Marks
3

- 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, determine the molality (in mol kg^{-1}) of NaCl in this sample. The molal freezing point depression constant for H_2O is 1.86 °C kg mol^{-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 ions dissolved in a kilogram of solvent.

If $\Delta T_f = 2.15$ °C and $K_f = 1.86$ °C m^{-1}:

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

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:

$$m_{NaCl} = 0.578 \text{ mol kg}^{-1}$$

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

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

• Explain the following terms or concept.	**Mark** **1**

Osmotic pressure

The pressure that needs to be applied to prevent water from flowing across a semipermeable membrane due to osmosis.

	Marks 2
• Isooctane, an important constituent of petrol, has a boiling point of 99.3 °C and a standard enthalpy of vaporisation of 37.7 kJ mol^{-1}. What is $\Delta S°$ (in J K^{-1} mol^{-1}) for the vaporisation of isooctane?	

At the boiling point, $\Delta_{vap}G° = \Delta_{vap}H° - T\Delta_{vap}S° = 0$. Hence:

$\Delta_{vap}S° = \Delta_{vap}H° / T = (37.7 \times 10^3$ J mol^{-1}) / (99.3 + 273) K = +101 J K^{-1} mol^{-1}

Answer: +101 J K^{-1} mol^{-1}

**Marks
3**

- An aqueous solution with a volume of 10.0 mL contains 0.025 g of a purified protein of unknown molecular weight. The osmotic pressure of the solution was measured in an osmometer to be 0.0036 atm at 20.0 °C. Assuming ideal behaviour and no dissociation of the protein, estimate its molar mass in g mol^{-1}.

A pressure of 0.0036 atm corresponds to (0.0036 × $101.3 × 10^3$) Pa = 360 Pa. The osmotic pressure, Π, is related to the concentration of the solute through

$$\Pi = cRT$$

Hence:

$c = \Pi / RT$ = 360 Pa / ((8.314 Pa m^3 mol^{-1} K^{-1})((20.0 + 273) K))
     = 0.15 mol m^{-3} = 0.15 × 10^{-3} mol L^{-1}

The number of moles in 10.0 mL is therefore:

number of moles = concentration × volume
               = (0.15 × 10^{-3} mol L^{-1}) × (0.0100 L) = 1.5 × 10^{-6} mol

As this is the number of moles in 0.025 g, the molar mass is:

molar mass = mass / number of moles
          = 0.025 g / 1.5 × 10^{-6} mol = 17000 g mol^{-1}

Answer: **17000 g mol^{-1}**

• Explain the following term or concept.	**Marks**
	1

Vapour pressure

The partial pressure that builds up above the surface of a volatile liquid in a closed container due to evaporation.

- The concentration of NaCl used in intravenous drips is 150 mM. Explain why this particular concentration is used and what the consequences would be for a patient if pure water were used instead.

2

150 mM is used because this corresponds to the NaCl concentration in extracellular fluids and thus prevents any change in osmotic conditions in the blood. If water were used, water would flow into the red blood cells causing them to swell up and possibly burst. This may have fatal consequences for the patient.

- At 21.0 °C, a solution of 18.26 g of a non-volatile, non-polar compound in 33.25 g of bromoethane, CH_3CH_2Br, has a vapour pressure of 4.42×10^4 Pa. The vapour pressure of pure bromoethane at this temperature is 5.26×10^4 Pa. What is the molar mass of the compound?

4

The molar mass of CH_3CH_2Br is:

> **molar mass = (2 × 12.01 (C) + 5 × 1.008 (H) + 79.90 (Br)) g mol^{-1}**
>   **= 108.96 g mol^{-1}**

The number of moles of CH_3CH_2Br in 18.26 g is therefore:

> **$n_{solvent}$ = mass / molar mass = 33.25 g / 108.96 g mol^{-1} = 0.305 mol**

From Raoult's law, the vapour pressure of a solution, $P_{solution}$, is related to the vapour pressure of the pure solvent, $P^o_{solvent}$ and its mole fraction, $X_{solvent}$:

> **$P_{solution} = X_{solvent} \times P^o_{solvent}$**

Using $P_{solution} = 4.42 \times 10^4$ Pa and $P^o_{solvent} = 5.26 \times 10^4$ Pa, the mole fraction must be:
> **$X_{solvent} = P_{solution} / P^o_{solvent} = (4.42 \times 10^4 / 5.26 \times 10^4) = 0.840$**

The mole fraction of solvent is given by the number of moles of solvent divided by the total number of moles of solvent and solute:

> **$X_{solvent} = n_{solvent} / (n_{solvent} + n_{solute}) = 0.840$**

Using $n_{solvent} = 0.305$ mol from above,

> **$n_{solvent} / (n_{solvent} + n_{solute}) = 0.305 / (0.305 + n_{solute}) = 0.840$**

> **$n_{solute} = (0.305 / 0.840 – 0.305)$ mol = 0.0581 mol**

As this corresponds to a mass of 18.26 g, the molar mass is:

> **molar mass = mass / number of moles = 18.26 g / 0.0581 mol = 315 g mol^{-1}**

Answer: **315 g mol^{-1}**

Marks 4

• Order the following salts from lowest to highest molar solubility.

Salt	CuCl	$Cd(IO_3)_2$	$BaSO_4$	Ag_2CrO_4
K_{sp}	1.9×10^{-7}	2.3×10^{-8}	1.1×10^{-10}	2.6×10^{-12}

(1) For CuCl, K_{sp} refers to the dissolution reaction:

$$CuCl(s) \rightleftharpoons Cu^+(aq) + Cl^-(aq) \qquad K_{sp} = [Cu^+(aq)][Cl^-(aq)]$$
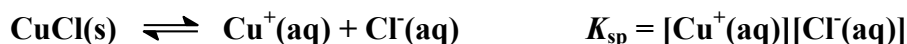
If the solubility is x mol L^{-1} then x mol of CuCl(s) dissolves in one litre. From the dissolution equation, this leads to $[Cu^+(aq)] = x$ M and $[Cl^-(aq)] = x$ M. Hence,

$$K_{sp} = (x)(x) = x^2 = 1.9 \times 10^{-7} \quad \text{or } x = (1.9 \times 10^{-7})^{1/2} \text{ M} = 4.4 \times 10^{-4} \text{ M}$$
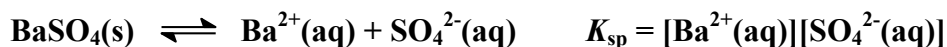
(2) For $Cd(IO_3)_2$, K_{sp} refers to the dissolution reaction:

$$Cd(IO_3)_2(s) \rightleftharpoons Cd^{2+}(aq) + 2IO_3^-(aq) \qquad K_{sp} = [Cd^{2+}(aq)][IO_3^-(aq)]^2$$
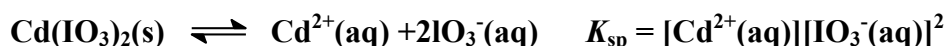
If the solubility is x mol L^{-1} then x mol of $Cd(IO_3)_2$ (s) dissolves in one litre. From the dissolution equation, this leads to $[Cd^{2+}(aq)] = x$ M and $[IO_3^-(aq)] = x$ M. Hence,

$$K_{sp} = (x)(2x)^2 = 4x^3 = 2.3 \times 10^{-8} \quad \text{or } x = (2.3 \times 10^{-8} / 4)^{1/3} \text{ M} = 1.8 \times 10^{-3} \text{ M}$$
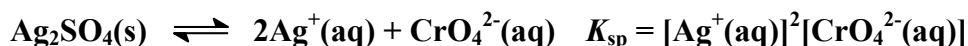
(3) For $BaSO_4$, K_{sp} refers to the dissolution reaction:

$$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq) \qquad K_{sp} = [Ba^{2+}(aq)][SO_4^{2-}(aq)]$$

If the solubility is x mol L^{-1} then x mol of $BaSO_4$(s) dissolves in one litre. From the dissolution equation, this leads to $[Ba^{2+}(aq)] = x$ M and $[SO_4^{2-}(aq)] = x$ M. Hence,

$$K_{sp} = (x)(x) = x^2 = 1.1 \times 10^{-10} \quad \text{or } x = (1.1 \times 10^{-10})^{1/2} \text{ M} = 1.0 \times 10^{-5} \text{ M}$$
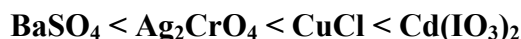
(4) For Ag_2CrO_4, K_{sp} refers to the dissolution reaction:

$$Ag_2SO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq) \qquad K_{sp} = [Ag^+(aq)]^2[CrO_4^{2-}(aq)]$$

If the solubility is x mol L^{-1} then x mol of Ag_2CrO_4(s) dissolves in one litre. From the dissolution equation, this leads to $[Ag^+(aq)] = 2x$ M and $[CrO_4^{2-}(aq)] = x$ M. Hence,

$$K_{sp} = (2x)^2(x) = 4x^3 = 2.6 \times 10^{-12} \quad \text{or } x = (2.6 \times 10^{-12} / 4)^{1/2} \text{ M} = 8.7 \times 10^{-5} \text{ M}$$

Overall, solubility increases in the order:

$$BaSO_4 < Ag_2CrO_4 < CuCl < Cd(IO_3)_2$$

- Isooctane, an important constituent of petrol, has a boiling point of 99.3 °C and an enthalpy of vaporisation of 37.7 kJ mol^{-1}. What is ΔS (in J K^{-1} mol^{-1}) for the vaporisation of isooctane?

2

At the boiling point, the system is at equilibrium and so $\Delta G = 0$:

$\Delta G = \Delta H - T\Delta S = 0$ **or** $\Delta S = \Delta H / T$

As $T_{\text{boiling}} = (99.3 + 273.0)$ K = 372.3 K,

$\Delta S = (+37.7 \times 10^3 \text{ J mol}^{-1}) / (373.2 \text{ K}) = +101 \text{ J K}^{-1} \text{ mol}^{-1}$

Answer: **+101 J K^{-1} mol^{-1}**

- An aqueous solution with a volume of 10.0 mL contains 0.025 g of a purified protein of unknown molecular weight. The osmotic pressure of the solution was measured in an osmometer to be 0.0036 atm at 20.0 °C. Assuming ideal behaviour and no dissociation of the protein, estimate its molar mass.

3

The osmotic pressure is given by

$$\Pi = cRT$$

As 1 atm = 1.013×10^5 Pa, the osmotic pressure is

$$\Pi = (0.0036 \times 1.013 \times 10^5) \text{ Pa} = 365 \text{ Pa}$$

Hence

$$c = \Pi / RT = (365 \text{ Pa}) / ((8.314 \text{ m}^3 \text{ Pa K}^{-1} \text{ mol}^{-1}) \times (293 \text{ K})) = 0.15 \text{ mol m}^{-3}$$

As 1 m^3 = 1000 L, this corresponds to

$$c = 1.5 \times 10^{-4} \text{ mol L}^{-1}$$

As 1 L contains 1.5×10^{-4} mol, the amount in 10.0 mL is:

amount of protein = (0.0100 L) $\times$ (1.5×10^{-4} mol L^{-1}) = 1.5×10^{-6} mol

As this amount has a mass of 0.025 g, the mass of 1 mol is:

molar mass = mass / number of moles
= (0.025 g) / (1.5×10^{-6} mol) = 17000 g mol^{-1}

Answer: **1.7×10^4 g mol^{-1}**

**Marks
4**

- Assuming ideal behaviour, calculate the mass of $MgCl_2 \cdot 6H_2O$ that should be dissolved in 1.0 L of water at 37 °C to obtain a solution with an osmotic pressure of 6.0 atm, the same as that of cell cytoplasm.

The molar mass of $MgCl_2 \cdot 6H_2O$ is:

(24.31 (Mg) + 2 × 35.45 (Cl) + 12 × 1.008 (H) + 6 × 16.00 (O)) g mol^{-1}

= 203.3 g mol^{-1}

An osmotic pressure of 6.0 atm corresponds to (6.0 × 101.3) kPa = 607.8 kPa.

The osmotic pressure, Π, is given by $\Pi = cRT$. Hence, the concentration, c, required is:

$c = \Pi / RT = (607.8 × 10^3 \text{ Pa}) / (8.314 \text{ m}^3 \text{ Pa K}^{-1} \text{ mol}^{-1} × ((37 + 273) \text{ K}))$

= 0.236 M

$MgCl_2 \cdot 6H_2O$ dissolves to give $Mg^{2+} + 2Cl^-$: three particles per mole. Hence, the number of moles of $MgCl_2 \cdot 6H_2O$ required to give this concentration of particles in 1.0 L is:

number of moles = 0.236 / 3 mol = 0.0786 mol

Hence, the mass of $MgCl_2 \cdot 6H_2O$ required is:

mass = number of moles × molar mass = 0.0786 mol × 203.3 g mol^{-1} = 16 g

Answer: **16 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$_2$O 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 $\Delta T_f = 2.15$ °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, $c = (2 \times 0.578)$ 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:

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

The tower would need to be 291 m in height.

Answer: **291 m**

**Marks
3**

- 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?

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

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

$$c = \frac{\pi}{RT} = \frac{0.00125\,atm}{(0.08206\,L\,atm\,K^{-1}\,mol^{-1}) \times ((25 + 273)\,K)} = 5.1 \times 10^{-5}\,M$$

As $c = \dfrac{n}{V}$, $n = cV = (5.1 \times 10^{-5}$ mol L$^{-1}) \times (0.210$ L$) = 1.1 \times 10^{-5}$ mol

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

$$M = \frac{m}{M} = \frac{0.150\,g}{1.1 \times 10^{-5}\,mol} = 1.40 \times 10^{4}\,g\,mol^{-1}$$

Answer: **1.40×10^{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.

3

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, $m = \dfrac{\Delta T_f}{K_f} = \dfrac{10.0\ °C}{1.86\ °C\ kg\ mol^{-1}} = 5.38$ 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 $m = \dfrac{amount\ of\ solute\,(mol)}{mass\ of\ solvent\,(kg)}$, 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
• 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}.	**6**

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

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

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:

$\Delta T_f = K_f\, m = 1.86 × 1.136 = 2.11$ °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 $(2 \times 1.008$ (H)) + 16.00 (O) = 18.016. Therefore, 0.995 kg of water corresponds to

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

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
4**

- Oral rehydration therapy (ORT) is a simple low-cost treatment that replaces fluid and electrolytes lost by sufferers of diarrhoea. To make the solution for ORT, 3.5 g NaCl, 2.9 g sodium citrate (which contains 1 citrate^{3-} and 3 Na^{+} ions and has a molar mass of 258 g mol^{-1}), 1.5 g KCl and 20.0 g glucose ($C_6H_{12}O_6$) are dissolved in water to make 1.0 L of solution. What is the osmotic pressure (in mmHg) of this solution at body temperature (37 °C)?

The osmotic pressure, Π, is given by $\Pi = i \times MRT$ where M is the molarity of the solution and i is the amount (mol) of particles in solution divided by the amount (mol) of dissolved solute. For this solution, the contributions from each dissolved species are summed, $\Pi = \Sigma (i \times MRT)$.

As NaCl dissolves to give two particles, Na^{+} + Cl^{-}, i_{NaCl} = 2. The formula mass of NaCl is 22.99 (Na) + 35.45 (Cl) g mol^{-1} = 58.44 g mol^{-1} so 3.5 g contains:

$$n_{NaCl} = \frac{mass}{formula\,mass} = \frac{3.5\,g}{58.44\,g\,mol^{-1}} = 0.060\,mol$$

$$M_{NaCl} = \frac{n}{V} = \frac{0.060\,mol}{1.0\,L} = 0.060\,M$$

As sodium citrate dissolves to give four particles, 3Na^{+} + citrate^{3-}, $i_{sodium\,citrate}$ = 4. The formula mass of sodium citrate is 258 g mol^{-1} so 2.9 g contains:

$$n_{sodium\,citrate} = \frac{2.9\,g}{258\,g\,mol^{-1}} = 0.011\,mol$$

$$M_{sodium\,citrate} = \frac{0.011\,mol}{1.0\,L} = 0.011\,M$$

As KCl dissolves to give two particles, K^{+} + Cl^{-}, i_{KCl} = 2. The formula mass of KCl is 39.10 (K) + 35.45 (Cl) g mol^{-1} = 74.55 g mol so 1.5 g contains:

$$n_{KCl} = \frac{1.5\,g}{74.55\,g\,mol^{-1}} = 0.020\,mol$$

$$M_{KCl} = \frac{0.020\,mol}{1.0\,L} = 0.020\,M$$

As glucose dissolves without dissociating, $i_{glucose}$ = 1. The molar mass of glucose is (6 × 12.01 (C)) + (12 × 1.008 (H)) + (6 × 16.00 (O)) g mol^{-1} = 180.156 g mol^{-1}. 20.0 g contains:

$$n_{glucose} = \frac{20\,g}{180.156\,g\,mol^{-1}} = 0.11\,mol$$

$$M_{glucose} = \frac{0.11\,mol}{1.0\,L} = 0.11\,M$$

The total osmotic pressure is therefore:

$$\Pi = \Pi_{NaCl} + \Pi_{sodium\ citrate} + \Pi_{KCl} + \Pi_{glucose}$$
$$= [(2\times0.060) + (4\times0.011)+(2\times0.020)+(1\times0.11)] \times (0.08206) \times (37.+ 273)\ \text{atm}$$
$$= 8.0\ \text{atm}$$

As 1 atm = 760 mmHg, 8.0 atm = (8.0 × 760) = 6.1 × 10^3 mmHg

	Answer: **6.1 × 10^3 mmHg**

This pressure is about the same as the osmotic pressure of blood. The calorie content of the solution can be increased by adding either more glucose or a polymer of glucose. Which would be preferable? Give a brief reason.

The polymer would be preferable as it would cause little change in the osmotic pressure of the solution.

- Ethylene glycol antifreeze, $C_2H_6O_2$, (1.00 kg) is added to a car radiator that contains 5.00 kg of water. What is the freezing point of the solution obtained?
  Data: The molal freezing point depression constant for water $K_f = 1.86$ °C kg mol^{-1}.

The molar mass of ethylene glycol is ((2 × 12.01 (C)) + (6 × 1.008 (H)) + (2 × 16.00 (O))) g mol^{-1}) = 62.068 g mol^{-1}. The number of moles in 1.00 kg is therefore:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{1.00 \times 10^3 \text{ g}}{62.068 \text{ g mol}^{-1}} = 16.1 \text{ mol}$$

The molality is:

$$\text{molality} = \frac{\text{number of moles of solute (mol)}}{\text{mass of solvent (kg)}} = \frac{16.1 \text{ mol}}{5.00 \text{ kg}} = 3.22 \text{ mol kg}^{-1}$$

The freezing point depression, ΔT_f, is given by:

$$\Delta T_f = K_f m$$

where K_f is the molal freezing point depression constant. Hence,

$$\Delta T_f = K_f m = (1.86 \text{ °C kg mol}^{-1}) \times (3.22 \text{ mol kg}^{-1}) = 5.99 \text{ °C}$$

At atmospheric pressure, the water freezes at 0 °C. The solution will freeze at -5.99 °C.

Marks
4

- A specific variety of haemoglobin associated with heart disease was isolated from a blood sample. A sample of this haemoglobin (21.5 mg) is dissolved in water at 25 °C to make 1.50 mL of solution. The osmotic pressure of the solution was measured and found to be 3.61 mmHg. What is the molar mass of this particular type of haemoglobin?

The osmotic pressure, Π, is related to the molarity M by $\Pi = MRT$

As 1 atm = 760 mmHg, Π = 3.81 mmHg = $\dfrac{3.81}{760}$ atm = 5.01×10^{-3} atm

Hence, $M = \dfrac{(5.01 \times 10^{-3}\text{ atm})}{(0.08206\,\text{L atm K}^{-1}\,\text{mol}^{-1}) \times ((25 + 273)\,\text{K})} = 2.05 \times 10^{-4}$ M

The number of moles of haemoglobin in 1.50 mL is thus:

$n(\text{haemoglobin}) = V \times c = \left(\dfrac{1.50}{1000}\,\text{L}\right) \times (2.05 \times 10^{-4}\text{ M}) = 3.08 \times 10^{-7}$ mol

This amount corresponds to 21.5 mg so the molar mass is:

$\text{molar mass} = \dfrac{\text{mass}}{\text{number of moles}} = \dfrac{(21.5 \times 10^{-3}\text{ g})}{(3.08 \times 10^{-7}\text{ mol})} = 6.99 \times 10^{4}$ g mol^{-1}

Answer: **6.99×10^{4} g mol^{-1}**

3

- Calcium oxalate (CaC_2O_4) is only slightly soluble in water (5.73 mg L^{-1} at 25 °C) and can be deposited in renal calculi (kidney stones). What is the molar solubility of calcium oxalate?

The formula mass is 40.08 (Ca) + (2×12.01 (C)) + (4×16.00 (O)) g mol^{1}
                                          = 128.1 g mol^{-1}

The molar solubility is = $\dfrac{\text{solubility}}{\text{formula mass}} = \dfrac{(5.73 \times 10^{-3}\text{ g L}^{-1})}{(128.1\,\text{g mol}^{-1})} = 4.47 \times 10^{-5}$M

Answer: **4.47×10^{-5} M**

Calculate the solubility product constant (K_{sp}) of calcium oxalate at 25 °C.

The solubility equilibrium is $CaC_2O_4(s) \rightleftharpoons Ca^{2+}(aq) + C_2O_4^{2-}(aq)$

As 1 mol of $Ca^{2+}(aq)$ and 1 mol of $C_2O_4^{2-}(aq)$ is produced by every mole of $CaC_2O_4(s)$ that dissolves, $[Ca^{2+}(aq)] = [C_2O_4^{2-}(aq)] = 4.47 \times 10^{-5}$ M.

Hence, $K_{sp} = [Ca^{2+}(aq)][C_2O_4^{2-}(aq)] = (4.47 \times 10^{-5}) \times (4.47 \times 10^{-5}) = 2.00 \times 10^{-9}$

$K_{sp} =$ **2.00×10^{-9}**