $CHCl_3$, is 0.526 atm at 316 K. A solution is prepared from equal masses of these two compounds at this temperature. Calculate the mole fraction of the chloroform in the vapor above the solution. If the vapor above the original solution is condensed and isolated into a separate flask, what will the vapor pressure of chloroform be above this new solution?

Let's assume that we made a solution of 50.0 g $CCl_4$ and 50.0 g $CHCl_3$.

$$50.0\ g\ CCl_4 \times \frac{1\ mol\ CCl_4}{153.82\ g} = 0.325\ mol\ CCl_4$$

$$50.0\ g\ CHCl_3 \times \frac{1\ mol\ CHCl_3}{119.38\ g} = 0.419\ mol\ CHCl_3$$

$$x_{CCl_4} = \frac{0.325\ mol}{0.325\ mol + 0.419\ mol} = 0.437$$

$$x_{CHCl_3} = \frac{0.419\ mol}{0.325\ mol + 0.419\ mol} = 0.563$$

$P_{CCl4} = x_{CCl4}P°_{CCl4}$ = (0.437)(0.354 atm) = 0.155 atm
$P_{CHCl3} = x_{CHCl3}P°_{CHCl3}$ = (0.563)(0.526 atm) = 0.296 atm

$$x_{CHCl_3}\ in\ vapor = \frac{0.296\ atm}{0.155\ atm + 0.296\ atm} = \mathbf{0.657}$$

After recondensing this vapor:

$P_{CHCl3} = x_{CHCl3}P°_{CHCl3}$ = (0.657)(0.526 atm) = **0.345 atm**