Marks 8

- Calcium chloride (3.42 g) is completely dissolved in 200 mL of water at 25.00 ºC in a 'coffee cup' calorimeter. The temperature of the water after dissolution is 27.95 ºC. Calculate the standard enthalpy of solution of $CaCl_2$ (in kJ mol^{-1}). The heat capacity of water is 4.184 J K^{-1} g^{-1}. Ignore the heat capacity of the $CaCl_2$.

As water has a density of 0.997 g mL^{-1}, 200 mL of water corresponds to a mass of:

$$m = \text{density} \times \text{volume} = (0.997 \text{ g mL}^{-1}) \times (200 \text{ mL}) = 199 \text{ g}$$

The temperature change, ΔT, is (27.95 – 25.00) ºC = 2.95 ºC (or 2.95 K)

The heat change accompanying a change in temperature is given by:

$$q = m \times C \times \Delta T = (199 \text{ g}) \times (4.184 \text{ J K}^{-1} \text{ g}^{-1}) \times (2.95 \text{ K}) = 2460 \text{ J} = 2.46 \text{ kJ}.$$

As the temperature of the water increases, the dissolution is exothermic and hence, ΔH = -2.46 kJ.

The formula mass of $CaCl_2$ is 40.08 (Ca) + (2 × 35.45 (Cl)) g mol^{-1} = 110.98 g mol^{-1}. Hence, 3.42 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{formula mass}} = \frac{3.42 \text{ g}}{110.98 \text{ g mol}^{-1}} = 0.0308 \text{ mol}$$

If this amount gives rise to an heat change of 2.46 kJ, the molar enthalpy change is given by $\frac{-2.46 \text{ kJ}}{0.0308 \text{ mol}} = -79.9 \text{ kJ mol}^{-1}$

Answer: **-79.9 kJ mol^{-1}**

ANSWER CONTIUNES ON THE NEXT PAGE

What would be the vapour pressure of water above this solution?
(P^0 (H_2O) = 3.17 kPa)

According to Raoult's Law, the vapour pressure of pure solvent, $P^o_{solvent}$ is lowered by the presence of solute by an amount ΔP:

$$\Delta P = i \times X_{solute} \times P^o_{solvent}$$

where X_{solute} is the mole fraction of the solute and i is the amount (mol) of particles in solution divided by the amount (mol) of dissolved solute.

The molar mass of H_2O is (2 × 1.008 (H)) + 16.00 (O) g mol^{-1} = 18.016 g mol^{1} so the number of moles of water in 199 g is:

$$\text{moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{199\,\text{g}}{18.016\,\text{g mol}^{-1}} = 36.0\,\text{mol}$$

As the solution contains 0.0308 mol of $CaCl_2$, the mole fraction of the solute is:

$$X_{solute} = \frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}} = \frac{0.0308}{(0.0308 + 36.0)} = 8.55 \times 10^{-4}$$

$CaCl_2$ dissolves to give three particles, Ca^{2+} + $2Cl^-$, so i = 3. Thus,

$$\Delta P = i \times X_{solute} \times P^o_{solvent} = 3 \times (8.55 \times 10^{-4}) \times 3.17\ \text{kPa} = 8.14 \times 10^{-3}\ \text{kPa}$$

The vapour pressure is lowered to (3.17 – 8.14 × 10^{-3}) kPa = 3.16 kPa.

Answer: **3.16 kPa**

ANSWER CONTIUNES ON THE NEXT PAGE

What would be the freezing point of this solution? The molal freezing point depression constant (K_f) for water is 1.86 °C kg mol^{-1}.

From above, 0.0308 mol of $CaCl_2$ are dissolved in 199 g of water. The molality is:

$$\text{molality} = \frac{\text{number of moles of solute (mol)}}{\text{mass of solvent (kg)}} = \frac{0.0308\ \text{mol}}{(199 \times 10^{-3}\ \text{kg})} = 0.155\ \text{mol kg}^{-1}$$

The freezing point depression, ΔT_f, is given by:

$$\Delta T_f = i \times K_f m$$

where K_f is the molal freezing point depression constant and m is the molality. As $i = 3$,

$$\Delta T_f = i \times K_f m = 3 \times (1.86\ °\text{C kg mol}^{-1}) \times (0.155\ \text{mol kg}^{-1}) = 0.865\ °\text{C}$$

At atmospheric pressure, the water freezes at 0 °C. The solution will freeze at -0.865 °C.

Answer: **-0.865 °C**

Which would you expect to cause the greater freezing point depression of water, 3.42 g of $CaCl_2$ or 3.42 g of NaCl? Explain your answer.

The formula mass of NaCl is 22.99 (Na) + 35.43 (Cl) g mol^{-1} = 58.42 g mol^{-1}. The number of moles of NaCl is therefore:

$$\text{number of moles} = \frac{\text{mass}}{\text{formula mass}} = \frac{3.42\ \text{g}}{58.42\ \text{g mol}^{-1}} = 0.0585\ \text{mol}$$

NaCl dissolves to give two particles, $Na^+ + Cl^-$ so $i = 2$.

As $\Delta T_f = i \times K_f m$, the freezing point depression of x kg of water is given by:

$$\Delta T_f = i \times K_f m = i \times K_f \times \frac{\text{moles of solute}}{x}$$

For NaCl, $\Delta T_f = 2 \times K_f \times \frac{0.0585}{x}$. For $CaCl_2$, $\Delta T_f = 3 \times K_f \times \frac{0.0308}{x}$.

Hence, ΔT_f α 0.117 for NaCl and ΔT_f α 0.0924 for $CaCl_2$. The freezing point depression is larger for NaCl.

Marks
6

- Assume that NaCl is the only significant solute in seawater. A 1.000 L sample of seawater at 25 ºC and 1 atm has a mass of 1.0275 kg and contains 33.0 g of NaCl. At what temperature would this seawater freeze? The freezing point depression constant of water is 1.86 ºC kg mol^{-1}.

The formula mass of NaCl is 22.99 (Na) + 35.45 (Cl) = 58.44. Therefore, 33.0 g corresponds to:

$$\textbf{number of moles of NaCl} = \frac{\text{mass}}{\text{formula mass}} = \frac{33.0}{58.44} = 0.565 \text{ mol}$$

As each mole of NaCl dissolves to give 2 moles of particles (Na^+(aq) and Cl^-(aq)), the number of moles of solute is 2 × 0.565 = 1.129 mol.

If salt water contains only water and NaCl,

mass of water = 1.0275 – 0.0330 = 0.995 kg

Hence, the molality is

$$m = \frac{\text{moles of solute}}{\text{mass of solvent}} = \frac{1.129}{0.995} = 1.136 \text{ mol kg}^{-1}$$

The freezing point depression is then:

$$\Delta T_f = K_f\, m = 1.86 \times 1.136 = 2.11\ °C$$

As water normally freezes at 0°C, this saltwater will freeze at -2.11 °C.

Answer: **-2.11 °C**

ANSWER CONTINUES ON THE NEXT PAGE

The vapour pressure above pure H_2O is 23.76 mmHg at 25 ºC and 1 atm. Calculate the vapour pressure above this seawater under the same conditions.

The molar mass of H_2O is (2 × 1.008 (H)) + 16.00 (O) = 18.016. Therefore, 0.995 kg of water corresponds to

$$\textbf{moles of water} = \frac{\text{mass}}{\text{molar mass}} = \frac{(0.995 \times 10^3)}{18.016} = 55.3\,\text{mol}$$

As 1.129 mol of solute is also present, the mole fraction, X, of water is

$$X_{water} = \frac{\text{number of moles of water}}{\text{total number of moles}} = \frac{55.3}{(55.3 + 1.129)} = 0.980$$

From Raoult's law,

$$P_{water} = X_{water}P_{water}^{o} = 0.980 \times 23.76 = 23.3\ \text{mmHg}$$

Answer: **23.3 mmHg**

The desalination of seawater by reverse osmosis has been suggested as a way of alleviating water shortages in Sydney. What pressure (in Pa) would need to be applied to this seawater in order to force it through a semi-permeable membrane, yielding pure H_2O?

The concentration of solute is:

$$\text{concentration} = c = \frac{\text{number of moles of solute}}{\text{volume}} = \frac{1.129}{1.000} = 1.129\,\text{M}$$

The osmotic pressure, Π, required is given by

$$\Pi = cRT = (1.129) \times (0.08206) \times (25 + 273) = 27.6\ \text{atm}$$

As 1 atm = 101.3 × 10^3 Pa,

$$\Pi = 27.6 \times (101.3 \times 10^3) = 2800000\ \text{Pa} = 2.80 \times 10^6\ \text{Pa}$$

Answer: **2.80×10^6 Pa**

Marks

- Lysozyme is an enzyme that breaks down bacterial cell walls. A solution containing 0.150 g of this enzyme in 210 mL of solution has an osmotic pressure of 0.00125 atm at 25 °C. What is the molar mass of lysozyme?

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The osmotic pressure, π, is given by $\pi = cRT$

Hence, if π = 0.00125 atm, the concentration at 25 °C is given by:

$$c = \frac{\pi}{RT} = \frac{(0.00125\,\text{atm})}{(0.08206\,\text{L atm K}^{-1}\,\text{mol}^{-1})\times((25+272)\,\text{K})} = 5.1\times10^{-5}\,\text{M}$$

As $c = \frac{n}{V}$, $n = cV = (5.1\times10^{-5}\ \text{mol L}^{-1})\times(0.210\ \text{L}) = 1.1\times10^{-5}\ \text{mol}$

This amount corresponds to 0.150 g, so the molar mass, *M*, is:

$$M = \frac{m}{M} = \frac{0.150\ \text{g}}{1.1\times10^{-5}\ \text{mol}} = 14000\ \text{g mol}^{-1}$$

Answer: **1.4×10^{4} g mol^{-1}**

- What mass of ethylene glycol, $HOCH_2CH_2OH$, is required to lower the freezing point of 1.00 L of water to –10.0 °C? The freezing point depression constant of water is 1.86 °C kg mol^{-1}. Assume the density of water is 1.00 g mL^{-1} at 0 °C.

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The freezing point depression, ΔT_f, is related to the molality, *m*, and the freezing point depression constant, K_f, by $\Delta T_f = K_f m$

Hence, $m = \frac{\Delta T_f}{K_f} = \frac{10.0\ °\text{C}}{1.86\ °\text{C kg mol}^{-1}} = 5.38\ \text{mol kg}^{-1}$

If the density of water is 1.00 g mL^{-1}, 1000 mL will have a mass of 1.00 kg.

As the molality is given $m = \frac{\text{amount of solute (mol)}}{\text{mass of solvent (kg)}}$**, the amount of solute is:**

amount of solute (mol) = molality (mol kg^{-1}) × mass of solvent (kg)
= 5.38 × 1.00 mol = 5.38 mol

The molar mass of $HOCH_2CH_2OH$ ($C_2H_6O_2$) is (2 × 12.01 (C)) + (6 × 1.008 (H)) + (2 × 16.00 (O)) = 62.068 g mol^{-1}. The mass of 5.38 mol is therefore:

mass (g) = molar mass (g mol^{-1}) × amount (mol) = 62.068 × 5.38 g = 334 g

Answer: **334 g**

Marks 5

- The freezing point of a sample of seawater is measured as –2.15 °C at 1 atm pressure. Assuming that the concentrations of other solutes are negligible, and that the salt does not significantly change the density of the water from 1.00 kg L^{-1}, determine the concentration (in mol L^{-1}) of NaCl in this sample. (The molal freezing point depression constant for H_2O is 1.86 °C m^{-1})

The freezing point depression, ΔT_f, is given by,

$$\Delta T_f = K_f m$$

where K_f is the molal freezing point depression and m is the molality. The molality is the number of moles of particles dissolved in a kilogram of solvent.

If ΔT_f = 2.15 °C and K_f = 1.86 °C m^{-1}:

$$m = \Delta T_f / K_f = (2.15\ °C) / (1.86\ °C\ m^{-1}) = 1.156\ m^{-1} = 1.156\ mol\ kg^{-1}$$

A mole of NaCl dissolves to give two particles (Na^+ and Cl^-) so (1.156 / 2) mol = 0.578 mol of NaCl per kilogram of water is needed.

As the density of the solution is 1.00 kg L^{-1}, a kilogram of solution has a volume of one litre. Hence:

concentration required = 0.578 mol L^{-1}

Answer: **0.578 mol L^{-1}**

In principle, it would be possible to desalinate this water by pumping it into a cylindrical tower, and allowing gravity to push pure water through a semipermeable membrane at the bottom. At 25 °C, how high would the tower need to be for this to work? (The density of liquid Hg at 25 °C is 13.53 g cm^{-3}.)

The osmotic pressure, Π, is given by $\Pi = cRT$ where c is the concentration of the particles. From above, c = (2 × 0.578) mol L^{-1} and so:

$$\Pi = (2 \times 0.578\ mol\ L^{-1}) \times (0.08206\ atm\ L\ mol^{-1}\ K^{-1}) \times (298\ K) = 28.3\ atm$$

As 1 atm = 760 mmHg, this corresponds to (28.3 × 760) mmHg = 21500 mmHg.

Considering the relative densities of water and Hg, the height of water required to exert this pressure would be:

$$21500\ mmHg = (21500 \times \frac{13.53}{1.000})\ mmH_2O = 291000\ mmH_2O \text{ or } 291\ mH_2O.$$

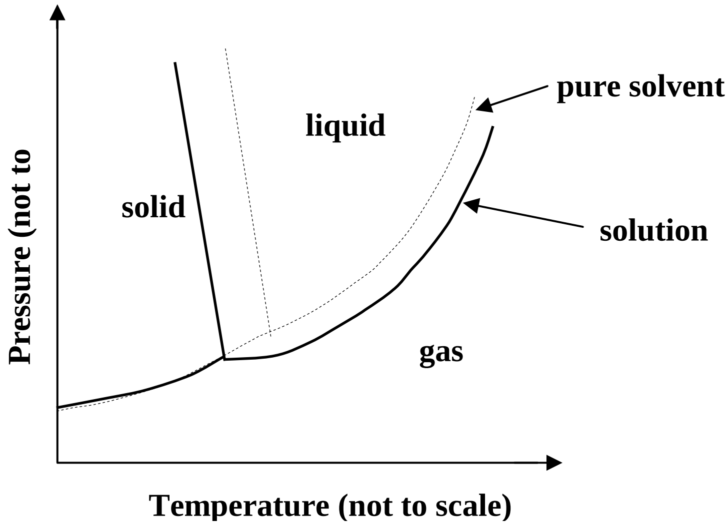
The tower would need to be 291 m in height.

Answer: **291 m**

Marks 3

- Explain why the freezing temperature of an aqueous salt solution is lower than that of pure water.

The presence of solute particles lowers the vapour pressure of the solution compared to that of the pure solvent. This results in a lowering of the freezing point as shown in the phase diagram.



What mass of sugar (sucrose, MW 342 g mol^{-1}) would have to be dissolved in 1.0 L of water to lower the freezing point as much as a water solution containing 11.1 g L^{-1} of $CaCl_2$?

The molar mass of $CaCl_2$ is (40.08 (Ca) + 2 × 35.45 (Cl)) g mol^{-1} = 110.98 g mol^{-1}. The number of moles in 11.1 g of $CaCl_2$ is therefore:

$$\textbf{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{11.1\ \text{g}}{110.98\ \text{g mol}^{-1}} = 0.100\ \text{mol}$$

As $CaCl_2$ dissolves to give 3 particles per mole (Ca^{2+} + $2Cl^-$), the number of moles of sucrose required is (3 × 0.100 mol) = 0.300 mol. The corresponds to a mass of:

mass = number of moles × molar mass = (0.300 mol) × (342 g mol^{-1}) = 103 g

Answer: **103 g**

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.