- Calculate the osmotic pressure of a 0.25 M aqueous solution of sucrose, $C_{12}H_{22}O_{11}$, at 37 °C 2

The osmotic pressure for strong electrolyte solutions is given by:

$$\boldsymbol{\Pi = i \times (cRT)}$$

where *i* is the amount (mol) of particles in solution divided by the amount (mol) of dissolved solute. For 0.25 M sucrose, *c* = 0.25 M and *i* = 1. Hence,

$$\begin{aligned} \Pi &= (0.25 \times 1 \text{ mol L}^{-1}) \times RT \\ &= (0.25 \text{ mol L}^{-1}) \times (0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times ((37 + 273) \text{ K}) = 6.4 \text{ atm} \end{aligned}$$

Answer: **6.4 atm**

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Answer: **6.4 atm**

Marks 4

- A saline solution used for intravenous injections contains 900 mg of sodium chloride in 100 mL. What is the concentration of this sodium chloride solution?

The molar mass of NaCl is (22.99 (Na) + Cl (35.45)) g mol^{-1} = 58.44 g mol^{-1}. The number of moles in 900 mg is therefore:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{900 \times 10^{-3}\ \text{g}}{58.44\ \text{g mol}^{-1}} = 0.0154\ \text{mol}$$

The concentration is therefore:

$$\text{concentration} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.0154\ \text{mol}}{0.100\ \text{L}} = 0.154\ \text{mol L}^{-1}$$

Answer: **0.154 M**

What is the osmotic pressure of this solution at 37 °C?

The osmotic pressure for strong electrolyte solutions is given by:

$$\Pi = i \times (cRT)$$

where i is the amount (mol) of particles in solution divided by the amount (mol) of dissolved solute.

For 0.154 M NaCl, c = 0.154 and i = 2 (as NaCl → Na^+ + Cl^-, each mole of NaCl produces two moles of particles). Hence:

$$\begin{aligned}\Pi &= (2 \times 0.154\ \text{mol L}^{-1}) \times RT \\ &= (2 \times 0.154\ \text{mol L}^{-1}) \times (0.08206\ \text{atm L K}^{-1}\ \text{mol}^{-1}) \times ((37 + 273)\ \text{K}) \\ &= 7.84\ \text{atm}\end{aligned}$$

Answer: **7.84 atm**

Why is it better to use a saline solution rather than pure water when administering drugs intravenously?

Saline solution is isotonic with blood plasma. Injection water would have a hypotonic effect and cause lysis of cells.

4

- Calculate the osmotic pressure of a solution of 1.0 g of glucose ($C_6H_{12}O_6$) in 1500 mL of water at 37 °C.

The osmotic pressure π = cRT where c is the concentration.

The molar mass of glucose is:

(6 × 12.01 (C)) + (12 × 1.008 (H)) + (6 × 16.00 (O)) = 180.156

1.0 g of glucose corresponds to $\frac{\text{mass}}{\text{molar mass}} = \frac{1.0}{180.156} = 0.0056 \text{ mol}$

The concentration when this amount is dissolved in 1500 mL = 1.5 L is:

$$c = \frac{\text{number of moles}}{\text{volume}} = \frac{0.056}{1.5} = 0.0037 \text{ M}$$

Hence, π = cRT = (0.0037) × (0.08206) × (273 + 37) = 0.094 atm.

Answer: **0.094 atm**

Explain why a drip for intravenous administration of fluids is made of a solution of NaCl at a particular concentration rather than pure water.

Blood plasma is isotonic with cells (same osmotic pressure). Using saline drip of same osmotic pressure as blood prevents haemolysis or crenation of red blood cells.

- The solubility of nitrogen in water at 25 °C and 1.0 atm is 0.018 g L^{-1}. What is its solubility at 0.50 atm and 25 °C? 1

The equilibrium of interest is $N_2(g) \leftrightarrows N_2(aq)$ with equilibrium constant:

$$K = \frac{[N_2(aq)]}{[N_2(g)]}$$

From the perfect gas law, PV = nRT or concentration = $\frac{n}{V} = \frac{P}{RT}$.

As a result, doubling the pressure doubles the concentration of $N_2(g)$. As the temperature is unchanged, the equilibrium constant does not change and so $[N_2(aq)]$ must halve to ensure that K_p is constant.

The solubility is thus halved to ½ × 0.018 = 0.009 g L^{-1}

Answer: **0.009 g L^{-1}**

Marks
4

- In the spaces provided, explain the meanings of the following terms. You may use an equation or diagram where appropriate.

(a) hydrogen bonding

An unusually strong dipole-dipole interaction that forms when a hydrogen atom is bonded to one of the very electronegative atoms F, O or N.

(b) colligative properties

Properties of a solution that depend only upon the number of moles of solute present, not the nature of the solute.

(c) hypotonic solution

A solution with lower osmotic pressure than cell fluid.

(d) isoelectric point

The pH at which there is no nett charge on a molecule containing both acidic and basic groups.

- The osmotic pressure of a solution containing 5.5 g L^{-1} of a polypeptide is 0.103 atm at 5 °C. Calculate the molar mass of the polypeptide. **2**

The osmotic pressure, *Π*, is related to the concentration, *c*, and the temperature, *T*:

$$\Pi = cRT$$

where R is the gas constant. With *Π* = 0.103 atm and *T* = 5 °C = (273 + 5) K,

$$0.103 \text{ atm} = c \times (0.08206 \text{ L atm K}^{-1} \text{ mol}^{1}) \times (278 \text{ K})$$

$$c = 0.00452 \text{ mol L}^{-1}$$

where *R* = 0.08206 L atm K^{-1} mol^{1} has been used to remove the need to convert the pressure in Pascals.

As the solution contains 5.5 g L^{-1} of the polypeptide and this is equivalent to 0.00452 mol L^{-1}, 5.5 g must contain 0.00452 mol. Hence,

$$\text{molar mass} = \frac{5.5 \text{ g}}{0.00452 \text{ mol}} = 1200 \text{ g mol}^{-1} = 1.2 \times 10^{3} \text{ g mol}^{1}$$

Answer: **1200 g mol^{-1} = 1.2 × 10^{3} g mol^{1}**

- Henry's law describes the solubility of a gas in a liquid phase. What methods are possible to ensure a patient receives enough oxygen during surgery? Which method is the most practical? Explain. **3**

Henry's law states that the higher the pressure of gas above a liquid, the greater the solubilityof the gas in that liquid:

c = kP

Normal air is 21% O_2. Anaesthetists can ensure a patient receives enough O_2 during surgery by increasing the % (*i.e.* partial pressure) of O_2 in the gas the patient breathes. This is the most practical and easy approach.

The alternative would be to get the patient to breathe a mixture of air at a pressure greater than 1 atm, but this would be more difficult to control and could lead to other problems (*e.g.*"the bends").

- A saline solution used to administer drugs intravenously is prepared by dissolving 0.90 g NaCl in 100.0 mL water. What mass of glucose ($C_6H_{12}O_6$) is required to prepare a 100.0 mL solution with the same osmotic pressure? **2**

The formula mass of NaCl is (22.99 (Na) + 35.45 (Cl)) g mol^{-1} = 58.55 g mol^{-1}. The number of moles in 0.90 g is therefore:

$$\text{number of moles} = \frac{\text{mass}}{\text{formula mass}} = \frac{0.90\text{ g}}{58.55\text{ g mol}^{-1}} = 0.015\text{ mol}$$

Dissolution of NaCl(s) produces Na^+(aq) and Cl^-(aq) so the total number of moles of ions present is 2 × 0.015 mol = 0.030 mol. The concentration of ions in a solution with volume 100.0 mL is thus:

$$\text{concentration} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.030\text{ mol}}{0.1000\text{ L}} = 0.30\text{ mol L}^{-1} = 0.30\text{ M}$$

The osmotic pressure is related to the concentration by $\Pi = cRT$. As glucose does not dissociate in solution, to produce the same concentration of particles in solution requires 0.030 mol of glucose.

The molar mass of glucose, $C_6H_{12}O_6$ is (6 × 12.01 (C) + 12 × 1.008 (H) + 6 × 16.00 (O)) g mol^{-1} = 180.156 g mol^{-1}.

The mass of 0.030 mol of glucose is therefore:

$$\text{mass} = \text{number of moles} \times \text{molar mass} = (0.030\text{ mol}) \times (108.156\text{ g mol}^{-1}) = 5.4\text{ g}$$

Answer: **5.4 g**

Marks 4

- Henry's law relates the solubility of a gas to its pressure. *i.e.* $c = kp$

The Henry's law constant for $N_2(g)$ at 298 K is 6.8×10^{-4} mol L^{-1} atm^{-1}. A diver descends to a depth where the pressure is 5 atm. If the diver's body contains about 5 L of blood, calculate the maximum amount of nitrogen gas dissolved in the diver's blood at 1 atm and at 5 atm. (Assume solubility of nitrogen in water and blood to be the same.)

At the surface, p = 1 atm and so:

$$\mathbf{c = \mathit{k}p = (6.8 \times 10^{-4}\ mol\ L^{-1}\ atm^{-1}) \times (1\ atm) = 6.8 \times 10^{-4}\ mol\ L^{-1}}$$

Hence, the amount of N_2 in 5 L is

$$\mathbf{number\ of\ moles = concentration \times volume}$$
$$\mathbf{= (6.8 \times 10^{-4}\ mol\ L^{-1}) \times (5\ L) = 0.003\ mol}$$

When p = 5 atm and so:

$$\mathbf{c = (6.8 \times 10^{-4}\ mol\ L^{-1}\ atm^{-1}) \times (5\ atm) = 3.4 \times 10^{-3}\ mol\ L^{-1}}$$

Hence, the amount of N_2 in 5 L is

$$\mathbf{number\ of\ moles = (3.4 \times 10^{-4}\ mol\ L^{-1}) \times (5\ L) = 0.02\ mol}$$

1 atm: **3×10^{-3} mol**	5 atm: **2×10^{-2} mol**

If all the gas dissolved at 5 atm were suddenly released, what volume would it occupy at 1 atm and 298 K?

Using the ideal gas law, *PV* = *nRT*, with *n* = 2×10^{-2} mol:

$$\mathbf{\mathit{V} = \mathit{nRT} / \mathit{P} = (2 \times 10^{-2}\ mol)(0.08206\ L\ atm\ K^{-1}\ mol^{-1})(298\ K) / (1\ atm) = 0.4\ L}$$

Answer: **0.4 L**

Marks
5

- The concentration of a dissolved gas is related to its partial pressure by c = *k*p. What is the concentration of CO_2 dissolved in blood if the partial pressure of CO_2 in the lungs is 0.053 atm? The *k* for CO_2 is 0.034 mol L^{-1} atm^{-1}.

Using *c* = *k*p,

c = (0.034 mol L^{-1} atm^{-1})(0.053 atm) = 0.0018 mol L^{-1}

Answer: **0.0018 mol L^{-1}**

Calculate the pH of blood if all of this CO_2 reacted to give H_2CO_3.
The K_a of H_2CO_3 is 4.5 × 10^{-7}.

If $[H_2CO_3(aq)]$ = 0.0018 mol L^{-1}, the pH can be calculated using the reaction table:

	H_2CO_3	H_2O	$\rightleftharpoons$	H_3O^+	HCO_3^-
initial	0.0018	large		0	0
change	$-x$	negligible		$+x$	$+x$
final	$0.0018 - x$	large		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = \frac{x^2}{0.0018 - x}$$

As $K_a = 4.5 \times 10^{-7}$ and is very small, $0.0018 - x \sim 0.0018$ and hence:

$$x^2 = 0.0018 \times (4.5 \times 10^{-7}) \quad \text{or} \quad x = 2.8 \times 10^{-5}\ \text{M} = [H_3O^+]$$

Hence:

$$\text{pH} = -\log_{10} [H_3O^+(aq)] = -\log_{10}(2.8 \times 10^{-5}) = 4.54$$

Answer: **4.54**

Hyperventilation results in a decrease in the partial pressure of CO_2 in the lungs. What effect will this have on the pH of the blood? Use a chemical equation to illustrate your answer.

If the CO_2 partial pressure decreases, the equilibrium below will shift to the left. This will decrease $[H^+(aq)]$ and the pH will increase.

$$CO_2(aq) + H_2O \rightleftharpoons H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$$

ANSWER CONTINUES ON THE NEXT PAGE

The pH of blood is maintained around 7.4 by the H_2CO_3 / HCO_3^- buffer system. Explain how a buffer works, illustrating your answer with chemical equations.

A buffer resists changes in pH. It contains substantial quantities of a weak acid and its conjugate base. In the H_2CO_3/HCO_3^- buffer, added acid is removed by the reaction:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2CO_3(aq)$$

Added base is removed by the reaction:

$$H_2CO_3(aq) + OH^-(aq) \rightarrow HCO_3^-(aq) + H_2O$$