

11.E: Solutions (Exercises)

These are homework exercises to accompany the Textmap created for "Principles of Modern Chemistry" by Oxtoby et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed [here](#).

Q.1

A student goes to the hospital and has blood work done and the results state that he has a bromide count of $3.17 \frac{mg}{dL}$.

- What is the molar concentration of the students blood?
- What is the molality of the student's blood if the density of his blood was found to be $1.06 \frac{g}{mL}$?
- How much of the students blood would be require to get 3g of bromide?

Solution

The molar mass of bromide is $79.90 \frac{g}{mol}$ $1L = 10dL$, $1g = 1000mg$

$$\left(\frac{3.17 \cancel{mg}}{1 \cancel{dL}} \right) \left(\frac{1 \cancel{g}}{1000 \cancel{mg}} \right) \left(\frac{10 \cancel{dL}}{1L} \right) \left(\frac{1mol}{79.90 \cancel{g}} \right) = 3.97 \times 10^{-4} \frac{mol}{L}$$

Assume there is 1L of solution. Modality (m) is calculated by:

$$3.17 \times 10^{-2} \cancel{g} \left(\frac{1mol}{79.90 \cancel{g}} \right) \left(\frac{1}{1.06 kg} \right) = 3.74 \times 10^{-4} \frac{mol}{kg}$$

If we know that there is $3.17 \times 10^{-2} g$ of bromide in 1L of blood, so all that needs to be done is divide 3g by $3.17 \times 10^{-2} \frac{g}{L}$:
Volume of Blood = 94.64 L

Q.5

Given a 3.1416 M aqueous solution of sucrose, $C_{12}H_{22}O_{11}$ with a density of $1.5986 \frac{g}{mL}$. Calculate the molality of this solution, remember that the molar mass of sucrose is $342.297 \frac{g}{mol}$

Solution

Before we calculate the molality of this solution, don't forget that molarity is just mol/L so we will multiply the molarity of the solution by 1 L to isolate the moles.

$$1.00 \cancel{L} \text{ Water} = \frac{3.1416 mol \text{ Sucrose}}{1.00 \cancel{L} \text{ Water}} = 3.1416 mol \text{ Sucrose}$$

Once we have the moles of Sucrose we will multiply it by the molar mass of sucrose in order to find the number of sucrose grams present in the solution:

$$3.1416 \cancel{mol} \text{ Sucrose} = \frac{342.297 g \text{ Sucrose}}{1.00 \cancel{mol} \text{ Sucrose}} = 1075.4 g \text{ Sucrose}$$

We next will multiply the density by 1000 mL so as get the number of grams per liter, this way we can find the total grams of the solution and can subtract the sucrose grams from the solution grams to find the grams of water.

$$1000 \cancel{mL} = \frac{1.5986 g \text{ SucroseSoln}}{1.00 \cancel{mL} \text{ SucroseSoln}} = 1598.6 g/L \text{ SucroseSoln}$$

$$\text{Water grams} = [1598.6g \text{ Sucrose Soln.}] - [1075.4g \text{ Sucrose}] = 523.2g \text{ Water} = 0.5232 kg \text{ Water}$$

Lastly we divide the moles of Sucrose by the kg of water in order to get molality.

$$\frac{3.1416 mol \text{ Sucrose}}{0.5232 kg \text{ Water}} = 6.005 molal$$

Q.9

A solution contains 75% of ethanol (C_2H_6O) by mass and the rest is water.

- What is the density of the solution if there is 15 mol of ethanol per liter of solution?
- To prepare a 3.5L of 2M ethanol, how many milliliters of the solution is needed?

Solution

a) For easy calculation, assume that there is 1 L of the solution. Using:

$$\text{mol} = \text{molarity} \times \text{volume} \quad (11.E.1)$$

The mol of 15M ethanol in 1L is 15 mol. The molar mass of ethanol is $46.068 \frac{g}{mol}$. Using the equation below:

$$\text{ethanol mol} = \frac{\text{mass ethanol}}{\text{ethanol molar mass}} \quad (11.E.2)$$

$$\text{ethanol mass in the solution} = \text{ethanol mol} \times \text{molar mass} = (15 \text{ mol}) \left(46.068 \frac{g}{mol} \right) = 691.02g \quad (11.E.3)$$

Ethanol is 75% of the solution by mass, therefore:

$$\text{mass of solution} \times 0.75 = \text{mass of ethanol} \quad (11.E.4)$$

$$\text{mass of solution} \times 0.75 = 691.02g \quad (11.E.5)$$

$$\text{mass of solution} = \frac{691.02g}{0.75} = 921.36g \quad (11.E.6)$$

The equation to calculate density is:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad (11.E.7)$$

$$\text{Density} = \frac{921.36g}{1L} = 0.921 \frac{g}{ml} \quad (11.E.8)$$

b) To prepare a 3.5L of 2M ethanol;

$$\text{ethanol mol needed} = \text{molarity} \times \text{volume} = (2 \text{ M})(3.5 \text{ L}) = 7 \text{ mol} \quad (11.E.9)$$

$$V_{\text{ethanol needed}} = \frac{\text{mol needed}}{\text{molarity}} = \frac{7 \text{ mol}}{15 \text{ mol}} \times 1 \text{ L} = 0.467 \text{ L} \quad (11.E.10)$$

0.467L of the solution is needed, which is 467mL.

Q.14

Rewrite the following balanced equations as net ionic equations.

- $H_2SO_4(aq) + 2 NaOH(aq) \rightarrow 2 H_2O(l) + Na_2SO_4(aq)$
- $2 Na_3PO_4(aq) + 3 CaCl_2(aq) \rightarrow 6 NaCl(aq) + Ca_3(PO_4)_2(s)$
- $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2 NaCl(aq)$
- $2 AgNO_3(aq) + ZnCl_2(aq) \rightarrow 2 AgCl(s) + Zn(NO_3)_2(aq)$

Solution

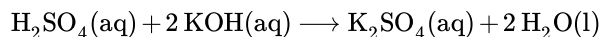
- $2 H^+(aq) + 2 OH^-(aq) \rightarrow 2 H_2O(l)$
- $2 PO_4^{3-}(aq) + 3 Ca^{2+}(aq) \rightarrow Ca_3(PO_4)_2(s)$
- $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
- $2 Ag^+(aq) + 2 Cl^-(aq) \rightarrow 2 AgCl(s)$

Q.15

Write the balanced chemical equation for the titration of H_2SO_4 with KOH , then find the volume of a 5.49 M KOH solution needed to neutralize 36.2 g of H_2SO_4 .

Solution

The balanced equation of H_2SO_4 with KOH is:



Calculate the amount of moles of H_2SO_4 being neutralized:

$$n_{\text{H}_2\text{SO}_4} = 36.2 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.079 \text{ g H}_2\text{SO}_4} = 0.369 \text{ mol H}_2\text{SO}_4$$

Using stoichiometry to find the amount of moles of KOH needed to neutralize the amount of H_2SO_4 present:

$$n_{\text{KOH}} = 0.369 \text{ mol H}_2\text{SO}_4 \times \frac{2 \text{ mol KOH}}{1 \text{ mol H}_2\text{SO}_4} = 0.738 \text{ mol KOH}$$

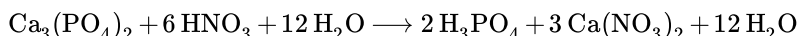
Calculate the volume of the 5.49 M KOH solution needed to neutralize the amount of H_2SO_4 present:

$$V_{\text{KOH}} = 0.738 \text{ mol KOH} \times \frac{1 \text{ L}}{5.49 \text{ mol KOH}} = 0.134 \text{ L} = 134 \text{ mL}$$

Therefore, approximately 134 mL of the 5.49 M KOH solution is needed to neutralize 36.2 g H_2SO_4 .

Q.16

Phosphoric acid is made industrially as a by-product of the extraction of calcium phosphate



What volume of 8.5 M phosphoric acid is generated by the reaction of 4.5 metric tons (4500 kg) of calcium phosphate?

Solution

$$4500 \text{ kg Ca}_3(\text{PO}_4)_2 \times \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \times \left(\frac{1 \text{ mol}}{310 \text{ g}} \right) \times \left(\frac{2 \text{ H}_3\text{PO}_4}{1 \text{ Ca}_3(\text{PO}_4)_2} \right) \times \left(\frac{1 \text{ L}}{8.5 \text{ moles}} \right) = 3415.56 \text{ L H}_3\text{PO}_4$$

Q.19

Write a balanced equation for the acid-base reaction that leads to the production of each of the following salts.

- BaSO_4
- MgSO_3
- PbSO_4
- AgCl

Solution

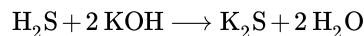
- $\text{Ba}(\text{OH})_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightleftharpoons \text{BaSO}_4(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$
- $\text{Mg}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_3(\text{aq}) \rightleftharpoons \text{MgSO}_3(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$
- $\text{Pb}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightleftharpoons \text{PbSO}_4(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$
- $\text{AgOH}(\text{s}) + \text{HCl}(\text{aq}) \rightleftharpoons \text{AgCl}(\text{s}) + \text{H}_2\text{O}(\text{l})$

Q.21

Hydrogen selenide is classified as a binary acid that has similar properties as hydrogen sulfide. Write down the balanced equation of hydrogen selenide reacting with the base potassium hydroxide and name the salt product of the reaction.

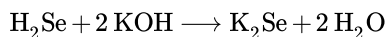
Solution

When the acid H_2S reacts with the base KOH , the reaction goes as follows:



Forming potassium sulfide.

Similarly, since H_2Se has similar properties as H_2S , a similar reaction occurs:



Potassium selenide is the salt product of the reaction.

Q.25

A student working with an unknown solution of hydrochloric acid is assigned the task of finding its concentration. The student places a 200.0 mL sample of the solution in a 350.0 mL Erlenmeyer flask and titrates the solution with a 0.1234 M solution of sodium hydroxide, where a phenolphthalein indicator is used to find the endpoint. The endpoint is found to be 31.416 mL, what is the concentration of the original hydrochloric acid solution sample?

Solution

Remember that in titration of a strong acid and base, when the endpoint is reached the number of moles used is equal, so the number of moles of NaOH at the endpoint are equal to that of HCl . Use this information to find moles of NaOH from the concentration and volume used to titrate.

$$\text{Concentration of X} = \frac{\text{moles of X}}{\text{liters of X}}$$

Use dimensional analysis to convert concentration of NaOH to moles using the known amount of volume titrated, in liters:

$$0.031416 \cancel{\text{ liter}} \text{ NaOH} = \frac{0.1234 \text{ mol NaOH}}{1.00 \cancel{\text{ liter}} \text{ NaOH}} = 0.003877 \text{ mol}$$

Now because we know that the moles of NaOH are equal to the moles of HCl we simply plug in our calculated moles and known original sample volume, in liters, to find the concentration of HCl :

$$\text{Concentration of HCl} = \frac{0.003877 \text{ mol HCl}}{0.200 \text{ liters HCl}} = 0.019385 \text{ M HCl}$$

The concentration of the original hydrochloric acid solution sample was, therefore, 0.019385 M.

Q.41

The vapor pressure of salicylic acid $\text{C}_7\text{H}_6\text{O}_3$ at 200°C is 0.1598 atm. A 19.0 g sample of ethanol $\text{C}_2\text{H}_6\text{O}$ is dissolved in 60.0 g of salicylic acid. Calculate the vapor pressure of salicylic acid above the resulting solution.

Solution

First find the molar masses of salicylic acid and ethanol. Then use them to compute the chemical amounts of the two in solution.

$$n_{\text{C}_2\text{H}_6\text{O}} = 19.0 \text{ g C}_2\text{H}_6\text{O} \times \frac{1 \text{ mol C}_2\text{H}_6\text{O}}{46.068 \text{ g C}_2\text{H}_6\text{O}} = 0.4124 \text{ mol C}_2\text{H}_6\text{O}$$

$$n_{\text{C}_7\text{H}_6\text{O}_3} = 60.0 \text{ g C}_7\text{H}_6\text{O}_3 \times \frac{1 \text{ mol C}_7\text{H}_6\text{O}_3}{138.121 \text{ g C}_7\text{H}_6\text{O}_3} = 0.4344 \text{ mol C}_7\text{H}_6\text{O}_3$$

The mole fraction of ethanol in the solution equals

$$X_{\text{C}_2\text{H}_6\text{O}} = \frac{0.4124}{0.4344 + 0.4124} = 0.4870$$

The change in vapor pressure of the salicylic acid due to the presence of the ethanol is

$$\Delta P_{\text{C}_4\text{H}_6\text{O}_3} = -X_{\text{C}_4\text{H}_6\text{O}_3} \times P_{\text{C}_4\text{H}_6\text{O}_3}^\circ = -0.5130 \times 0.1518 \text{ atm} = -0.08198$$

The final vapor pressure of the salicylic acid equals its original vapor pressure P° plus the change. This is 0.1598 atm minus 0.08198 atm, which is 0.07782 atm.

Q.43

When 53 mol of an unknown compound was added to 100 g of water, the normal boiling point of water increased to 101.1°C. What is the boiling point elevation constant of water? The compound does not dissociate in solution.

Solution

The relevant relationship here is the colligative property associated with boiling point elevation

$$\Delta T_b = i K_b m \quad (11.E.11)$$

$$\Delta T_b = 101.1^\circ \text{C} - 100^\circ \text{C} = 1.1^\circ \text{C} \quad (11.E.12)$$

Since the compound was told to not dissociate, we know that the van't Hoff factors is 1:

$$i = 1 \quad (11.E.13)$$

$$m = \frac{\text{mol solute}}{\text{kg solvent}} = \frac{53 \text{ mol}}{0.100 \text{ kg}} = 0.53 \frac{\text{mol}}{\text{g}} \quad (11.E.14)$$

$$1.1^\circ \text{C} = (1) \times K_b \times (0.53 \frac{\text{mol}}{\text{g}}) \quad (11.E.15)$$

$$K_b = \frac{1.1^\circ \text{C}}{0.53 \frac{\text{mol}}{\text{kg}}} = 2.075 \frac{^\circ \text{C kg}}{\text{mol}} \quad (11.E.16)$$

This problem does not give the correct constant though 0.513:

[https://chem.libretexts.org/Reference...erties/B5%3A_Ebullioscopic_\(Boiling_Point_Elevation\)_Constants](https://chem.libretexts.org/Reference%20Data/Ebullioscopic_(Boiling_Point_Elevation)_Constants)

Q.45

When 5.82 g of uniformly dissociating unknown salt, YX, is dissolved in 100.0 g of water, the boiling point of the water is raised by 0.20°C. When 7.09 g of uniformly dissociating unknown salt, ZX, is dissolved in 200.0 g of water, the boiling point of the water is raised by 0.30°C. Assuming both have a van't hoff factor of 2, identify the elements that X, Y, and Z represent.

Solution

For both salts, one can utilize the boiling point elevation expression (including the van't hoff factor):

$$\Delta T = i \times K_b \times m$$

and by substituting moles solute over mass solvent into the molality, m , one can derive that:

$$n_{\text{salt}} = \frac{\Delta T \times m_{\text{water}}}{i \times K_b}$$

From this we get the moles of each salt present, and can therefore get their molar mass by dividing their mass by moles present. Next, subtracting the heavier salt by the lighter salt will **yield** the molar mass of X, $\sim 35.5 \frac{\text{g}}{\text{mol}}$. Subtracting this from the molar masses of each salt will then **yield** Y and Z respectively.

$$\Delta T = i \times m \times K_b = 2 \times m \times 0.512 = 0.2, \text{ m of YX} = 0.1953 \frac{\text{mol}}{\text{kg}}; m \times 0.1 \text{ kg} = 0.0195 \text{ mol}; \frac{5.82}{0.0195} = 298.46 \frac{\text{g}}{\text{mol}}$$

$$\Delta T = i \times m \times K_b = 2 \times m \times 0.512 = 0.3, \text{ m of ZX} = 0.2930 \frac{\text{mol}}{\text{kg}}; m \times 0.2 \text{ kg} = 0.0586 \text{ mol}; \frac{7.09}{0.0586} = 120.99 \frac{\text{g}}{\text{mol}}$$

$$\text{molar mass difference between Y and Z} = 298.46 - 120.99 = 177.47 \frac{\text{g}}{\text{mol}}$$

Q.47

Camphor is a useful agent for determining solute's molar mass. Camphor's melting point and K_f are given at 451.55 K and $37.7 \frac{\text{K kg}}{\text{mol}}$. Find the freezing point of the solution of 0.72 g of glucose is dissolved in the 30.0 g of camphor.

Solution

Since we are evaluating glucose as the solute, $i = 1$.

$$\Delta T = -iK_fm \quad (11.E.17)$$

First find number of moles of glucose in the solution using molar mass.

$$\text{Moles}_{\text{glucose}} = \frac{(0.72\text{g})}{(180.1559 \frac{\text{g}}{\text{mol}})} = 0.0040\text{mol} \quad (11.E.18)$$

Molality then can be determined.

$$m = \frac{(0.0040\text{mol})}{(0.030\text{kg})} = 0.13 \frac{\text{mol}}{\text{kg}} \quad (11.E.19)$$

Freezing then can be easily found as such:

$$\Delta T = -iK_fm = -(1)(37.7 \frac{\text{K kg}}{\text{mol}})(0.13 \frac{\text{mol}}{\text{kg}}) = -4.9^\circ\text{C} \quad (11.E.20)$$

The freezing is $451.55\text{K} - 4.9\text{K} = 446.65\text{K} = 173.5^\circ\text{C}$.

Q.48

It is observed that 2 grams of a non-electrolyte solute dissolved in 100 grams of benzene lowered the freezing point of the solution by 0.40°C compared to pure benzene. Given that the freezing point depression constant of benzene is 5.12 K kg/mol , find the molar mass of the solute. If it is given that the molecular formula of the solute is X_8 where X is an element, what is the formula of the solute?

Solution

$$\Delta T = -iK_fm$$

It is given that

$$\Delta T = 0.40^\circ\text{C} \quad (11.E.21)$$

Since the solution is a non-electrolyte, $i=1$

Hence,

$$\begin{aligned} \frac{0.40}{K_f} &= m \\ m &= \frac{\frac{2}{\text{Molecular Mass}}}{100 \times 10^{-3}} = \frac{\frac{2}{\text{Molecular Mass}}}{0.1} \end{aligned}$$

Hence, we get

$$\begin{aligned} \frac{0.40 \times 0.1}{K_f} &= \frac{2}{\text{Molecular Mass}} \\ \text{Molecular Mass} &= \frac{2K_f}{0.40 \times 0.1} \end{aligned}$$

Hence, the Molecular Mass of the solute is $256 \frac{\text{g}}{\text{mol}}$

Since the Molecular Mass of the solute is $256 \frac{\text{g}}{\text{mol}}$, we can say that,

$$\text{Molecular Mass of X} = \frac{256}{8} = 32$$

The solute is S_8 .

Q.49

At what temperature does the first ice crystals begin to form in a 27% salt (by mass) aqueous solution of NaCl? As the crystallization of water carries on, the remaining solution becomes more concentrated, so what happens to the freezing point of the solution? Given that the freezing point depression constant (K_f) of water is $1.86 \frac{^\circ\text{C} \cdot \text{kg}}{\text{mol}}$.

Solution

$$27\% \text{ NaCl(aq)} = \frac{27 \text{ g NaCl}}{73 \text{ g H}_2\text{O}} \quad (11.E.22)$$

$$\text{molality} = \frac{27 \text{ g NaCl}}{73 \text{ g H}_2\text{O}} \times \frac{\text{mol NaCl}}{58.44 \text{ g}} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 6.33 \text{ m} \quad (11.E.23)$$

$$\Delta T = -i \times m \times K_f = -2 \times 6.33 \frac{\text{mol}}{\text{kg}} \times 1.86 \frac{^\circ\text{C} \cdot \text{kg}}{\text{mol}} = -23.5^\circ\text{C} \quad (11.E.24)$$

The first ice crystal begins to appear at -23.5°C , as the solution becomes more concentrated, its freezing point decreases further because freezing point depression is a colligative property, so the freezing point of a solution is lower than that of the pure solvent and is directly proportional to the molality of the solute.

Q.50

When homemade ice cream is being made, the temperature ranging downward from -3°C are needed. Ice cubes from a freezer have a temperature of about -12°C , which is cold enough; however, when the ice cream mixture is mixed with the ice cubes, the liquid balances out to 0°C , which is too warm. To obtain a liquid that is cold enough, salt NaCl is dissolved in water and ice is added to the saltwater. The salt lowers the freezing point of the water enough so that it can freeze the liquid inside the ice cream maker. The instruction for the ice cream maker say to add one part salt to 11 parts water (by mass). What is the freezing point of this solution in $^\circ\text{C}$? Assume the NaCl dissociates fully into ions, and that the solution is ideal.

Solution

$$\Delta T_f = -ik_fm \quad (11.E.25)$$

$$\frac{1.00 \text{ g NaCl}}{53.5 \frac{\text{g}}{\text{mol}}} = 0.01869 \text{ mol NaCl} \quad (11.E.26)$$

$$m = \frac{0.01869 \text{ mol}}{11.00 \text{ g} \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right)} = 1.699 \text{ m} \quad (11.E.27)$$

$$x - (-12^\circ\text{C}) = -(2) \times \left(1.86 \frac{^\circ\text{C}}{\text{m}} \right) \times (1.699 \text{ m}) \quad (11.E.28)$$

$$x = -18.32^\circ\text{C} \quad (11.E.29)$$

Q.53

A sample of a purified unknown compound is dissolved in toluene, diluting the solvent to a volume of 1.05 mL. The resulting solution has an osmotic pressure of 0.025 atm at 275 K. If the solute has a molar mass of 46.06 g/mol, how many grams of the unknown compound were added?

Solution

$$\Pi = \frac{nRT}{V}$$

and we know $T = 275\text{ K}$, $R = 0.0821 \frac{\text{L atm}}{\text{mol K}}$ and $\Pi = 0.025 \text{ atm}$

Thus $V = 1.05 \text{ mL}$

Since we need V in liters because the units of R are $\frac{\text{L atm}}{\text{mol K}}$, we use unit conversions:

$$1.05 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00105 \text{ L}$$

Plug in the values we know to find n

$$\begin{aligned} 0.025 &= \frac{n \times 0.0821 \times 275}{0.00105} \\ \frac{0.025 \times 0.00105}{0.0821 \times 275} &= n \\ n &= 1.163 \times 10^{-6} \end{aligned}$$

Now that we know number of moles we just have to multiply the number of moles by the molar mass to find grams of substance.

$$(1.163 \times 10^{-6}) \times 46.06844 \frac{\text{g}}{\text{mol}} = 5.356 \times 10^{-5} \text{ g}$$

Q.57

An amount of hydrogen gas (H_2) with a partial pressure of 4.6 atm is dissolved in water and this solution is sealed. The Henry's Law constant K for this solution at 25°C is $7.8 \times 10^{-4} \frac{\text{M}}{\text{atm}}$.

- How many moles of hydrogen gas will dissolve per liter of water? Assume the density of water is $1.0 \frac{\text{grams}}{\text{cm}^3}$ and has not been affected by the temperature.
- What will the dissolved gas molecules do if the sealed solution is unsealed?

Solution

- According to [Henry's Law](#),

$$C = k \times P_{\text{gas}}$$

In this equation, C = concentration of gas (M), k = Henry's Law constant ($\frac{\text{M}}{\text{atm}}$), and P_{gas} = partial pressure of gas.

So, given k and P_{gas} :

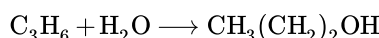
$$C = 7.8 \times 10^{-4} \frac{\text{M}}{\text{atm}} \times 4.6 \text{ atm}$$

$$C = 0.003588 \frac{\text{mol H}_2}{\text{L}_{\text{water}}}$$

- The partial pressure of H_2 in the Earth's atmosphere is less than 1 atm . When the solution is unsealed, the H_2 molecules will no longer be under enough pressure to dissolve in water, so they will quickly bubble out of the solution and escape as gas into the surroundings.

Q.59

Propene is a colorless organic compound with a weak but unpleasant smell. However, a company can take advantage of its chemical properties to produce propanol:



- The heat of formation of propene is $20.4 \frac{\text{kJ}}{\text{mol}}$.
- The heat of formation of H_2O is $-241.8 \frac{\text{kJ}}{\text{mol}}$.
- The heat of formation of $\text{CH}_3(\text{CH}_2)_2\text{OH}$ is $-104.6 \frac{\text{kJ}}{\text{mol}}$.

What suggestions can you give to the company regarding the conditions of pressure and temperature that will have a maximum yield of propanol at equilibrium without calculating the actual data?

Solution

$$\Delta H_{f(\text{propene})}^{\circ} = 20.4 \frac{\text{kJ}}{\text{mol}} \quad (11.E.30)$$

$$\Delta H_{f(\text{water})}^{\circ} = -241.8 \frac{\text{kJ}}{\text{mol}} \quad (11.E.31)$$

$$\Delta H_{f(\text{CH}_3(\text{CH}_2)_2\text{OH})}^{\circ} = -104.6 \frac{\text{kJ}}{\text{mol}} \quad (11.E.32)$$

$$\Delta H^{\circ} = \Delta H_{f(\text{CH}_3(\text{CH}_2)_2\text{OH})}^{\circ} - \Delta H_{f(\text{propene})}^{\circ} - \Delta H_{f(\text{water})}^{\circ} \quad (11.E.33)$$

$$\begin{aligned} &= -104.6 \frac{\text{kJ}}{\text{mol}} - (-241.8 \frac{\text{kJ}}{\text{mol}}) - 20.4 \frac{\text{kJ}}{\text{mol}} \\ &= 116.8 \frac{\text{kJ}}{\text{mol}} > 0 \end{aligned} \quad (11.E.34)$$

It is an endothermic reaction. As a result, a higher temperature is required to maximize the yield of product.

Also, according to the ideal gas law, because the moles of gas are proportional to the volume, and the volume is inverse proportional to the pressure; a higher pressure is required to maximize the yield of product.

Q.61

Suppose heptane and octane form an ideal solution. At 20.0 °C pure heptane has a vapor pressure of $P_1^{\circ} = 0.0562$ atm while pure octane has a vapor pressure of $P_2^{\circ} = 0.0145$ atm. What is the mole fraction of heptane vapor over a solution of heptane and octane if the solution contained 0.300 mol heptane and 0.700 mol octane?

Solution

Since we have been told to assume the solution is ideal, we can use [Raoult's law](#) to obtain the partial pressures of heptane and octane, respectively.

$$P_{\text{heptane}} = \left(\frac{0.300 \text{ mol}}{0.300 \text{ mol} + 0.700 \text{ mol}} \right) \times (0.0562 \text{ atm}) = 0.01686 \text{ atm}$$

$$P_{\text{octane}} = \left(\frac{0.700 \text{ mol}}{0.300 \text{ mol} + 0.700 \text{ mol}} \right) \times (0.0145 \text{ atm}) = 0.01015 \text{ atm}$$

Using the [relationship between partial pressure and mole fraction](#), we can obtain the mole fraction of heptane over the solution.

Partial pressure and mole fraction can be related through the ideal gas law $P = n \frac{RT}{V}$. Which shows that under constant temperature and pressure, pressure and number of moles are directly related.

$$\chi_{\text{heptane}} = \left(\frac{0.01686 \text{ atm}}{0.01686 \text{ atm} + 0.01015 \text{ atm}} \right) = 0.624$$

Q.63

At 50°C, benzene (C_6H_6) has a vapor pressure of 42.5 kPa and isobutene (C_4H_8) has a vapor pressure of 0.83 torr. When 10.0 g of benzene and 30 g of isobutene are mixed in a solution, they form a nearly ideal solution.

- Calculate the mole fraction of benzene in the solution.
- What is the total vapor pressure above the solution in kPa at 50°C ?
- Calculate the mole fraction of isobutene that exists above the solution in the vapor phase.

Solution

The mole fraction is equal to:

$$\left(\frac{X_A \text{ mol}}{X_A + X_B \text{ mol}} \right)$$

Use dimensional analysis to convert grams to moles:

$$(10.0 \cancel{\text{ g C}_6\text{H}_6}) \left(\frac{1 \text{ mol}}{78.108 \cancel{\text{ g C}_6\text{H}_6}} \right) = 0.128 \text{ mol}$$

$$(30.0 \cancel{\text{ g C}_3\text{H}_8}) \left(\frac{1 \text{ mol}}{44.094 \cancel{\text{ g C}_3\text{H}_8}} \right) = 0.680 \text{ mol}$$

Let:

X_A = the moles of $\text{C}_6\text{H}_6 = 0.128 \text{ mol}$, X_B = the moles of $\text{C}_3\text{H}_8 = 0.680 \text{ mol}$,

$$\chi_B = \frac{0.0929 \text{ kPa}}{6.81 \text{ kPa}} = 0.0136$$

Plug into the equation and solve:

$$\left(\frac{0.128 \text{ mol}}{0.128 + 0.680 \text{ mol}} \right) = 0.158$$

b) Since the problem states that the result of mixing the two solutions is nearly ideal, we can assume an ideal solution is formed and use Raoult's Law and we can choose to keep pressure in units of kPa OR torr.

It's [Raoult's Law](#):

$$P_A = \chi_A \times P_A^\circ$$

Here we have chosen to use units of kPa:

$$P_{\text{C}_6\text{H}_6} = \chi_{\text{C}_6\text{H}_6} \times P_{\text{C}_6\text{H}_6}^\circ$$

$$P_{\text{C}_6\text{H}_6} = 0.158 \times 42.5 \text{ kPa} = 6.72 \text{ kPa}$$

To convert torr to kPa:

$$(0.83 \cancel{\text{ torr}}) \left(\frac{0.133 \text{ kPa}}{1 \cancel{\text{ torr}}} \right) = 0.1104 \text{ kPa}$$

$$P_{\text{C}_3\text{H}_8} = \chi_{\text{C}_3\text{H}_8} \times P_{\text{C}_3\text{H}_8}^\circ$$

$$P_{\text{C}_3\text{H}_8} = 0.842 \times 0.1104 \text{ kPa} = 0.0929 \text{ kPa}$$

Then we add these up to find the total pressure:

$$P_{\text{Total}} = \chi_A P_A^\circ + \chi_B P_B^\circ$$

$$P_{\text{Total}} = 6.72 \text{ kPa} + 0.0929 \text{ kPa} = 6.81 \text{ kPa}$$

c) To find the mole fraction of isobutene **only** in the vapor phase, we can use partial pressure to represent the amount of gas, because in mixed gas $n \propto p$:

$$\chi_{\text{isobutene}} = \frac{n_{\text{isobutene}}}{n_{\text{tot}}} = \frac{P_{\text{isobutene}}}{P_{\text{tot}}} = \frac{0.83 \text{ torr}}{6.81 \text{ kPa}} = 0.0162$$

Q.65

A solution is prepared by mixing 2g of $\text{CaCl}_2(\text{s})$, 2g of $\text{MgCl}_2(\text{s})$, and 3g of $\text{MnO}_2(\text{s})$ in enough water to make a total volume of 150 mL. Calculate the concentration of chlorine in the solution in grams per liter.

Solution

$$m_{\text{Cl}} = 2g \text{ CaCl}_2 \left(\frac{2(35.5g \text{ Cl})}{110.98g \text{ CaCl}_2} \right) + 2g \text{ MgCl}_2 \left(\frac{2(35.5g \text{ Cl})}{95.12g \text{ MgCl}_2} \right)$$

$$m_{\text{Cl}} = 2.77g \text{ Cl}$$

$$\text{mass of chlorine per liter: } \frac{2.77g \text{ Cl}}{150 \times 10^{-3} L} = 18.5 \frac{g}{L}$$

Q.69

Chromium ion, Cr^{2+} , is a good reducing agent, often being itself turned into Cr^{3+} . Suppose that 10 mL of 0.1 M $\text{Cr}(\text{OH})_2$ was needed to reduce completely an 0.05 grams of unknown substance X.

- If each molecule of X accept just one electron, what is the molecular weight of X?
- If each molecule of X accept five electrons, what is the molecular weight of X?

Solution

$$\text{(a) Number of moles of } \text{Cr}(\text{OH})_2 \text{ used} = M \times V$$

$$= 0.1 \times 10 \text{ mL} = 1 \text{ mmol} = 10^{-3} \text{ moles}$$

Since only one mole of X was used for every mole of Chromium Hydroxide, the number of moles reacted are equal.

$$\frac{\text{Mass}}{\text{Molar mass}} = \text{number of moles, hence for X,}$$

$$\frac{0.05 \text{ g}}{(0.01L \times 0.1M)} = 50 \frac{g}{\text{mol}}$$

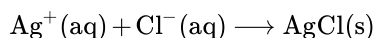
$$\text{(b) } \frac{0.05 \text{ g}}{(0.01L \times 0.1M \times 5)} = 10 \frac{g}{\text{mol}}$$

Q.71

A 0.200 g sample of chloride with unknown purity was dissolved in 200 mL of water was then titrated with AgNO_3 solution with a molarity of 0.1 M. After the titration it was determined that 28 mL of AgNO_3 was required to fully titrate the sample of chloride. What's the mass percent of chloride in the original dry sample?

Solution

First write the precipitation reaction:



Next we need to determine the amount of moles of Ag used during the titration:

$$\text{mol Ag}^+ = 28 \cancel{\text{mL}} \left(\frac{1 \cancel{\text{L}}}{1000 \cancel{\text{mL}}} \right) \left(\frac{0.1 \text{ mol}}{1 \cancel{\text{L}}} \right) = 2.80 \times 10^{-3} \text{ mol Ag}^+$$

Now that we know the amount of moles of Ag reacted because Ag and Cl combine at a one to one ration we can find the mass of Cl used:

$$\text{mass Cl}^- = 2.80 \times 10^{-3} \cancel{\text{mol}} \left(\frac{35.45 \text{ g}}{1 \cancel{\text{mol}}} \right) = 9.92 \times 10^{-2} \text{ g Cl}^-$$

Now that we know the mass of Cl in the unknown we can calculate the mass percent:

$$100\% \left(\frac{0.0992 \text{ g}}{0.2g} \right) = 49.63\%$$

Q.73

At 20°C, the vapor pressure of pure water is 17.5 mmHg. When 110 g of a substance X_y is dissolved in 500 g of water, the vapor pressure of the solution is 15.0 mmHg. Given that this substance has a molecular mass of 12.011 g/mol, what is the molecular formula of this substance X_y . (i.e. what is the number value of subscript y).

Solution

For this question, we can use [Raoult's Law](#)

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

First, we plug in the known values of the pressure of the solution and the pressure of the pure solvent, water, to find the mole fraction.

$$15.0 \text{ mmHg} = \chi_{\text{solvent}} \times 17.5 \text{ mmHg}$$

$$\chi_{\text{solvent}} = 0.8571$$

Next, we can calculate the the number of moles of the solvent from its mass and molecular weight.

$$\text{Moles of solvent} = 500 \text{ g} \times \frac{1}{18.015 \text{ g/mol}} = 27.75 \text{ moles}$$

Next, we can calculate the moles of the solute from the mole fraction of the solvent

$$0.8571 = \frac{27.75 \text{ moles}}{27.75 \text{ moles} + \text{moles of } X_y}$$

$$\text{moles of } X_y = 4.63 \text{ moles}$$

From here, we can calculate the molecular formula of the substance by dividing the amount of grams per mole used in the solution by the molecular mass of the substance.

$$\frac{110 \text{ g}}{4.63 \text{ moles}} = 23.76 \text{ g/mol}$$

$$(23.76 \text{ g/mol}) \times \left(\frac{1}{12.011 \text{ g/mol}} \right) = 1.98$$

$$1.98 \approx 2$$

Thus, the molecular formula of our substance is X_2 .

Q.81

The cell membrane of a plant cell is permeable to water. In a salt solution, the plant cells shrivel (lose water) when the concentration is high, and swell (take up water) when the concentration is low. Scientists perform an experiment at 30°C , where an aqueous solution of NaCl with a freezing point of -0.028°C is used to water the plant cells. This causes the plant cells to neither swell nor shrink indicating that the osmotic pressure of the plant cell contents is equal to that of the NaCl solution. Assuming the molarity and molality of the plant cell contents are equal, calculate the osmotic pressure of the solution inside the plant cells. The K_f for the solution is $1.86 \frac{\text{K kg}}{\text{mol}}$.

Solution

Plug in the given values into the equation for freezing point depression to solve for the molality of the solution

$$\Delta T_f = -K_f m$$

$$-0.0218 = -1.86 K \cdot \text{kg} \cdot \text{mol}^{-1} m$$

$$m = 0.015 \frac{\text{kg}}{\text{mol}}$$

Because the plant cell neither swells nor shrinks, the molality of the solution is equivalent to the molality of the plant cell contents. According to the question, assume the molarity and molality are equal so concentration

$$c = 0.15 M$$

Use the following equation to calculate for the osmotic pressure, π .

$$\begin{aligned}\pi &= cRT \\ &= (0.15M)(0.08206 \frac{L \text{ atm}}{\text{mol K}})(303.15K) \\ &= 0.374 \text{ atm}\end{aligned}$$

Q.83

Jim loves the smell of grass. He loves it so much that he frequently sniffs the fumes from 1-Hexanol, which is a chemical compound that is known to smell like grass trimmings. Because of his love, he decides to make his garage into a “gas den”, and aimed to have 60% of the air in it as 1-Hexanol gas. He achieves this by filling a pool with a mixture of 1-Hexanol and Water. Assume that the air in the “gas den” consists of only the fumes from this mixture. If the pool has 10000L of water in it, how many litres of 1-Hexanol does he need? Use Henry’s Law to determine this amount.

Relevant Information:

1-Hexanol:

- Density: $814 \frac{kg}{m^3}$
- Molar mass: $107.1748 \frac{g}{mol}$
- Henry’s Law Constant: $64 \frac{molal}{bar}$

Water:

- Density: $1000 \frac{kg}{m^3}$

Solution

This problem asks for the amount, in litres, of 1-Hexanol that is required to make Jim’s dreams come true. The information given is the density of 1-Hexanol and its molar mass, but since they don’t have very much to do with pressure, they can be set aside for now. The question explicitly mentions that Henry’s Law should be used, so start by identifying what Henry’s Law actually does:

$$C = kP$$

Where C is concentration, k is Henry’s Law constant, and P is partial pressure. The question gives out the fact that Jim wants the air in his garage to be 60% 1-Hexanol vapor. Remember the total pressure of air is $1 \text{ atm} \approx 1 \text{ bar}$ (ignore the error here). Therefore, the partial pressure of 1-hexanol is $1 \text{ bar} \times 60\% = 0.6 \text{ bar}$. The Henry’s Law constant is also provided, and by plugging those two in, the concentration of 1-Hexanol in the desired “solution” (more accurately, giant pool of chemicals) is found. This process is shown below:

$$C = kP = (64 \frac{molal}{bar}) \times 0.6 \text{ bar} = 38.4$$

This concentration, from the units of k, is in *molality*, which is $\frac{mol_{solute}}{kg_{solvent}}$. Now all that is left is to convert molality to litres. By utilizing the molar mass of 1-Hexanol:

$$38.4 \frac{mol_{solute}}{kg_{solvent}} \times 107.1748 \frac{g}{mol} \approx 4115.51 \frac{g_{solute}}{kg_{solvent}}$$

We know that the solvent is water, and that water has a density of $1000 \frac{kg}{m^3}$. Also, the conversion rate of m^3 to L is 1:1000. Therefore, the equation can be solved as follows.

$$4115.51 \frac{g_{solute}}{kg_{solvent}} \times 1000 \frac{kg_{H_2O}}{m^3} \times \frac{m^3}{1000 L_{H_2O}} \approx 4115.51 \frac{g_{solute}}{L_{solvent}}$$

We now have the value for how many grams of solute (1-Hexanol) there should be per litre of solvent. simply divide this value by the density to obtain litres of 1-Hexanol per litre of water. Multiply this number by the amount of water in the pool to get the desired amount of solute that is needed for the solution! (note that the density was converted to $\frac{g}{L}$,

$$4115.51 \frac{g_{\text{solute}}}{L_{\text{solvent}}} \div \left(814000 \frac{g_{1\text{-hexanol}}}{L} \right) \times 10000 L_{H_2O} = 50.56 L$$

~~Therefore, Jim needs 50,560L of 1-Hexanol to put in his pool. That's going to take a big pool!~~

Therefore, Jim needs 50.56L of 1-Hexanol to put in his pool.

Please note that, while various units do appear in the conversions, H_2O can be considered as solute and 1-Hexanol as the solvent, so the units can cancel.

Abstract: Utilize henry's law to find concentration (molality). From that, convert molality to g_{solute} to L_{solute} .

Q.87

A 3.5 gram sample is decomposed into a compound containing nitrogen, hydrogen, and oxygen, together creating 4.00 g of NO_2 and 7.00 g of H_2O . The molar mass of the sample is $78.5 \frac{g}{mol}$. Find the molecular formula of this sample.

Solution

First, the moles of each Nitrogen, Hydrogen, and Oxygen in solution are calculated:

$$mol_N = 4g NO_2 \times \left(\frac{1 mol NO_2}{46.005 \frac{g}{mol}} \right) \times \left(\frac{1 mol N}{1 mol NO_2} \right) = 0.0869 mol$$

$$mass_N = (14.00 \frac{g}{mol})(0.0869 mol) = 1.2172g N$$

$$mol_H = 7g H_2O \times \left(\frac{1 mol H_2O}{18.02 \frac{g}{mol}} \right) \times \left(\frac{2 mol H}{1 mol H_2O} \right) = 0.7769 mol$$

$$mass_H = (1.01 \frac{g}{mol})(0.7769 mol) = 0.7847g H$$

From these values, moles of O can be calculated to be:

$$3.5 gram sample = grams_N + grams_H + grams_O$$

$$3.5 = 1.2172g + 0.7847g + grams_O$$

$$grams_O = 1.4981g$$

$$mol_O = \frac{1.4981g}{15.9994 \frac{g}{mol}} = 0.0936 mol$$

The mole values for each element create a ratio of $H_{0.7769}N_{0.0869}O_{0.0936}$, creating an empirical formula H_9NO

From here, the molecular formula can be established using the molar mass:

$$\text{Empirical formula mass} = 9(1.01 \frac{g}{mol H}) + (14.00 \frac{g}{mol N}) + (15.9994 \frac{g}{mol O})$$

$$\text{Empirical formula mass} = 39.0894 \frac{g}{mol}$$

$$\text{Sample Mass} = 78.5 \frac{g}{mol}$$

$$\frac{78.5 \frac{g}{mol}}{39.084 \frac{g}{mol}} = 2.00$$

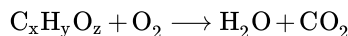
$$2(H_9NO) = H_{18}N_2O_2 = \text{Molecular Formula}$$

Q.87

5 grams of a substance containing hydrogen, carbon and oxygen were fully combusted and produced 3.00 grams of H_2O and 7.33 grams of CO_2 . When 10.64 grams of the substance was mixed with 0.5 kg of water, the freezing point was found to decrease by $0.0605^\circ C$. What is the substance?

Solution

The generic equation for this reaction is



The molar masses of carbon, hydrogen, and oxygen are 12.01 g/mol, 1.008 g/mol, and 16.00 g/mol respectively. Therefore the molar masses of water and carbon dioxide are $18.016 \frac{g}{mol}$ and $44.01 \frac{g}{mol}$. These values can be used to determine the mass percent of hydrogen in water

$$\frac{2(1.008) \frac{g}{mol}}{18.016 \frac{g}{mol}} \times 100 = 11.2\%$$

And the mass percent of carbon in carbon dioxide

$$\frac{12.01 \frac{g}{mol}}{44.01 \frac{g}{mol}} \times 100 = 27.3\%$$

The mass percents can be used to determine the amount of hydrogen and carbon are present.

$$3.00 \text{ g} \times 0.112 = 0.336 \text{ g of Hydrogen}$$

$$7.33 \text{ g} \times 0.273 = 2.001 \text{ g of Carbon}$$

By subtracting the masses of Carbon and Hydrogen from the original mass of the substance, the mass of oxygen present in the substance can be found

$$5 \text{ g} - (0.336 \text{ g} + 2.001 \text{ g}) = 2.663 \text{ g of Oxygen}$$

By dividing the masses present by the molar mass, the amount of moles of each element can be determined

$$\frac{0.336 \text{ g}}{1.008 \frac{g}{mol}} = 0.333 \text{ mol of Hydrogen}$$

$$\frac{2.001 \text{ g}}{12.01 \frac{g}{mol}} = 0.167 \text{ mol of Carbon}$$

$$\frac{2.663 \text{ g}}{16.00 \frac{g}{mol}} = 0.166 \text{ mol of Oxygen}$$

By looking at the amount of moles of each element, it can be noticed that there are equal amounts of carbon and oxygen and twice as much hydrogen as carbon or oxygen. Therefore the ratio of C : H : O in the substance is 1 : 2 : 1. However, the molar mass of the substance remains unknown, but it can be found using the information about the freezing point depression.

From the formula for freezing point depression

$$\Delta T = -i \times k_f \times m$$

$$-0.0605^\circ C = -1 \times 0.512^\circ C \text{ m}^{-1} \times m$$

$$m = \frac{0.0605^\circ C}{0.512 \frac{^\circ C}{m}} = 0.118 \frac{mol}{kg}$$

By multiply by the mass of solvent (water), we can determine the amount of moles that was dissolved in the water

$$0.118 \frac{mol}{kg} \times 0.5 \text{ kg} = 0.059 \text{ mol}$$

The last step in finding the molar mass of the substance is to divide the mass dissolved by the moles dissolved

$$\frac{10.64 \text{ g}}{0.059 \text{ mol}} = 180.3 \frac{g}{mol}$$

Now we can find the formula of the substance. The first step is to find the number of moles of the substance combusted

$$\frac{5 \text{ g}}{180.3 \frac{\text{g}}{\text{mol}}} = 0.0277 \text{ mol}$$

We can compare this to the number of carbon moles combusted to determine the number of carbon atoms per molecule of the substance

$$\frac{0.167 \text{ g of Carbon}}{0.0277 \text{ mol}} = 6 \text{ Carbon atoms}$$

When we apply the number of carbon atoms to the ratio of elements found earlier, the chemical formula comes to be $\text{C}_6\text{H}_{12}\text{O}_6$ which is glucose.

Answer: Glucose

11.E: Solutions (Exercises) is shared under a [CC BY-NC-SA 4.0](#) license and was authored, remixed, and/or curated by LibreTexts.