COLLIGATIVE PROPERTIES OF DILUTE SOLUTION

The solution in which the amount of solute is very less than the amount of solvent is known as dilute solution. That means in a dilute solution the amount of solute is negligible to that of the solvent. For example, if n_1 is the number of moles of solvent and n_2 is the number of moles of solute in a dilute solution then total number of moles in the solution would be considered as $n_1+n_2 \approx n_1$.

Colligative properties of dilute solution

The properties of dilute solution which depends on only the amount (number of moles) of molecules of solute present in the solution not on the chemical nature or size (molecular weight) of the solute are known as colligative properties of dilute solution. There are four colligative properties. They are as follows

1. *Lowering of vapor pressure:* The presence of a non-volatile solute lowers the vapor pressure of the solvent. This is because the solute particles occupy some of the surface area of the solution, reducing the number of solvent molecules that can escape into the gas phase.

2. *Elevation of boiling point:* The boiling point of a solution is higher than that of the pure solvent. This is because the solute particles interfere with the ability of the solvent molecules to escape into the gas phase, requiring a higher temperature to overcome the increased intermolecular forces.

3. *Depression of freezing point:* The freezing point of a solution is lower than that of the pure solvent. This is because the solute particles disrupt the crystal lattice formation of the solvent, making it more difficult for the solvent molecules to arrange themselves into a solid structure.

4. *Osmosis and osmotic pressure:* Osmotic pressure is the pressure required to prevent the flow of solvent across a semipermeable membrane into a solution. This pressure arises due to the difference in concentration of solute particles on either side of the membrane.

These properties are closely related to each other through a common explanation. For these reason they are grouped together under the class name colligative properties. Latin word *colligatus* means collected together or grouped together or bound together.

Lowering of vapor pressure is the most important among the colligative properties. Because the other properties can be explained with the help of it. If a solute is dissolved in a solvent its vapor pressure would be lower. For low vapor pressure the solution has to be heated to higher temperature to get its vapor pressure equal to the atmospheric pressure. That means its boiling point increases. Similarly freezing point decreases for lowering of vapor pressure. For the difference of vapor pressure between solvent and solution, osmotic pressure develops and solvent molecules move from solvent (lower concentration region) to solution (higher concentration region).

Note that, colligative properties are applicable for dilute solution and the solute must be non-volatile and non-electrolyte.

These colligative properties have various applications, including:

Determining the molecular weight of a solute: By measuring the change in a colligative property, such as freezing point depression or boiling point elevation, the molecular weight of an unknown solute can be determined.

Preserving food: Adding salt or sugar to food can lower its freezing point, preventing the growth of bacteria and fungi.

De-icing roads: Salt is added to roads in winter to lower the freezing point of water, preventing ice formation.

LOWERING OF VAPOR PRESSURE

When a small amount of a non-volatile solute is dissolved in a solvent, its vapor pressure decreases. That means the vapor pressure of resulting solution is lower than that of the pure solvent. If the vapor pressure of pure solvent is p^o and that of the resulting solution is p, then the lowering of vapor pressure is (p^o – p).
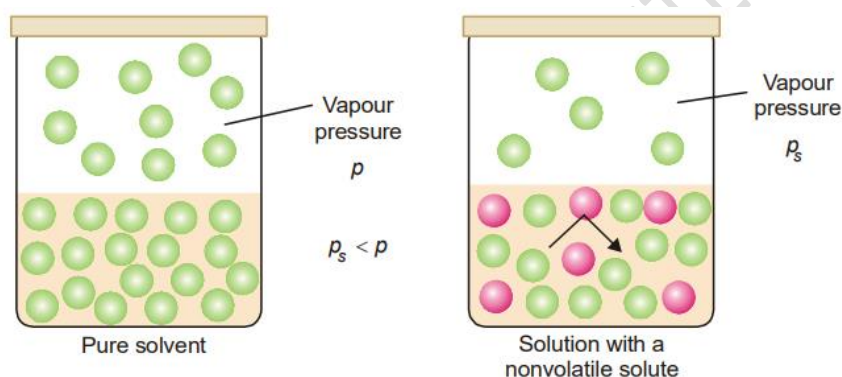

Figure 1: Lowering of vapour pressure by a nonvolatile solute. The particles of the solute block the escape of solvent molecules from the surface of the solution.

Vapor pressure of pure solvent, p^o

p^o

Lowering of vapor pressure, p^o – p

Vapor pressure of solution, p

p

Relative lowering of vapor pressure, $\frac{p^o - p}{p^o}$

Raoult's law of lowering of vapor pressure

When a small amount of a non-volatile solute is dissolved in a solvent, it is found that at a given temperature the vapor pressure of the resulting solution is less than that of the pure solvent. Studying on this, scientist F. M. Raoult established a relation between relative lowering of vapor pressure of the solution and the amount of solute (in mole fraction). This relation is well known as Rauolt's law of lowering of vapor pressure.

This law states that, *the relative lowering of vapor pressure of a dilute solution produced by dissolving a non-volatile solute is equal to the mole fraction of solute in the solution.*

Explanation: let,

n_1 = number of moles of solvent

n_2 = number of moles of solute

$$mole\ fraction\ of\ solvent, x_1 = \frac{n_1}{n_1 + n_2}$$

$$mole\ fraction\ of\ solute, x_2 = \frac{n_2}{n_1 + n_2}$$

p^o = vapor pressure of pure solvent

p = vapor pressure of solution

Lowering of vapor pressure = $p^o - p$ (NB: $p^o > p$)

Relative lowering of vapor pressure = $\frac{p^o - p}{p^o}$

According to the **Rauolt's law of lowering of vapor pressure**,

$$\frac{p^o - p}{p^o} = \frac{n_2}{n_1 + n_2}$$

$$\frac{p^o - p}{p^o} = x_2$$

The alternative expression and statement of Raoult's law: From above equation we can write

$$1 - \frac{p}{p^o} = x_2$$

$$\frac{p}{p^o} = 1 - x_2$$

$$\frac{p}{p^o} = x_1 \qquad \text{as } [x_1 + x_2 = 1]$$

$$\boxed{p = X_1 \times p^o}$$

The alternative statement of Raoult's law, *the vapor pressure of solution is equal to mole fraction of solvent multiplied by vapor pressure of pure solvent at the same temperature.*

Determination of molecular weight of solute using Rauolt's law

Using Rauolt's law of lowering of vapor pressure molecular weight of a solute can be determined. From Raoult's law we know that

$$\frac{p^o - p}{p^o} = \frac{n_2}{n_1 + n_2}$$

For very dilute solution N+n≈N, so

$$\frac{p^o - p}{p^o} = \frac{n_2}{n_1}$$

If w_2 g of solute of molecular weight M_2 is dissolved in w_1 g of solvent of molecular weight M_1, then

$$number\ of\ moles\ of\ solute, n_2 = \frac{w_2}{M_2}$$

$$number\ of\ moles\ of\ solvent, n_1 = \frac{w_1}{M_1}$$

From the above equation we can write

$$\frac{p^o - p}{p^o} = \frac{w_2/M_2}{w_1/M_1}$$

$$\boxed{M_2 = \frac{w_2 M_1}{w_1} \times \frac{p^o}{p^o - p}}$$

Knowing the experimental value of $(p^o - p)/p^o$ and molecular weight of solvent (M_1), molecular weight of solute (M_2) can be determined using the above equation.

Problem: The vapor pressure of ether (mw=74) is 442 mm Hg at 30 °C. If 3.0 g of a compound A is dissolved in 50.0 g of ether at this temperature, the vapor pressure falls to 426 mm Hg. Assuming a very dilute solution, calculate the molecular weight of A. (Ans 123)

Problem: 18.2 g of urea is dissolved in 100.0 g of water at 50 °C. The lowering of vapor pressure is 5.0 mm Hg. Calculate the molecular weight of urea. The vapor pressure of pure water at 50 °C is 92.0 mm Hg. (Ans 57.05)   60.26gm/mol
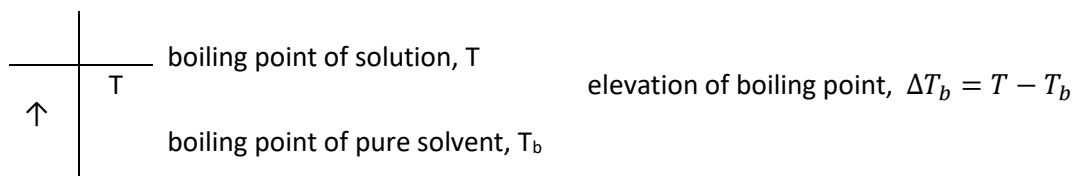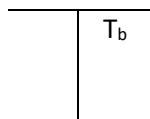
Deviation from Raoult's law

Elevation of boiling point

When a liquid is heated, its vapor pressure rises and when it equals to the atmospheric pressure, the liquid boils. The temperature at which the boiling occurs is the boiling point of the liquid. So *boiling point is the temperature at which the vapor pressure of a liquid is equal to the atmospheric pressure.*

Addition of a non-volatile solute lowers the vapor pressure and consequently elevates the boiling point as the solution has to be heated to higher temperature to make its vapor pressure equal to atmospheric pressure. If the boiling point of pure solvent is T_b and that of solution is T, the difference between the boiling points (ΔT_b) is called the elevation of boiling point.

$$\Delta T_b = T - T_b$$

boiling point of solution, T

elevation of boiling point, $\Delta T_b = T - T_b$

boiling point of pure solvent, T_b

T_b

The vapor pressure-temperature curves of pure solvent, and solution-1 and solution-2 with different concentrations of solute are shown below.
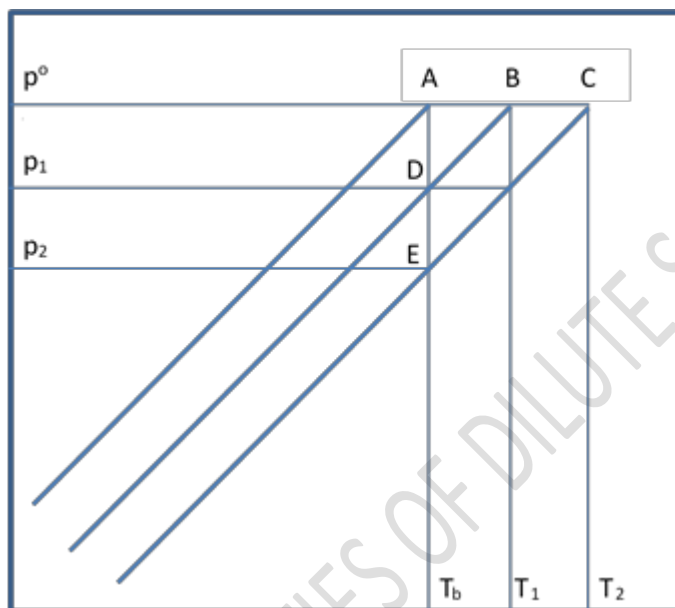

Figure 2: The vapor pressure-temperature curves of pure solvent, and solution-1 and solution-2 with different concentrations of solute.

For dilute solutions, the curves BD and CE are approximately straight line and parallel. Therefore for similar triangles of ABD and ACE, we have

$$\frac{AB}{AC} = \frac{AD}{AE}$$

$$\frac{T_1 - T_b}{T_2 - T_b} = \frac{p^o - p_1}{p^o - p_2}$$

$$\frac{\Delta T_{b1}}{\Delta T_{b2}} = \frac{p^o - p_1}{p^o - p_2}$$

or, $\frac{\Delta T_{b1}}{\Delta T_{b2}} = \frac{(p^o - p_1)/p^o}{(p^o - p_2)/p^o}$

Where $p^o - p_1$ and $p^o - p_2$ are lowering of vapor pressure for solution-1 and solution-2 respectively. Hence the elevation of boiling point is directly proportional to the lowering of vapor pressure.

$$\Delta T_b \propto p^o - p$$

Or, $\Delta T_b \propto \frac{(p^o - p)}{p^o}$

Determination of molecular weight of solute from elevation of boiling point

We know that, elevation of boiling point is directly proportional to the lowering of vapor pressure. Since p is constant for the same solvent at a fixed temperature, from above equation we can write,

$$\Delta T_b \propto \frac{p^o - p}{p}$$

But from Raoult's law for dilute solution,

$$\frac{p^o - p}{p} = x_2$$

$$\frac{p^o - p}{p} = \frac{n_2}{n_1 + n_2}$$

$$\frac{p^o - p}{p} = \frac{n_2}{n_1} \qquad\qquad for\ dilute\ solution$$

$$\frac{p^o - p}{p} = \frac{w_2/M_2}{w_1/M_1}$$

Hence

$$\Delta T_b \propto \frac{w_2/M_2}{w_1/M_1}$$

$$\Delta T_b = K \frac{w_2/M_2}{w_1/M_1} \qquad\qquad here, K\ is\ a\ proportionality\ constant$$

$$\Delta T_b = \frac{K \times M_1}{1000} \times \frac{w_2 \times 1000}{M_2 \times w_1}$$

$$\Delta T_b = K_b \frac{w_2 \times 1000}{M_2 \times w_1} \qquad Where,\ another\ constant, K_b = \frac{K \times M_1}{1000}$$

If w_2 g of solute of molecular weight M_2 is dissolved in w_1 g of solvent, then the molality of the solution by definition

$$C_m = \frac{w \times 1000}{m \times W}$$

So we can write,

$$\Delta T_b = K_b \times C_m$$

Here K_b is called boiling point constant or molal elevation constant or ebullioscopic constant. It can be defined as elevation of boiling point when one mole of solute in dissolved in one kg of solvent (one molal solution). In other word, K_b is the elevation of boiling point for one molal solution.

Problem: The boiling point of a solution containing 0.2 g of a substance in 20.0 g of ether as solvent is 0.17 K higher than that of pure solvent. Calculate the molecular weight of the substance. K_b = 2.16 K. (Ans 127.06).

Depression of freezing point

On cooling a dilute solution, a temperature is ultimately reached at which solid solvent begins to separate from the solution. The temperature at which the precipitation from the solid begins is the freezing point of the solution. At this temperature, solid solvent and solution are in equilibrium and consequently they must have the same vapor pressure.

It is observed that freezing point of a solution is lower than that of pure solvent. This is a direct consequence of the lowering of vapor pressure of the pure solvent by the dissolved solute. So depression of freezing point is the difference between the freezing point of the pure solvent T_o and that T_f of solution.

$\Delta T_f = T_o - T_f$	Here,
	ΔT_f = depression of freezing point
	T_o = Freezing point of pure solvent
	T_f = Freezing point of solution

freezing point of pure solvent, T_o

T_o

↓

freezing point of solution, T_f

T_f

Depression of freezing point, $\Delta T_f = T_o - T_f$

Relation between depression of freezing point and lowering of vapor pressure

The vapor pressure of a pure liquid changes with temperature as shown by the curve ABC, in Figure 3. There is a sharp break at point B, where in fact, freezing point curve commences. Thus the point B corresponds to the freezing point of pure solvent, T_f. The vapor pressure curve of a solution (solution 1) of a non-volatile solute in the same solvent is also shown in the Figure 3. It is similar to the vapor pressure curve of the pure solvent and meets the freezing point at F, indicating that T_1 is the freezing point of the solution. The difference of the freezing point of the pure solvent and the solution 1 is referred to as the **depression of freezing point**. It is represented by the symbol ΔT or ΔT_f.

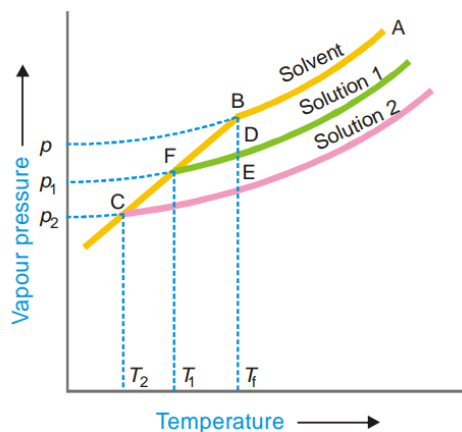Depression of freezing point, $\Delta T = T_f - T_1$

Figure 3. Relation between lowering of vapour-pressure and depression of freezing point.

When more of the solute is added to the solution 1, we get a more concentrated solution (solution 2). The vapor pressure of solution 2 meets the freezing point at C, indicating a further lowering of freezing point to T_2.

For dilute solution FD and CE are approximately parallel straight lines and BC is also a straight line. Since the triangles BDF and BEC are similar,

$$\frac{FD}{CE} = \frac{BD}{BE}$$

$$\frac{T_f - T_1}{T_f - T_2} = \frac{p^o - p_1}{p^o - p_2}$$

Where, p_1 and p_2 are vapor pressure of solution 1 and solution 2 respectively. Hence depression of freezing point is directly proportional to the lowering of vapor pressure. So

$$\Delta T_f \alpha \, (p^o - p)$$

Determination of molecular weight of solute from depression of freezing point.

Since p_o is constant for the same solvent at a fixed temperature, from the above equation we can write

$$\Delta T_f \alpha \, \frac{(p - p)}{p}$$

……….not written here………

WHAT IS OSMOSIS?

Let us consider a pure solvent and solution separated by a membrane which permits the passage to solvent molecules but not to solute molecules. Only the solvent will diffuse through the membrane into solution. A membrane which is permeable to solvent and not to solute, is called a semipermeable membrane.
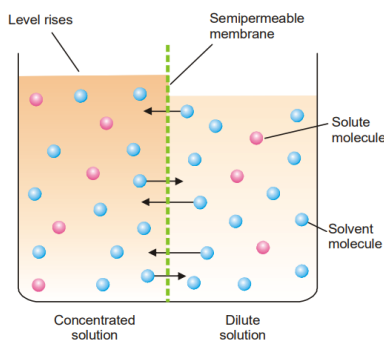
Figure 4: Illustration of the process of osmosis. The solvent molecules can pass through the membrane from either side of the semipermeable membrane. There is a net flow of solvent from the dilute solution to the concentrated solution. It shows up in the rise of level in the concentrated solution compartment.

The flow of the solvent through a semipermeable membrane from pure solvent to solution, or from a dilute solution to concentrated solution, is termed Osmosis (Greek Osmos = to push). This process continues until the concentration of solute on both sides of the membrane is equal.
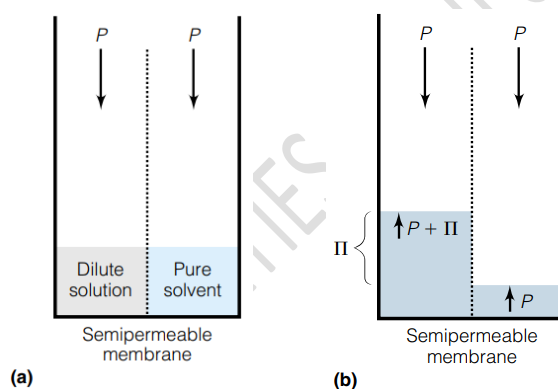

Figure 5: The two-part system is filled with pure solvent on one side and a dilute solution on the other. (a) Initially, the liquid levels are even with each other. However, it is not at equilibrium. Solvent will pass through the semipermeable membrane in a preferential direction. (b) At equilibrium, the two levels are uneven. The difference between the two levels is defined as the osmotic pressure.

Key points about osmosis:

> **Semipermeable membrane:** A membrane that allows certain molecules to pass through while blocking others.
> Solvent: The substance that dissolves the solute.
> Solute: The substance that is dissolved in the solvent.
> Concentration gradient: The difference in concentration between two regions.
> **Passive process:** Osmosis occurs without the need for energy input.

Importance of Osmosis:

❖ **Biological processes:** Osmosis plays a crucial role in many biological processes, such as the uptake of water by plants and the regulation of blood pressure in animals.

- ❖ Cell function: Osmosis helps maintain the balance of water and solutes within cells.
- ❖ Food preservation: Osmosis is used in food preservation techniques like salting and pickling.
- ❖ Water purification: Reverse osmosis is a process used to purify water by removing impurities.

LAWS OF OSMOTIC PRESSURE

From a study of the experimental results obtained by Pfeffer, van't Hoff showed that for dilute solutions:

a) The osmotic pressure of a solution at a given temperature is directly proportional to its concentration.
b) The osmotic pressure of a solution at a given concentration is directly proportional to the absolute temperature.

From the above findings, van't Hoff (1877) established the laws of osmotic pressure and pointed out that these were closely related to the gas laws.

Boyle-van't Hoff Law for Solutions

If π is the osmotic pressure and C its concentration, from (a) we can write,

$$\pi \propto C$$, if temperature (T) is constant.

If the concentration of the solute is expressed in moles per litre and V is the volume of the solution that contains 1 mole of solute,

$$C = \frac{1}{V}$$

$$\pi \propto \frac{1}{V}$$

This relationship is similar to the Boyle's law for gases and is known as the Boyle-van' t Hoff law for solutions.

Charles-van't Hoff Law for Solutions

If T is the absolute temperature, from the statement (b), we can write,

$$\pi \propto T$$, if temperature is constant.

This relationship is similar to the Charles' Law for gaseous state of matter and is known as Charles-van't Hoff law for solutions.

Van't Hoff Equation for Solutions

As shown above the osmotic pressure (π) of a dilute solution is inversely proportional to the volume (V) containing 1 mole of the solute and is directly proportional to the absolute temperature (T). This is,

$$\pi \propto \frac{1}{V} \quad \text{................. (1) and}$$

$$\pi \propto T \quad \text{.................. (2)}$$

Combining the above two equations (1) and (2), van't Hoff gave the general relationship,

$$\pi \propto \frac{T}{V}$$

$$\pi = R'\frac{T}{V}$$

$$\pi V = R'T \quad \text{.................... (3)}$$

Equation (3) is known as ***van't Hoff Equation.*** Where R' is a proportionality constant. He showed that this equation was parallel to the general Gas Equation (PV = RT), as the value of R' calculated from the experimental values of π, V, and T came out to be almost the same as of the *Gas constant, R*.

It is noteworthy that the van't Hoff Equation (3) was derived for 1 mole of solute dissolved in a volume of V litres. If n moles of solute are dissolved in V litres of solution, this equation assumes the form as follows,

$$\pi V = nRT$$

SOLVED PROBLEM. A solution of cane-sugar (mol mass = 342) containing 34.2 g litre^{-1} has an osmotic pressure of 2.4 atm at 20°C. Calculate the value of R in litre-atmospheres.

SOLUTION: From the van't Hoff equation

$$\pi V = RT$$

$$R = \frac{\pi V}{T}$$

Where,

π = osmotic pressure,

V = volume of solution in litres containing 1 mole of solute, and

T = absolute temperature. In the present case,

$$\pi = 2.4\ atm$$

$$V = \frac{1}{34.2} \times 342 = 10 \; litres$$

$$T = 20 + 273 = 293 \; K$$

Substituting the values in the expression,

$$R = \frac{2.4 \times 10}{293} = \mathbf{0.0819 \; litre\text{--} atmosphere \; \textit{K}^{-1}\textit{mole}^{-1}}$$

SOLVED PROBLEM 2. Experiment shows that 10 g of cane sugar (mol mass = 342) in $1 \times 10^{-3} \, m^3$ of solution produces an osmotic pressure of $6.68 \times 10^4 \, N \, m^{-2}$ at 273 K. Calculate the value of R in *SI* units.

SOLUTION

$$R = \frac{\pi V}{T} \qquad \qquad \text{...van't Hoff equation}$$

In this case,

$$\pi = 6.68 \times 10^4 \, N \, m^{-2}$$

$$V = \frac{1 \times 10^{-3} \times 342}{10} = 3.42 \times 10^{-2} \, m^3$$

Substituting the values in van't Hofff equation,

$$R = \frac{6.68 \times 10^4 \times 3.42 \times 10^{-2}}{273}$$

$$= \mathbf{8.36 \; J \; mol^{-1} \; K^{-1}}$$

SOLUTION

From van't Hoff equation

$$R = \frac{\pi V}{T} \qquad \qquad \text{...(1)}$$

where π = osmotic pressure, V = volume of solution in litres containing 1 mole of solute, T = absolute temperature.

In the present case,

$$\pi = 2.4 \, atm$$

$$V = \frac{1}{34.2} \times 342 = 10 \; litres$$

$$T = 20 + 273 = 293 \, K$$

Substituting the values in the expression (1),

$$R = \frac{2.4 \times 10}{293} = \mathbf{0.0819 \; litre \text{ - } atmosphere \; K^{-1} \, mol^{-1}}$$

Avogadro-van't Hoff Law for Solutions

Writing Van't Hoff equation for two different dilute solutions,

$$\pi_1 V_1 = n_1 RT_1 \quad \ldots\ldots\ldots\ldots (1)$$

$$\pi_2 V_2 = n_2 RT_2 \quad \ldots\ldots\ldots\ldots (2)$$

Where, n_1 and n_2 are the number of moles (molecules) of the solute in V_1 and V_2 litres of the two solutions respectively.

If $\pi_1 = \pi_2$ and $T_1 = T_2$, it follows from (1) and (2) that,

$$\frac{n_1}{V_1} = \frac{n_2}{V_2}$$

Hence when osmotic pressure and temperature are the same, equal volumes of solutions would contain equal number of moles (molecules) of the solute. This relationship is analogous to Avogadro's law of gases and is referred to as Avogadro-van't Hoff law for solutions. Alternatively, this law may be stated as:

Isotonic solutions at the same temperature have equimolar concentrations.

SOLVED PROBLEM 1. Calculate the osmotic pressure of a 5% solution of glucose (mol wt = 180) at 18°C.

SOLUTION $\quad\quad\quad\quad\quad\quad \pi V = nRT \quad\quad\quad\quad\quad\quad\quad$ ...van't Hoff Equation

where π = osmotic pressure in atmospheres; V = volume in litres; n = number of moles of solute (w/M), w being the weight in grams and M its molecular weight; R = gas constant.

In this case :

$$\pi = ?$$

$$V = 100\,\text{ml} = \frac{1}{10}\,\text{litre}$$

$$n = \frac{5}{M} = \frac{5}{180}$$

$$R = 0.0821 \text{ litre-atmosphere deg}^{-1}\text{ mole}^{-1}$$

$$T = 273 + 18 = 291 \text{ K}$$

Substituting the values in van't Hoff Equation

$$\pi \times \frac{1}{10} = \frac{5}{180} \times 0.0821 \times 291$$

$$\pi = \textbf{6.64 atm}$$

SOLVED PROBLEM 2. Find the osmotic pressure in millimetres of mercury at 15°C of a solution of naphthalene ($C_{10}H_8$) in benzene containing 14 g of naphthalene per litre of solution.

SOLUTION

$$\pi V = nRT \qquad \text{...vant Hoff equation}$$

Here,

π = ? atmosphere

V = 1 litre

$n = \dfrac{14}{128}$ (mol mass of naphthalene = 128)

R = 0.0821 litre-atmosphere $K^{-1}\,mol^{-1}$

T = 15 + 273 = 288 K

Substituting values in van't Hoff equation,

$$\pi \times 1 = \frac{14}{128} \times 0.0821 \times 288$$

$$\pi = 2.586 \,atm$$

$$= 2.586 \times 760 = \textbf{1965 mm}$$

DETERMINATION OF MOLECULAR WEIGHT FROM OSMOTIC PRESSURE

Osmotic pressure measurements are less accurate and difficult to carry out. Therefore, this method is not widely used in the determination of molecular weights. The chief use of this method is in finding molecular weights of polymers e.g., plastics, proteins, and starch, which have very high molecular weights.

Knowing the osmotic pressure of a given solution, the molecular weight of the solute can be calculated as follows from van't Hoff equation.

M = molecular mass of the solute

w = amount of solute in grams

R = 0.0821 litre-atmosphere $K^{-1}mole^{-1}$

T = (t°C + 273) K π = osmotic pressure in atmospheres

V = volume of solution in litres

SOLVED PROBLEM 1. Morse and Frazer observed that a solution of glucose containing 18 grams per litre had an osmotic pressure of 2.39 atmospheres at 23°C. Find the molecular weight of glucose.

SOLUTION

We know that:

$$M = \frac{wRT}{\pi V} \qquad \qquad ...(1)$$

Here,

$w = 18\,g$

$R = 0.0821$ litre-atmosphere

$T = 23+273 = 296\,K$

$\pi = 2.39$ atmospheres

$V = 1$ litre

Substituting these values in relation (1)

$$M = \frac{18 \times 0.0821 \times 296}{2.39 \times 1}$$

$$= \textbf{183.0}$$

SOLVED PROBLEM 2. A solution of glycol containing 1.821 g per litre has an osmotic pressure of 51.8 cm of mercury at 10°C. What is the molecular mass of glycol?

SOLUTION

We know that from van't Hoff equation

$$M = \frac{wRT}{\pi V} \qquad \qquad ...(1)$$

In this case,

$w = 1.821g$

$R = 0.0821$ litre-atmosphere

$T = 10+273 = 283\,K$

$\pi = \dfrac{51.8}{76}$ atmosphere

$V = 1$ litre

Substituting these values in expression (1),

$$M = \frac{1.821 \times 0.0821 \times 283 \times 76}{51.8 \times 1}$$

$$= \textbf{62.07}$$

Q1. Derive the Van't Hoff equation for solution.

Q2. Define lowering of vapor pressure and elevation of boiling point. Explain the interrelation between them as colligative properties.

*Q. **State and explain** Rauolt's law of lowering of vapor pressure.*

*Q. **How to determine the molecular weight of a solute from the** Rauolt's law of lowering of vapor pressure.*

Q3. What is the relation between vapor pressure and boiling point? Clarify your answer with examples.

Q4. Why the vapor pressure of a solution is lower than that of the pure solvent?

Q5. Why the boiling point of a solution is higher than that of the pure solvent?

Q6. Define vapor pressure, lowering of vapor pressure and relative lowering of vapor pressure.

Q7. Define boiling point, elevation of boiling point.

Q7. Define freezing point, depression of freezing point.

Q3. Lowering of vapor pressure, elevation of boiling point, depression of freezing point, and osmotic pressure are osmotic pressure. Explain from the mathematical expression.

Q. Give some examples of colligative properties in daily life.