

Chapter 2

Solutions

1. How many grams of CuSO_4 are required to make 650. mL of a 0.115-M solution?

$$650. \text{ mL solution} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \frac{0.115 \text{ mol CuSO}_4}{\text{L solution}} \times \frac{159.608 \text{ g CuSO}_4}{1 \text{ mol CuSO}_4} = 11.9 \text{ g CuSO}_4$$

3. How many grams of Na_2SO_4 are required to make 90.0 mL of a solution that is 0.200 M in Na^{1+} ?

$$90.0 \text{ mL solution} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \frac{0.200 \text{ mol Na}^{1+}}{\text{L solution}} \times \frac{1 \text{ mol Na}_2\text{SO}_4}{2 \text{ mol Na}^{1+}} \times \frac{142.042 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} = 1.28 \text{ g Na}_2\text{SO}_4$$

5. How many mmoles of chloride ion are in 55.0 mL of 0.0688 M BaCl_2 solution?

$$55.0 \text{ mL solution} \times \frac{0.0688 \text{ mmol BaCl}_2}{1 \text{ mL solution}} \times \frac{2 \text{ mmol Cl}^{1-}}{1 \text{ mmol BaCl}_2} = 7.57 \text{ mmol Cl}^{1-}$$

7. How many mL of 0.124 M Ba(OH)_2 are required to deliver 38.6 mmol of hydroxide ion ?

$$38.6 \text{ mmol OH}^{1-} \times \frac{1 \text{ mmol Ba(OH)}_2}{2 \text{ mmol OH}^{1-}} \times \frac{1 \text{ mL solution}}{0.124 \text{ mmol Ba(OH)}_2} = 156 \text{ mL solution}$$

9. How many grams of sodium should be added to 15.0 g Hg to make a mixture in which the mole fraction of sodium is 0.800?

$$\text{moles of mercury} = 15.0 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} = 0.0748 \text{ mol Hg}$$

$$\text{Use moles of Hg and the mole fraction equation: } X_{\text{Na}} = 0.800 = \frac{n}{0.0748 + n},$$

$$\text{rearrange to eliminate denominator: } 0.0598 + 0.800n = n \Rightarrow n = \frac{0.0598}{0.200} = 0.299 \text{ mol Na}$$

$$\text{Calculate the grams of Na: } 0.299 \text{ mol Na} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 6.88 \text{ g Na}$$

11. The density of a 1.140-M solution of NH_4Cl at 20 °C is 1.0186 g/mL. What mass of water does 100. mL of this solution contain?

The total mass of this solution is $1.0186 \text{ g/mL} \times 100. \text{ mL} = 102$ (three significant figures)

$$\text{The mass of NH}_4\text{Cl in this solution is } 0.100 \text{ L} \times \frac{1.140 \text{ mol NH}_4\text{Cl}}{1 \text{ L NH}_4\text{Cl}} \times \frac{53.49 \text{ g}}{1 \text{ mol NH}_4\text{Cl}} = 6.10 \text{ g NH}_4\text{Cl}$$

The mass of water is then the difference between the total mass of the solution and the mass of the NH_4Cl :
 $102 \text{ g solution} - 6 \text{ g NH}_4\text{Cl} = 96 \text{ g water}$.

13. What mass of $\text{Fe(ClO}_4)_3$ is required to make 275 mL of a solution that is 0.100 M in ClO_4^{1-} ?

$$0.0275 \text{ L} \times \frac{0.100 \text{ mol ClO}_4^{1-}}{1 \text{ L}} \times \frac{1 \text{ mol Fe(ClO}_4)_3}{3 \text{ mol ClO}_4^{1-}} \times \frac{354 \text{ g Fe(ClO}_4)_3}{1 \text{ mol}} = 3.25 \text{ g Fe(ClO}_4)_3$$

15. How many grams of CaCl_2 should be added to 50.0 g of water to make a solution in which the mole fraction of Cl^{1-} is 0.150?

$$\text{First get the number of moles of water: } 50.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = 2.78 \text{ mol H}_2\text{O}$$

If n mol CaCl_2 are dissolved, n mol Ca^{2+} and $2n$ mol Cl^{1-} form in solution, or $3n$ mol ions.

The mole fraction of chloride ion is 0.150, so

$$0.150 = \frac{2n \text{ mol Cl}^{1-}}{3n \text{ mol ions} + 2.78 \text{ mol of H}_2\text{O}}$$

$$\text{Rearrange and solve for } n: 0.450n + 0.417 = 2n \Rightarrow n = 0.269 \text{ mol CaCl}_2$$

$$0.269 \text{ mol CaCl}_2 \times \frac{111 \text{ g CaCl}_2}{1 \text{ mol}} = 29.9 \text{ g CaCl}_2$$

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17. A concentrated solution of phosphoric acid is 75% H_3PO_4 by mass and has a density of 1.57 g/mL.

a) What is the molarity of the concentrated phosphoric acid?

Assume 100. g of concentrated phosphoric acid then 75% translates to 75 g of H_3PO_4 . Calculate the moles of H_3PO_4 .

$$75 \text{ g H}_3\text{PO}_4 \times \frac{1 \text{ mol H}_3\text{PO}_4}{97.994 \text{ g}} = 0.77 \text{ mol H}_3\text{PO}_4$$

Use the density to calculate the volume of 100 g of concentrated phosphoric acid.

$$100. \text{ g conc H}_3\text{PO}_4 \times \frac{1 \text{ mL}}{1.57 \text{ g}} = 63.7 \text{ mL conc H}_3\text{PO}_4$$

The molarity of the solution can then be calculated, $\frac{0.77 \text{ mol H}_3\text{PO}_4}{0.0637 \text{ L solution}} = 12. \text{ M H}_3\text{PO}_4$

b) How many mL of the concentrated acid would be required to prepare 1.5 L of a 0.20-M solution?

Use dilution equation, $C_i V_i = C_f V_f$: $(0.20 \text{ M H}_3\text{PO}_4)(1.5 \text{ L}) = (12 \text{ M H}_3\text{PO}_4)(V_f)$

$$V_f = \frac{(1.5 \text{ L})(0.20 \text{ M H}_3\text{PO}_4)}{12 \text{ M H}_3\text{PO}_4} = 0.025 \text{ L} = 25 \text{ mL}$$

19. A 60/40 solder is a solution which is 60% Pb and 40% Sn. What are the two mole fractions in the solder?

Work with 100. g of solder, which contains 60. g of Pb and 40. g of Sn.

Calculate the moles of Pb and the moles of Sn

$$60. \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.19 \text{ g}} = 0.290 \text{ mol Pb} \quad \text{and} \quad 40. \text{ g Sn} \times \frac{1 \text{ mol Sn}}{118.71 \text{ g}} = 0.3370 \text{ mol Sn}$$

Use Equation 2.2 to determine the mole fraction of Pb. The mole fraction of Sn can be determined by difference because the sum of the two mole fractions must be one.

$$X_{\text{Pb}} = \frac{0.290}{0.290 + 0.3370} = 0.46 \quad \text{and} \quad X_{\text{Sn}} = 1 - X_{\text{Pb}} = 1 - 0.46 = 0.54$$

21. What is the molality of a sucrose solution if its mole fraction is 0.025?

A mole fraction of 0.025 means that there are 0.025 mol of sucrose in 1 mole of particles (sucrose + H_2O).

Thus, the number of moles of water present must be 1.000 mol total – 0.025 mol sucrose = 0.975 mol H_2O .

Thus, the solution has 0.025 mol sucrose for each 0.975 mol H_2O . Starting here, we can more easily convert between concentrations.

$$\frac{0.025 \text{ mol sucrose}}{0.975 \text{ mol H}_2\text{O}} = \frac{0.025 \text{ mol sucrose}}{(0.975 \text{ mol H}_2\text{O}) \left(\frac{18.0 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} \right) \left(\frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} \right)} = \frac{0.025 \text{ mol sucrose}}{0.01755 \text{ kg H}_2\text{O}} = 1.4 \text{ m}$$

23. Pure, crystalline Si is a poor conductor, but it can be converted into a p-type semiconductor by dissolving 1.0 ppm Ga in the Si crystal. This is a process called "doping" in the semiconductor industry. Given that the molar mass of Si is 28.086 and of Ga is 69.72, what is the mole fraction of Ga in the crystal? The density of the p-type Si semiconductor is 2.33 g/cm³.

Use the definition of ppm and assume a 1.0 g sample of semiconductor to obtain the mass of gallium.

$$\text{ppm Ga} = \frac{\text{g Ga}}{\text{g semiconductor}} \times 10^6 \Rightarrow \text{g Ga} = \frac{1.0 \text{ ppm} \times 1.0 \text{ g semiconductor}}{10^6} \text{ g} = 1.0 \times 10^{-6} \text{ g Ga}$$

Then calculate the moles of each component: $\text{Moles Ga} = \frac{1.0 \times 10^{-6} \text{ g}}{69.72 \text{ g/mol}} = 1.4 \times 10^{-8} \text{ mol Ga}$

The mass of the Si would then be calculated as the difference between the total mass, 1.0 g, and the mass of Ga. However, the mass of Ga is negligibly small, so it can be ignored. Thus, the mass of silicon is 1.0 g. Next,

determine the moles of silicon: $\text{moles Si} = 1.0 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.086 \text{ g Si}} = 0.0356 \text{ mol Si}$

$$X_{\text{Ga}} = \frac{\text{mole Ga}}{\text{mole Ga} + \text{mole Si}} = \frac{1.4 \times 10^{-8} \text{ mol Ga}}{1.4 \times 10^{-8} \text{ mol Ga} + 0.0356 \text{ mol Si}} = 4.02 \times 10^{-7}$$

25. What is the molality of a 2.06-M solution of H₂SO₄ that has a density of 1.124 g/mL?

Assume a total volume of 1 L, which has a total mass of (1000 mL)(1.124 g/mL) = 1124 g = 1.124 kg.

The mass of H₂SO₄ in 1 L of solution is (1 L)(2.06 mol/L)(98.1 g/mol) = 202 g = 0.202 kg H₂SO₄

The mass of water is obtained by difference: 1.124 kg solution – 0.202 kg H₂SO₄ = 0.922 kg H₂O

$$\text{the molality is: } m = \frac{2.06 \text{ mol H}_2\text{SO}_4}{0.922 \text{ kg H}_2\text{O}} = 2.23 \text{ m}$$

27. What is the molarity of a solution prepared by adding 30.0 mL of water to 57.0 mL of a 0.114 M CuSO₄ solution? Assume additive volumes.

Calculate the moles of CuSO₄ and divide by the total volume.

$$\frac{0.114 \text{ mol CuSO}_4}{\text{L H}_2\text{O}} \times 0.057 \text{ L H}_2\text{O} = 6.50 \times 10^{-3} \text{ mol CuSO}_4; \quad V = \frac{6.50 \times 10^{-3} \text{ mol CuSO}_4}{0.030 \text{ L H}_2\text{O} + 0.057 \text{ L H}_2\text{O}} = 0.0747 \text{ M}$$

Thus, 10 mL of 12 M HCl should be diluted to 240 mL to prepare a 0.50 M solution.

29. What volume of an 18.0 M "stock solution" of H₂SO₄ is needed in order to make 100. mL of a 3.0 M H₂SO₄ solution?

Using Equation 2.5, C_iV_i = C_fV_f, to solve for the volume of the H₂SO₄ needed to make 100 mL of the 3.0 M solution:

$$V \times 18.0 \text{ M H}_2\text{SO}_4 = 100 \text{ mL} \times 3.0 \text{ M H}_2\text{SO}_4$$

$$\text{Rearranging the equation and solving for } V \text{ yields: } V = \frac{100 \text{ mL} \times 3.0 \text{ M H}_2\text{SO}_4}{18.0 \text{ M H}_2\text{SO}_4} = 17 \text{ mL H}_2\text{SO}_4$$

17 mL of the stock solution should be diluted to a total volume of 100 mL.

31. What is the [H¹⁺] in a solution prepared by diluting 20.0 mL of 12.0 M HCl to 350. mL?

Use the dilution equation: C_iV_i = C_fV_f

$$(12.0 \text{ M})(20.0 \text{ mL}) = C_f(350 \text{ mL}), \text{ so } C_f = \frac{(12.0)(20.0) \text{ mmol H}^{1+}}{350 \text{ mL}} = 0.69 \text{ M}$$

33. What is the molarity of the chloride ion in a solution prepared by mixing 75 mL of 0.20 M NaCl and 55 mL of 0.15 M MgCl₂?

$$[\text{Cl}^{1-}] = \frac{\text{mmol Cl}^{1-}}{\text{total milliliters}} = \frac{(75 \text{ mL})(0.20 \text{ mmol Cl}^{1-}/\text{mL}) + (55 \text{ mL})(0.15 \text{ mmol MgCl}_2/\text{mL})(2 \text{ mmol Cl}^{1-}/\text{mmol MgCl}_2)}{75 \text{ mL} + 55 \text{ mL}}$$

$$[\text{Cl}^{1-}] = \frac{31.5 \text{ mmol Cl}^{1-}}{130 \text{ mL}} = 0.24 \text{ M}$$

35. 100.0 mL of a stock solution of hydrochloric acid was diluted to 250.0 mL. A 10.0 mL of the resulting solution was then diluted 250.0 mL. The final solution was prepared by diluting 30.0 mL of this solution to 100.0 mL. A 40.0-mL sample of the final solution was titrated with 0.0887 M NaOH. If the titration required 32.6 mL of the base, what is the concentration of the original stock solution.

The reaction is H¹⁺ + OH¹⁻ → H₂O, so the acid:base stoichiometry is 1:1, and we can use M_AV_A = M_BV_B to determine the concentration of the acid in the diluted solution.

$$M_A = \frac{(0.0887)(32.6) \text{ mmol}}{40.0 \text{ mL}} = 0.0723 \text{ M}$$

This is the concentration of the 40.0 mL sample, which is the same concentration as the 100.0 mL solution it came from.

Next, use the dilution equation to take 100 mL of 0.0723 M solution to the original solution.

C_f = C_i × $\frac{V_i}{V_f}$. Each successive concentration is multiplied by the ratio of volumes to yield the following:

$$C = (0.0723 \text{ M}) \times \frac{100.0 \text{ mL}}{30.0 \text{ mL}} \times \frac{250.0 \text{ mL}}{10.0 \text{ mL}} \times \frac{250.0 \text{ mL}}{100.0 \text{ mL}} = 15.1 \text{ M}$$

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37. A stock solution that is 0.4762 M undergoes the following successive dilutions: 1:3, 2:5, 1:10, and 2:15. What is the concentration of the final solution?

$$c = 0.4762 \text{ M} \left(\frac{1}{3}\right) \left(\frac{2}{5}\right) \left(\frac{1}{10}\right) \left(\frac{2}{15}\right) = 8.466 \times 10^{-4} \text{ M} = 0.8466 \text{ mM}$$

39. Formaldehyde can be measured in aqueous solutions by reaction with 1,3,5-trihydroxybenzene to produce a red-orange dye. The absorbance is measured with a spectrophotometer at a 470 nm wavelength. A standard solution was prepared by dissolving 1.50 mL of formaldehyde gas (at 1.00 atm pressure and 25 °C) in 1.00 L of water. This solution was treated with an excess of 1,3,5-trihydroxybenzene. A couple of milliliters of the resulting solution was placed in a spectrophotometric cell with a thickness of 1.00 cm, and the absorbance was measured to be 0.967. A 1.00-L sample of rainwater was similarly treated with 1,3,5-trihydroxybenzene and the absorbance was found to be 0.426 in the same spectrophotometric cell. What is the concentration of formaldehyde in the rain water?

First the concentration of the formaldehyde must be found for the standard solution in units of mols per liter. To do this, use the ideal gas law, $PV = nRT$, to calculate mols formaldehyde, then divide by the total volume of 1.0 L to get the concentration:

$$n = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 0.0015 \text{ L}}{298 \text{ K} \times 0.082057 \text{ L} \cdot \text{atm/mol} \cdot \text{K}} = 6.13 \times 10^{-5} \text{ mols formaldehyde}$$

$$\text{Molarity} = \frac{\text{moles formaldehyde}}{\text{L solution}} = \frac{6.13 \times 10^{-5} \text{ mols formaldehyde}}{1.00 \text{ L}} = 6.13 \times 10^{-5} \text{ M}$$

The absorbances can be determined from Beer's law (Eq. 2.6) as: $A_{\text{std}} = \epsilon l c_{\text{std}}$ $A_{\text{sample}} = \epsilon l c_{\text{sample}}$ where A_{std} is the absorbance of the standard and A_{sample} is the absorbance of the sample. However, the absorbance of the same complex is being measured and the two measurements use the same path length, so the values of ϵ and l are the same in both the standard and the sample solutions. Divide the A_{std} expression by the A_{sample} expression to eliminate ϵ and l .

$$\frac{A_{\text{std}}}{A_{\text{sample}}} = \frac{c_{\text{std}}}{c_{\text{sample}}} \Rightarrow c_{\text{sample}} = c_{\text{std}} \times \frac{A_{\text{sample}}}{A_{\text{std}}} = 6.13 \times 10^{-5} \text{ M} \times \frac{0.426}{0.967} = 2.70 \times 10^{-5} \text{ M}$$

Recall that $\text{ppm} = \frac{\text{g solute}}{\text{g solution}} \times 10^6$ and that 1000. mL water has a mass of 1000. g. Therefore:

$$\frac{2.70 \times 10^{-5} \text{ mols CH}_2\text{O}}{1 \text{ L}} \times \frac{30.03 \text{ g CH}_2\text{O}}{1 \text{ mol CH}_2\text{O}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1000 \text{ mL}}{1000 \text{ g}} \times 10^6 = 0.811 \text{ ppm}$$

41. $\text{Ru}(\text{bpy})_3^{2+}$ (bpy = 2,2'-bipyridine) is being investigated for use in solar energy conversion because it is deeply colored and has interesting redox capabilities in the excited state. A solution was prepared by dissolving 58.2 mg $\text{Ru}(\text{bpy})_3(\text{ClO}_4)_2$ ($M_m = 612 \text{ g/mol}$) in enough water to make 100.0 mL of solution. However, the resulting solution absorbed too strongly at 450 nm to be measured, so 5.0 mL of the solution was diluted to 100.0 mL. The absorbance of the final solution at 450 nm in a 1.00-cm cell was determined to be 0.571.

- a) What is the molar absorptivity of $\text{Ru}(\text{bpy})_3^{2+}$ at 450 nm?

$$58.2 \text{ mg Ru}(\text{bpy})_3^{2+} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol}}{612 \text{ g}} = 9.51 \times 10^{-5} \text{ mol Ru}(\text{bpy})_3^{2+}$$

$$[\text{Ru}(\text{bpy})_3^{2+}] = \frac{9.51 \times 10^{-5} \text{ mol Ru}(\text{bpy})_3^{2+}}{0.100 \text{ L}} = 9.51 \times 10^{-4} \text{ M}$$

Use $M_1V_1 = M_2V_2$ to account for dilution.

$$(9.51 \times 10^{-4} \text{ M})(0.005 \text{ L}) = (M_2)(0.100 \text{ L}) \Rightarrow M_2 = 4.75 \times 10^{-5} \text{ M}$$

Use $A = \epsilon l c$ to solve for ϵ

$$0.571 = \epsilon (1.00)(4.75 \times 10^{-5} \text{ M}) \Rightarrow \epsilon = 1.20 \times 10^4 \text{ M}^{-1} \cdot \text{cm}^{-1}$$

- b) What is the concentration of $\text{Ru}(\text{bpy})_3^{2+}$ in a solution with an absorbance of 0.885 at 450 nm?

$$A = \epsilon l c \Rightarrow 0.885 = (1.20 \times 10^4)(1.00)(c) \Rightarrow c = 7.37 \times 10^{-5} \text{ M} = [\text{Ru}(\text{bpy})_3^{2+}]$$

43. What is the colligative molality of a solution prepared by dissolving 15.2 g of AlCl_3 in 155 mL of water?

Determine the molality:

$$\text{Mass of water} = (155 \text{ mL})(1.00 \text{ g/mL})(1 \text{ kg}/1000 \text{ g}) = 0.155 \text{ kg water}$$

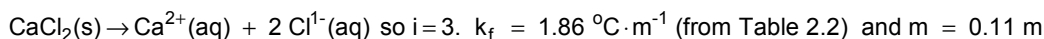
$$\text{Moles of } \text{AlCl}_3: 15.2 \text{ g } \text{AlCl}_3 \times \frac{1 \text{ mol } \text{AlCl}_3}{133.3 \text{ g } \text{AlCl}_3} = 0.114 \text{ mol } \text{AlCl}_3$$

$$\text{Molality of } \text{AlCl}_3 \text{ in solution} = \frac{0.114 \text{ mol } \text{AlCl}_3}{0.155 \text{ kg } \text{H}_2\text{O}} = 0.736 \text{ m}$$

$\text{AlCl}_3 \rightarrow \text{Al}^{3+} + 3\text{Cl}^{-}$, so the van't Hoff factor is $i = 4$; $m_c = im = 4(0.736) = 2.94 \text{ m}$

45. What is the freezing point of a 0.11 m aqueous CaCl_2 solution?

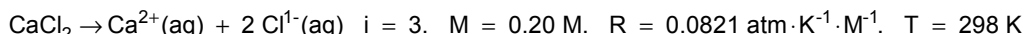
The solution can be started by calculating the freezing point depression:



$$\Delta T_f = ik_f m = (3)(1.86 \text{ }^\circ\text{C} \cdot \text{m}^{-1})(0.11 \text{ m}) \Rightarrow \Delta T_f = 0.61 \text{ }^\circ\text{C} \Rightarrow \text{fp} = 0.00 \text{ }^\circ\text{C} - 0.61 \text{ }^\circ\text{C} = -0.61 \text{ }^\circ\text{C}$$

47. What is the osmotic pressure of a 0.20 M CaCl_2 solution at 298 K?

The osmotic pressure can be calculated as follows:



$$\Pi = iMRT = (3)(0.20 \text{ M})(0.0821 \text{ atm} \cdot \text{K}^{-1} \cdot \text{M}^{-1})(298 \text{ K}) = 15 \text{ atm}$$

49. 1.0 g of CaCl_2 is dissolved in 100 mL of water to produce a solution with a density 1.1 g/mL.**a) What is the vapor pressure of the solution at 20 °C? The vapor pressure of water at 20 °C is 17.5 mm Hg.**

Use equation 2.10b to calculate the vapor pressure of the solution, but first the mole fraction of solvent must be determined. Any amount of solvent can be used to the numbers of moles, so we assume 100 mL. In 100 mL of solution, there are

$$100. \text{ mL water} \times \frac{1 \text{ g water}}{1 \text{ mL water}} \times \frac{1 \text{ mol water}}{18.02 \text{ g water}} = 5.549 \text{ mol water} \quad \text{and}$$

CaCl_2 is ionic, so it dissolves as ions: $\text{CaCl}_2 \rightarrow \text{Ca}^{2+} + 2\text{Cl}^{-}$. Thus, the van't Hoff factor is $i = 3$.

$$1.0 \text{ g } \text{CaCl}_2 \times \frac{1 \text{ mol } \text{CaCl}_2}{110.98 \text{ g } \text{CaCl}_2} \times \frac{3 \text{ mol particles}}{1 \text{ mol } \text{CaCl}_2} = 0.027 \text{ mol particles}$$

$$\text{The mole fraction of water is } X_{\text{solvent}} = \frac{5.549 \text{ mol water}}{(5.549 \text{ mol water} + 0.027 \text{ mol solute particles})} = 0.995$$

so the vapor pressure of the solution is $P_{\text{solution}} = X_{\text{solvent}} \times P_{\text{solvent}}^{\circ} = 0.995 \times 17.5 \text{ mm Hg} = 17.4 \text{ mm Hg}$

b) What is the boiling point of the solution?

To calculate the boiling point of the solution, we must first find the colligative molality of the CaCl_2 solution. From part (a), we already know that there is 0.027 moles of particles (ions) are in 100 mL of solution, so we must now determine the kilograms of solvent.

$$100. \text{ mL water} \times \frac{1 \text{ g water}}{1 \text{ mL water}} \times \frac{1 \text{ kg water}}{1000 \text{ g water}} = 0.10 \text{ kg water}$$

Therefore, the colligative molality is found to be $m_c = (0.027 \text{ mol particles})/(0.10 \text{ kg water}) = 0.27 \text{ m}$.

Now, using Table 2.3 to find k_b for water and using Eq. 2.9 (Note: in this case the solute dissociates into 3 components; one Ca^{2+} and two Cl^{-} ions, therefore, $i = 3$):

$$\Delta T_b = ik_b m = k_b m_c = (0.512 \text{ }^\circ\text{C}/\text{m})(0.27 \text{ m}) = 0.14 \text{ }^\circ\text{C}; \quad T_b = 100 + \Delta T_b = 100.14 \text{ }^\circ\text{C}$$

c) What is the freezing point of the solution?

The procedure for this problem is very similar to that in part (b).

$$\Delta T_f = i \times k_f \times m = k_f m_c = (1.86 \text{ }^\circ\text{C}/\text{m})(0.27 \text{ m}) = 0.50 \text{ }^\circ\text{C}; \quad T_f = 0.0 - \Delta T_f = -0.50 \text{ }^\circ\text{C}$$

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- d) What osmotic pressure would develop at 25 °C when the solution is placed in contact with pure water at a semipermeable membrane?

For this problem we need to calculate the molarity of the CaCl_2 . From part (a), we know that there is 0.027 moles of particles in solution, so the colligative molarity of particles in the solution is.

$$M_c = (0.027 \text{ mol particles}) / (0.100 \text{ L water}) = 0.27 \text{ M}$$

$$\text{Using Equation 2.13, we obtain } \Pi = M_c RT = (0.27 \text{ M})(0.0821 \text{ atm}\cdot\text{K}^{-1}\cdot\text{M}^{-1})(298 \text{ K}) = 6.6 \text{ atm}$$

51. The colligative molarity of living cells is typically 0.3 M. What osmotic pressure would develop when the cells are placed in pure water at 37 °C? What effect would this have on the cell?

$$\Pi = M_c RT = (0.3 \text{ mol}\cdot\text{L})(0.0821 \text{ L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})(310 \text{ K}) = 8 \text{ atm}$$

Water moves from higher concentration to lower concentration, so the 8 atm is the pressure pushing water into the cell. Consequently, water will flow into the cell, and since the cell membrane cannot withstand such pressures, it would most likely rupture.

53. 15.8 mg of a protein is dissolved in enough water to make a 5.00-mL solution. What is the molar mass of the protein if the osmotic pressure of the solution at 15 °C is 4.65 torr?

Proteins do not dissociate in water ($i = 1$). $T = 15 + 273 = 288 \text{ K}$. $R = 0.0821 \text{ atm}\cdot\text{K}^{-1}\cdot\text{M}^{-1}$.

$$\Pi = iMRT = 4.65 \text{ torr} \Rightarrow M = \frac{\Pi}{iRT} = \frac{4.65 \text{ torr} (1 \text{ atm}/760 \text{ torr})}{(1)(0.0821 \text{ atm}\cdot\text{K}^{-1}\cdot\text{M}^{-1})(288 \text{ K})} = 2.59 \cdot 10^{-4} \text{ M}$$

Notice that the pressure was converted from torr to atm.

The number of moles of protein can be calculated from the relationship $n = M \times V$.

$$n = (2.59 \cdot 10^{-4} \text{ M})(5.00 \cdot 10^{-3} \text{ L}) = 1.29 \cdot 10^{-6} \text{ mol}$$

$$M_m = \frac{15.8 \cdot 10^{-3} \text{ g protein}}{1.29 \cdot 10^{-6} \text{ mol protein}} = 12,200 \text{ g/mol}$$

55. Vitamin K is involved in normal blood clotting. When 1.00 g of vitamin K is dissolved in 20.0 g of camphor, ($k_f = 40.0 \text{ }^\circ\text{C}/\text{m}$) the freezing point of the solution is lowered by 4.43 °C. What is the molar mass of vitamin K?

Use Eq 2.12 to find the molality of vitamin K (vitamin K does not dissociate in solution, so $i = 1$)

$$\Delta T_f = i \times k_f \times m = 1 \times 40.0 \text{ }^\circ\text{C}/\text{m} \times m = 4.43 \text{ }^\circ\text{C}$$

$$\text{Solving for } m \text{ yields } m = \frac{\Delta T_f}{k} = \frac{4.43 \text{ }^\circ\text{C}}{40.0 \text{ }^\circ\text{C}\cdot\text{m}^{-1}} = 0.111 \text{ m}$$

Determine the number of moles of pepsin dissolved in the 20.0 g (0.0200 kg) of solvent

$$0.0200 \text{ kg camphor} \times \frac{0.111 \text{ mol vitamin K}}{\text{kg camphor}} = 0.00222 \text{ mol vitamin K}$$

The molar mass is obtained as the ratio of the mass of vitamin K to the number of moles that the mass represents.

$$M_m = \frac{\text{mass of vitamin K}}{\text{moles of vitamin K}} = \frac{1.00 \text{ g vitamin K}}{0.00222 \text{ mol vitamin K}} = 451 \text{ g/mol}$$

57. What is the concentration of the excess reactant in a solution prepared by mixing 25.0 mL of 0.242 M HCl with 36.3 mL of 0.167 M $\text{Ba}(\text{OH})_2$? Is this solution acidic, basic or neutral?

$$25.0 \text{ mL} \times \frac{0.242 \text{ mmol H}_3\text{O}^{1+}}{\text{mL}} = 6.05 \text{ mmol H}_3\text{O}^{1+} \quad \& \quad 36.3 \text{ mL} \times \frac{0.167 \text{ mmol Ba}(\text{OH})_2}{1 \text{ mL}} \times \frac{2 \text{ mmol OH}^{1-}}{1 \text{ mmol Ba}(\text{OH})_2} = 12.1 \text{ mmol OH}^{1-}$$

The reaction table:

	H_3O^{1+}	$+$	OH^{1-}	\rightarrow	$2\text{H}_2\text{O}$
initial	6.05		12.1		mmol
Δ	-6.05		-6.05		mmol
final	~		6.05		mmol

Note that the 5 is written as a subscript in the final mmol of OH^{1-} to indicate that it is not significant but that it will be used in future calculations. Therefore, the solution is basic because there is an excess amount of $\text{Ba}(\text{OH})_2$. The

$$\text{concentration of hydroxide in the final solution can be calculated as follows: } [\text{OH}^{1-}] = \frac{6.05 \text{ mmol OH}^{1-}}{(36.3 + 25.0) \text{ mL}} = 0.099 \text{ M}$$

59. Determine the amount of precipitate that will form if 50.0 mL of 0.100 M Na₂SO₄(aq) and 100. mL of 0.300 M BaCl₂ are mixed.

The reaction is Ba²⁺(aq) + SO₄²⁻(aq) → BaSO₄(s). (50.0 mL)(0.100 mmol/mL) = 5.00 mmol SO₄²⁻ and (100. mL)(0.300 mmol/mL) = 30.0 mmol Ba²⁺ are mixed, so SO₄²⁻ ion is the limiting reagent.

$$5.00 \text{ mmol SO}_4^{2-} \times \frac{1 \text{ mmol BaSO}_4}{1 \text{ mmol SO}_4^{2-}} \times \frac{233.4 \text{ mg BaSO}_4}{1 \text{ mmol BaSO}_4} = 1.17 \times 10^3 \text{ mg BaSO}_4 \text{ so } 1.17 \text{ g of BaSO}_4 \text{ precipitate.}$$

61. Construct the reaction table for the *net* reaction of 36.6 mL of 0.0668 M FeCl₃ and 24.2 mL of 0.100 K₂S.

The solubility rules indicate that KCl is soluble, but that Fe₂S₃ should be insoluble. Thus, the reaction is between Fe³⁺ and S²⁻. Next, determine the number of millimoles of each reactant

$$36.6 \text{ mL soln} \times \frac{0.0668 \text{ mmol FeCl}_3}{\text{mL soln}} \times \frac{1 \text{ mmol Fe}^{3+}}{1 \text{ mmol FeCl}_3} = 2.44 \text{ mmol Fe}^{3+}$$

$$24.2 \text{ mL soln} \times \frac{0.100 \text{ mmol K}_2\text{S}}{\text{mL soln}} \times \frac{1 \text{ mmol S}^{2-}}{1 \text{ mmol K}_2\text{S}} = 2.42 \text{ mmol S}^{2-}$$

Sulfide is the limiting reactant as it can produce only $\frac{1}{3}(2.42) = 0.807 \text{ mmol Fe}_2\text{S}_3$ while Fe³⁺ can produce $\frac{1}{2}(2.44) = 1.22 \text{ mmol Fe}_2\text{S}_3$.

The reaction table:

	2Fe ³⁺	+ 3S ²⁻	→	Fe ₂ S ₃	
initial	2.44	2.42		0	mmol
Δ	-1.62	-2.42		+0.81	mmol
final	0.82	0		0.81	mmol

a) What mass of Fe₂S₃ precipitates? $0.807 \text{ mmol Fe}_2\text{S}_3 \times \frac{207.9 \text{ mg Fe}_2\text{S}_3}{\text{mmol Fe}_2\text{S}_3} = 168 \text{ mg Fe}_2\text{S}_3$

- b) What is the concentration of the excess reactant after complete reaction? Assume additive volumes.

Total volume of the reaction solution is 60.8 mL, so $[\text{Fe}^{3+}] = \frac{0.82 \text{ mmol Fe}^{3+}}{60.8 \text{ mL solution}} = 0.014 \text{ M}$

63. Given the unbalanced chemical equation: $\underline{\hspace{1cm}} \text{Zn(s)} + \underline{\hspace{1cm}} \text{HCl(aq)} \rightarrow \underline{\hspace{1cm}} \text{H}_2(\text{g}) + \underline{\hspace{1cm}} \text{ZnCl}_2(\text{aq})$

What is the molarity of an HCl solution if 130. mL of the solution reacts with an excess of Zn to produce 7.00 L of H₂(g) at 273 K and 1.00 atm? $\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow 2\text{H}_2(\text{g}) + \text{ZnCl}_2(\text{aq})$

Use the ideal gas law to determine the number of moles of H₂(g) and the balanced equation to get mol HCl.

$$n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(7.00 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(273 \text{ K})} \text{ mol H}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol H}_2} = 0.624 \text{ mol HCl reacted}$$

$$[\text{HCl}] = \frac{0.624 \text{ mol HCl}}{0.130 \text{ L soln}} = 4.81 \text{ M HCl}$$

65. A solution is prepared by dissolving 5.863 g of impure NaOH in sufficient water to make 1.000 L of solution. The solution is added to a buret and used to titrate 25.00 mL of a 0.1173 M solution of HCl.

The balanced chemical equation for the titration is H¹⁺ + OH¹⁻ → H₂O.

- a) What is the molarity of the NaOH solution if the titration requires 28.04 mL of base?

$$\text{mmol NaOH required: } 25.00 \text{ mL} \times \frac{0.1173 \text{ mmol H}^{1+}}{1 \text{ mL}} \times \frac{1 \text{ mmol OH}^{1-}}{1 \text{ mmol H}^{1+}} \times \frac{1 \text{ mmol NaOH}}{1 \text{ mmol OH}^{1-}} = 2.932 \text{ mmol NaOH}$$

$$\text{molarity of the NaOH solution. } \frac{2.932 \text{ mmol NaOH}}{28.04 \text{ mL}} = 0.1046 \text{ M NaOH}$$

Solutions

b) What is the percent purity of the original NaOH sample?

Determine the amount of analytically pure NaOH in the base solution in milligrams/mL. The molar mass of NaOH is $39.995 \text{ g mol}^{-1}$.

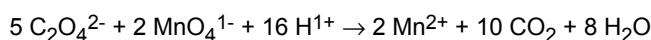
$$2.932 \text{ mmol NaOH} \times \frac{39.995 \text{ mg NaOH}}{1 \text{ mmol}} = 117.3 \text{ mg NaOH in } 28.04 \text{ mL}$$

$$\text{Therefore there were } 117.3 \text{ mg NaOH} \times \frac{1000 \text{ mL}}{28.04 \text{ mL}} = 4183 \text{ mg NaOH in } 1 \text{ L of solution}$$

$$\text{Percent purity} = \frac{\text{mass pure NaOH}}{\text{mass of sample}} \times 100\% = \frac{4183 \text{ mg}}{5863 \text{ mg}} \times 100\% = 71.35\%$$

67. A 10.00-mL blood sample was diluted to 100.00 mL. 10.00 mL of the resulting solution was analyzed for calcium by precipitating all of the calcium as calcium oxalate, CaC_2O_4 . The solid CaC_2O_4 was then redissolved in H_2SO_4 and titrated with 0.00886 M KMnO_4 . What is the concentration of Ca^{2+} in the blood sample expressed as milligrams Ca^{2+}/mL blood if the endpoint was reached with the addition of 1.14 mL of the KMnO_4 solution?

The balanced chemical reaction for the titration is



Use the factor-label method, determine the moles of Ca^{2+} ion required to reach the equivalence point.

$$1.14 \text{ mL} \times \frac{0.00886 \text{ mmol MnO}_4^{1-}}{1 \text{ mL}} \times \frac{5 \text{ mmol C}_2\text{O}_4^{2-}}{2 \text{ mmol MnO}_4^{1-}} \times \frac{1 \text{ mmol CaC}_2\text{O}_4}{1 \text{ mmol C}_2\text{O}_4^{2-}} \times \frac{1 \text{ mmol Ca}^{2+}}{1 \text{ mmol CaC}_2\text{O}_4} = 0.0253 \text{ mmol Ca}^{2+}$$

$$0.0253 \text{ mmol Ca}^{2+} \times \frac{40.08 \text{ mg Ca}^{2+}}{1 \text{ mmol Ca}^{2+}} = 1.01 \text{ mg Ca}^{2+}. \text{ However, this is the calcium in only } 1/10 \text{ of the}$$

sample. Therefore, there are $10.00 \times 1.01 \text{ mg Ca}^{2+} = 10.1 \text{ mg Ca}^{2+}$ in 10.00 mL of blood, so the calcium

$$\text{concentration is } \frac{10.1 \text{ mg Ca}^{2+}}{10.00 \text{ mL blood}} = 1.01 \text{ mg mL}^{-1} \text{ Ca}^{2+}$$

69. All of the iron in 1.314 g of an ore sample was converted to an aqueous solution of Fe^{2+} , which was analyzed by titration with dichromate. What is the mass percent of iron in the ore if 38.64 mL of $0.02063 \text{ M K}_2\text{Cr}_2\text{O}_7$ was required to reach the equivalence point?

The balanced chemical equation for the titration is $14 \text{ H}^+ + \text{Cr}_2\text{O}_7^{2-} + 6 \text{ Fe}^{2+} \rightarrow 6 \text{ Fe}^{3+} + 2 \text{ Cr}^{3+} + 7 \text{ H}_2\text{O}$

Using the factor-label method, determine the mass of Fe^{2+} present in the iron ore sample.

$$38.64 \text{ mL} \times \frac{0.02063 \text{ mmol Cr}_2\text{O}_7^{2-}}{1 \text{ mL}} \times \frac{6 \text{ mmol Fe}^{2+}}{1 \text{ mmol Cr}_2\text{O}_7^{2-}} \times \frac{55.847 \text{ mg Fe}^{2+}}{1 \text{ mmol Fe}^{2+}} = 267.1 \text{ mg Fe}^{2+}$$

$$\text{Calculate the mass percent of iron in the iron ore sample. } \frac{0.2671 \text{ g Fe}^{2+}}{1.314 \text{ g iron ore}} \times 100\% = 20.33\%$$

71. The concentration of a stock solution of Rhodamine B, a commercial dye ($\epsilon = 1.060 \times 10^5 \text{ M}^{-1} \cdot \text{cm}^{-1}$ at 543 nm) is determined spectrophotometrically, but the absorbance of the stock solution was too great to read on a spectrometer, so 15.00 mL of the solution was diluted to 250.00 mL . The resulting solution was still too concentrated, so 10.00 mL of that solution was diluted to 500.00 mL , but the solution was still too concentrated. However, a dilution of 25.00 mL of the resulting solution to a final volume of 500.00 mL provided an acceptable solution with an absorbance of 0.1814 in a 1.000 cm cell. What is the concentration of the stock solution?

$$\text{Concentration of the final solution: } c = \frac{A}{\epsilon l} = \frac{0.1814}{(1.060 \times 10^5 \text{ M}^{-1} \cdot \text{cm}^{-1})(1.000 \text{ cm})} = 1.711 \times 10^{-6} \text{ M}$$

$$\text{Dilutions: } 1.711 \times 10^{-6} \text{ M} = c_{\text{stock}} \left(\frac{15.00}{250.00} \right) \left(\frac{10.00}{500.00} \right) \left(\frac{25.00}{500.00} \right) = c_{\text{stock}} (6.000 \times 10^{-5})$$

$$\text{Concentration of stock: } c_{\text{stock}} = \frac{1.711 \times 10^{-6} \text{ M}}{6.000 \times 10^{-5}} = 0.02852 \text{ M}$$

What volume of stock solution would be required to make 2.000 L of a solution A = 1.000 in a 1.000 cm cell?

Use Beer's law to determine the concentration of the final solution,

$$c = \frac{A}{\epsilon l} = \frac{1.000}{(1.060 \times 10^5 \text{ M}^{-1} \cdot \text{cm}^{-1})(1.000 \text{ cm})} = 9.434 \times 10^{-6} \text{ M}$$

Then use the dilution to determine the initial volume.

$$V_i = \frac{C_f V_f}{C_i} = \frac{(9.434 \times 10^{-6} \text{ M})(2000 \text{ mL})}{0.02852 \text{ M}} = 0.6617 \text{ mL} = 661.7 \mu\text{L}$$

73. How many grams of NaCl must be added to 120 mL of water at 25 °C to prepare a solution with a vapor pressure of 22.0 torr?

The vapor pressure of pure water at 25 °C from Table 2.1 is 23.8 torr. Use that vapor pressure and the given vapor pressure to determine the required mole fraction of water.

$$X_{\text{water}} = \frac{P}{P^\circ} = \frac{22.0}{23.8} = 0.924. \text{ The mole fraction of particles is } 1.000 - 0.924 = 0.076.$$

$$\text{Determine the number of moles of water in 126 mL: } 120 \text{ mL} \times \frac{1.0 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{18.0 \text{ g}} = 6.7 \text{ mol}$$

Let y = moles of NaCl required. The van't Hoff factor for NaCl is 2, so $2y$ moles of particles are in solution. Use the mole fraction and moles of water above to determine y .

$$X = 0.076 = \frac{\text{moles particles}}{\text{moles water} + \text{moles particles}} = \frac{2y}{6.7 + 2y}$$

$$0.076(6.7 + 2y) = 2y$$

$$0.51 = (2 - 0.15)y$$

$$y = \frac{0.51}{1.85} = 0.28 \text{ mol NaCl}$$

$$(0.28 \text{ mol})(58.5 \text{ g/mol}) = 16 \text{ g NaCl}$$

