BYJU'S Classes

Solutions and Colligative Properties

Solution

Na^+Cl^-

Homogeneous mixture i.e., a single phase containing more than one component dispersed on a molecular scale.

Homogeneous Mixture

When salt is dissolved in water it forms a homogeneous mixture. Composition is same throughout the salt solution.

Solutions around us

Soda can

Saline solution

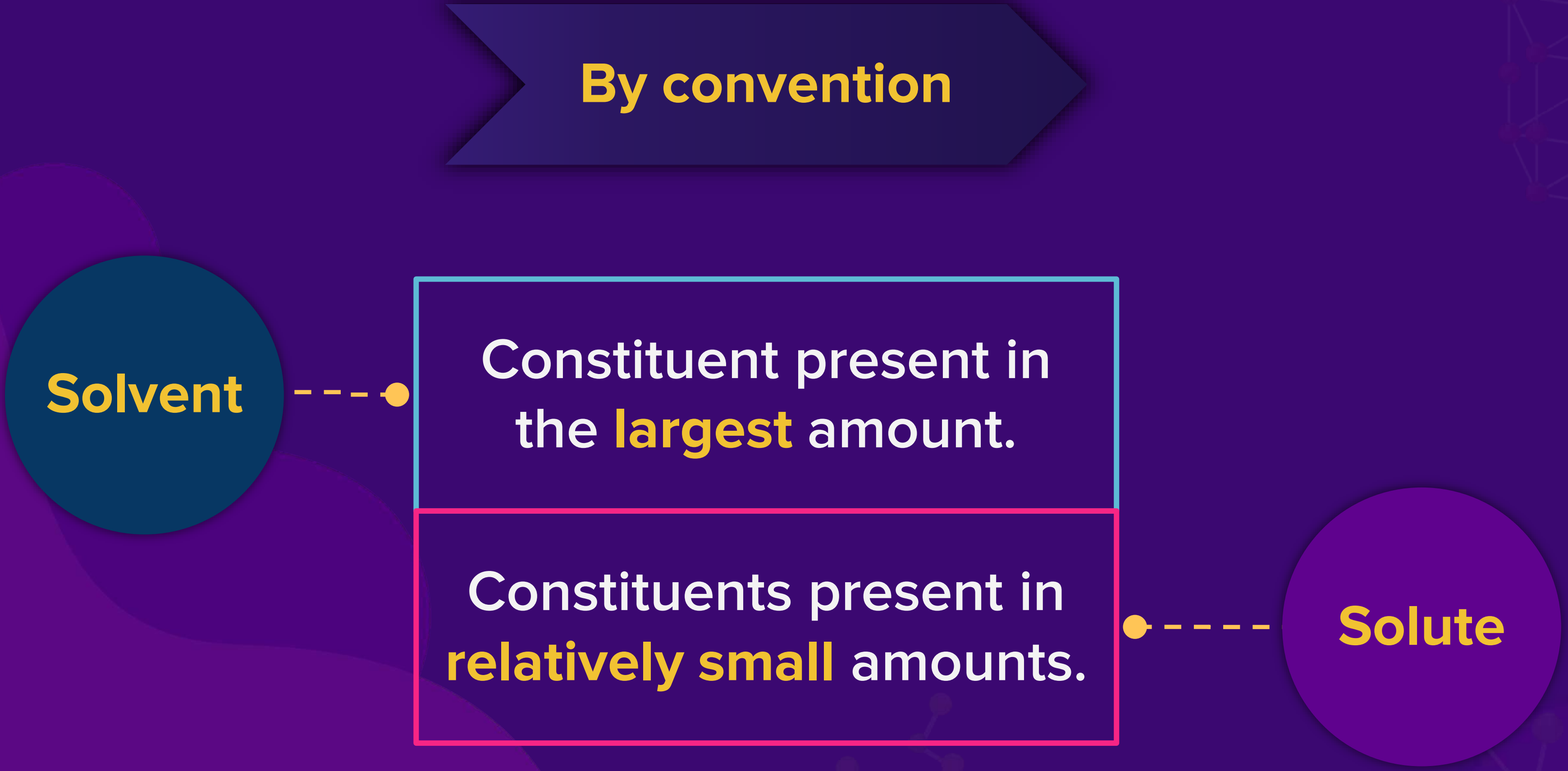
Components of a Solution

Solute

Solvent

Solution

Components of a Solution



By convention

Solvent

Constituent present in the **largest** amount.

Solute

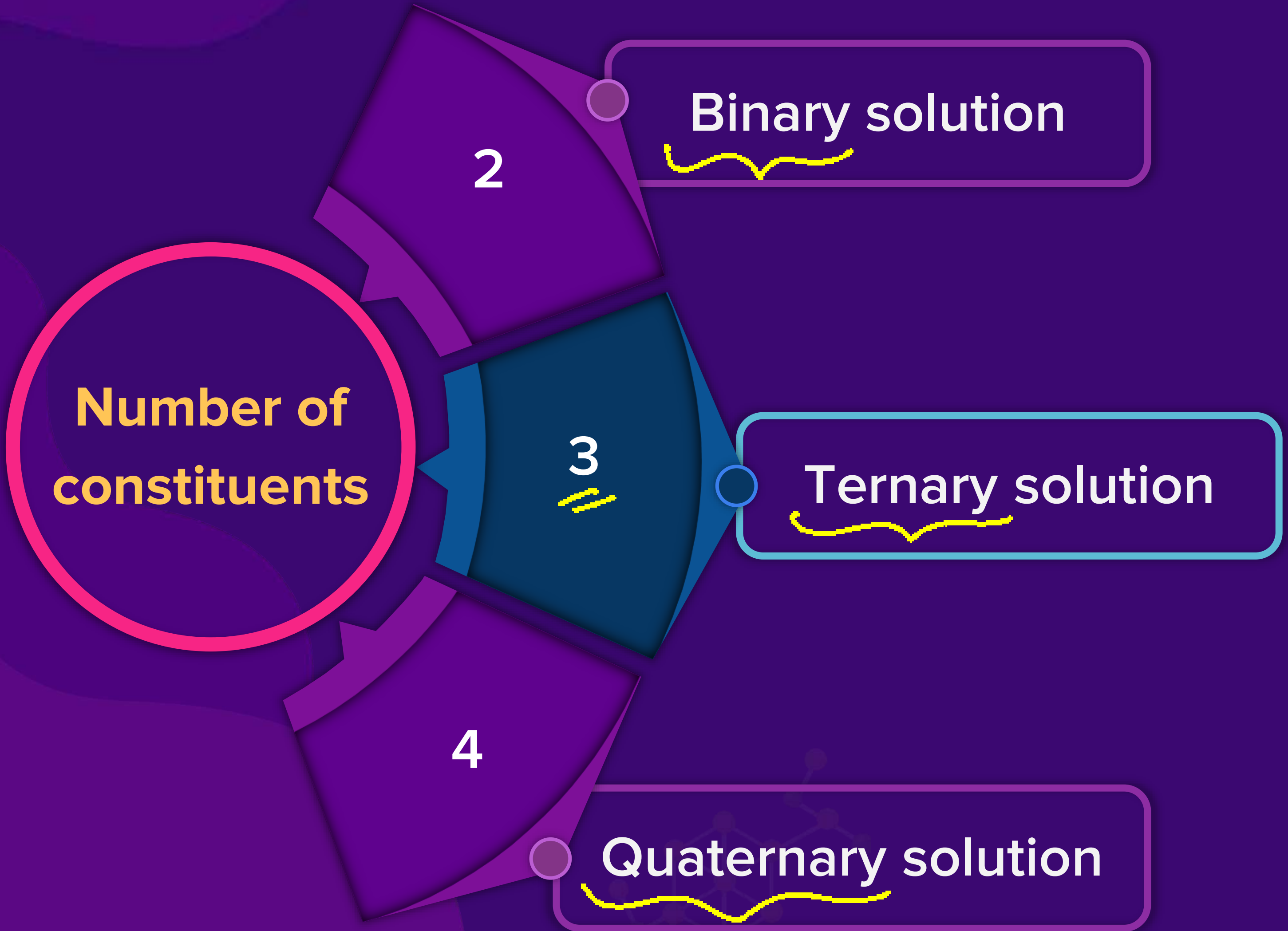
Constituents present in **relatively small** amounts.

Solution

Solvent determines the **physical state** in which the solution exists.

If one of the components of a solution is **water**, it will always be considered as a **solvent** even when it is present in a very less amount.

Solution



Number of constituents

2 → Binary solution

3 → Ternary solution

4 → Quaternary solution

Types of Liquid Solutions

- Liquid solutions
  - Gas in liquid
  - Liquid in liquid
  - Solid in liquid

Types of Liquid Solutions

1) Gas in liquid :

Solvent - Liquid & Solute - Gas
Example: CO_2 dissolved in water.

2) Liquid in liquid :

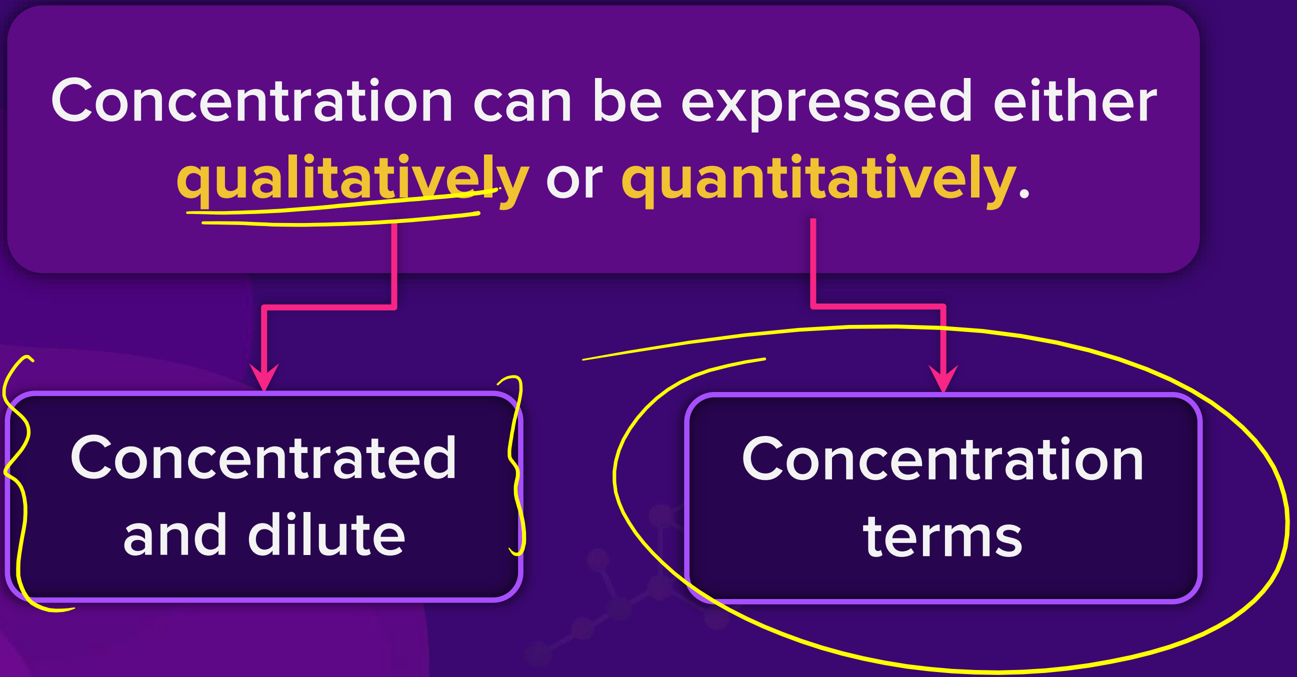
Solvent - Liquid & Solute - liquid
Example: Methanol dissolved in water .

3) Solid in liquid :

Solvent - Liquid & Solute - solid
Example: Sugar in water, common salt dissolved in water.

Concentration Term

Composition of a solution can be described by expressing its **concentration**.



Concentration can be expressed either **qualitatively** or **quantitatively**.

Concentrated and dilute

Concentration terms

Recall

Percentage Concentration Terms

- % w/w
- % w/V
- % V/V

% by mass

or

% by weight

weight / volume

% w/w

Amount of **solute** in grams dissolved **per 100 g** of solution.

$$\% \text{ w/w} = \frac{\text{Weight of solute (g)}}{\text{Weight of solution (g)}} \times 100$$

% w/V

→(in g)

Amount of **solute** in grams dissolved **per 100 mL** of solution.

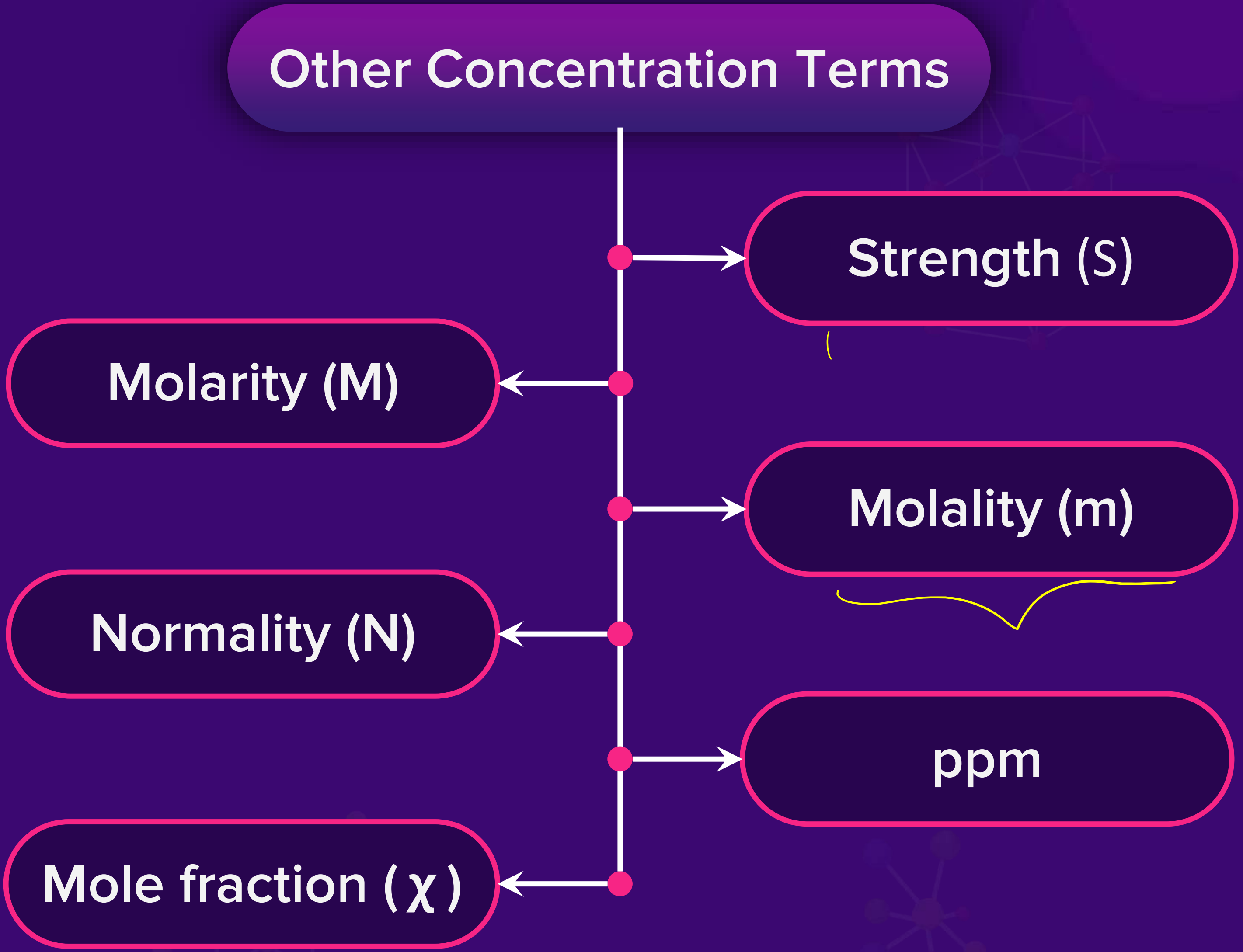
$$\% \text{ w/V} = \frac{\text{Weight of solute (g)}}{\text{Volume of solution (mL)}} \times 100$$

% V/V

Volume of a **solute (in mL)** dissolved **per 100 mL** of solution.

$$\% \text{ V/V} = \frac{\text{Volume of solute (mL)}}{\text{Volume of solution (mL)}} \times 100$$

Other Concentration Term



Other Concentration Terms

- Strength (S)
- Molarity (M)
- Molality (m)
- Normality (N)
- ppm
- Mole fraction (χ)

Strength of Solution $= 10 \times (\% \ w/v)$

$g.L^{-1}$

Weight of **solute** (in gram) per litre (1000 mL) of **solution.**

If we have 6% w/w urea solution with density 1.060 g/mL, then its strength in g/L is:

BOARDS

Solution

Given:

% w/w = 6%

Density of solution = 1.060 g mL^{-1}

Strength of solution = ?

6% w/w urea solution means 6 *g* urea present in 100 g solution

Therefore, mass of the solution= 100 *g*

Density is defined as the ratio of mass and volume. Thus,

$$\text{Density of solution} = \frac{\text{Mass of solution}}{\text{Volume of solution}}$$

$$\text{Volume of solution} = \frac{100 \text{ g}}{1.060 \text{ g mL}^{-1}}$$

Volume of solution = 94.34 mL
or volume of solution = 0.09434 L

Strength of the solution is given as follow:

$$\text{Strength of solution} = \frac{\text{Mass of the solute } (g)}{\text{Volume of solution } (L)}$$

$$\text{Strength of solution} = \frac{6\ g}{0.09434\ L}$$

Hence, strength of solution = 63.6 %

Molarity (M)

Number of **moles** of **solute** per litre of solution.

$$M = \frac{\text{No. of moles of solute (n)}}{\text{Volume of solution (L)}}$$

6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of urea solution is:

BOARDS

Solution

Number of molecules of urea = 6.02×10^{20} molecules

Number of moles of urea $= \dfrac{N}{N_A} = \dfrac{6.02 \times 10^{20}}{6.02 \times 10^{23}}$

Number of moles of urea = 0.001 mol

Volume solution = 100 mL = 0.1 L

$$\textbf{Molarity} = \frac{n}{V} = \frac{0.001}{0.1} = 0.01\ \text{M}$$

Therefore, molarity of urea solution = 0.01 M

Molality (m)

Number of **moles** of the **solute** per **1000 g or 1 kg** of **solvent**.

$$m = \frac{\text{No. of moles of solute (n)}}{\text{Mass of solvent (kg)}}$$

The molality of 1.2 M H_2SO_4 solution is:
(Given density of solution = 1.4 g/mL)

$1.2\,\frac{mol}{L}$

Consider 1L soln

	mass = n×M	M	mol (n)	Volume
solute	1.2×98 g = 117.6 g	98 g/mol	1.2 mol	
solvent	1222.4 g			
solution	$\frac{1.4g}{mL}$ × 1L = 1400 g			1L

$\rho_{soln} \times V_{soln} = mass_{soln}$

∴ $m = \frac{1.2\,mol}{1.2224\,kg} \lesssim 1$

$$molarity = \frac{n_u}{V_{soln}}$$

$$\frac{1.2\,mol}{\cancel{L}} = \frac{n_u}{\cancel{L}}$$

∴ $n_{solute} = 1.2$

= 1222.4

$$\frac{1}{m} = \frac{L}{10^{-3}} = 10^{+3}$$

$$\frac{1.4\,g}{mL} = \frac{1.4\,g}{10^{-3}\,L} = 1400\,\frac{g}{L}$$

Solute + solvent = sln

∴ solvent = soln − solute

= 1400 g − 117.6 g

Normality (N)

Number of **gram equivalents** of **solute** dissolved **per litre** of **solution.**

Normality (N)

$$\text{Normality} = \frac{\left\{\text{Number of equivalents of solute}\right\}}{\text{Volume of solution (L)}}$$

$$\text{Number of equivalents} = \frac{\text{Mass of the species}}{\text{Equivalent mass}}$$

$$= \frac{\text{Mass of the species}}{\dfrac{\text{Molar mass}}{\text{n - factor}}}$$

① no. of equivalents = no. of moles × n

② equivalent mass = $\frac{\text{Molar mass}}{(n)}$

③ Normality = molarity × n

n = equivalane factor

'n' - factor

- For oxidising /reducing agents
- For acid/ base
- For salt

For Oxidising/Reducing Agents

Number of **electrons involved** in oxidation/ reduction half reaction per mole of oxidising/ reducing agent.

$$5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + H_2O$$

n-factor = 5

For Acid/Base

Number of H^+ ions displaced/OH^- ions displaced per mole of acid/base.

Example: H_2SO_4 n-factor = 2

For Salts

For simple salts, n-factor is a total **charge on cations** or a total **charge on anions.**

Example: $Al_2(SO_4)_3$

n-factor = charge on the cation

$= 2 \times 3 = 6$

What will be the n-factor for the following reactions?

$\Delta = \text{change}$

(i) I_2 ⟶ $2IO_3^-$ (n-factor for

zero $\qquad$ +5

1 I atom ; $\Delta = +5$

for I : $n = 5$

for $I_2 = 2 \times 5 = \underline{10}$

$n = 2$

(ii) $C_2O_4^{2-}$ ⟶ $2CO_2$ (n-factor for C)

+3 $\qquad$ +4 $\qquad$ $+2e^-$

(iii) SO_2 ⟶ SO_4^{2-} (n-factor for S)

+4 $\qquad$ +6

$\Delta = 2 = n_f$

Hence, n-factor is 10, 2 and 2 respectively for the given reactions.

Parts per Million (ppm)

The number of **parts of solute** present in **1 million parts of solution.**

Parts per Million (ppm)

$$\text{ppm (w/w)} = \frac{\text{Weight of solute (g)}}{\text{Weight of solution (g)}} \times 10^6$$

$$\text{ppm (w/V)} = \frac{\text{Weight of solute (g)}}{\text{Volume of solution (mL)}} \times 10^6$$

$$\text{ppm (moles/moles)} = \frac{\text{Moles of solute}}{\text{Moles of solution}} \times 10^6$$

Mole Fraction ($\mathcal{X}$)

Ratio of the **number of moles** of a **particular component** to the **total number of moles** of all the components.

Mole Fraction (χ)

For a binary solution,

$$\chi_{\text{solute}} = \frac{(\text{Moles of solute})}{\text{Total moles in solutions}} = \frac{n}{n + N}$$

$$\chi_{\text{solvent}} = \frac{\text{Moles of solvent}}{\text{Total moles in solutions}} = \frac{N}{n + N}$$

$$\chi_{\text{solute}} + \chi_{\text{solvent}} = 1$$

What is the mole fraction of the solute in a 1.00 m aqueous solution?

Solution

1kg solvent → H_2O → 1 mol solute

$1000\,g\ H_2O = \frac{1000}{18} = \text{~~55.55 mol~~}\ H_2O$

$n_{Total} = \frac{1000}{18} + 1 = \frac{1018}{18}$

$x_{solute} = \frac{1\ mol}{\frac{1018}{18}\ mol}$

$= \frac{18}{1018} \lesssim \frac{18}{1000}$

$\frac{18}{1018} \lesssim 0.018$

a ← num.

b ← den.

den ↑

number ↓

den ↓

num ↑

- a 1.770
- b 0.0354
- c $0.01\overset{8}{7}7$
- d 0.177

All volume dependent concentration terms are dependent on **temperature.**

Which of the following is **dependent on temperature**?

BOARDS

$\frac{mol}{L}$

a. Molarity

$\frac{n_u}{n_{soln}}$

b. Mole fraction

$\frac{m_u}{m_{soln}} \times 100$

c. Weight percentage

$\frac{n_u}{m_w}$

d. Molality

Solution

Volume dependent concentration terms are temperature dependent, here, molarity is volume dependent and it also dependent on temperature.

Hence, option (a) is the correct answer.

2.5 litre of 1 M NaOH solution is mixed with another 3 litre of 0.5 M NaOH solution. Then find out molarity of resultant solution.

a 0.80 M

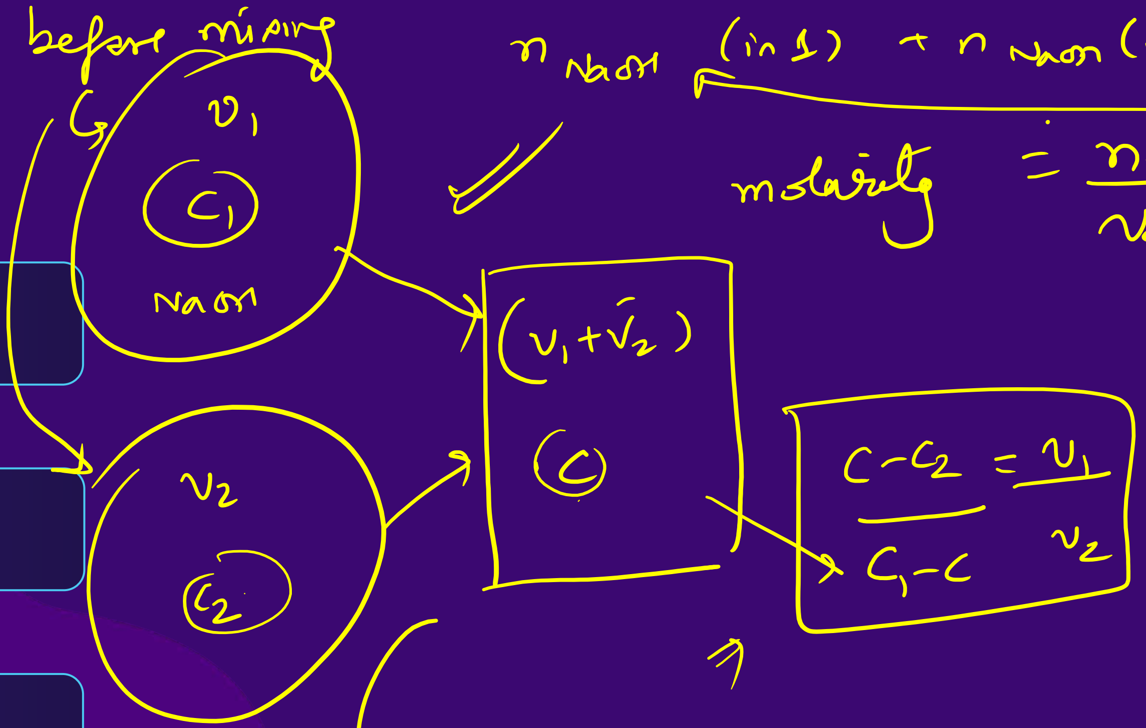
b 1.0 M

c 0.73 M

d 0.50 M

n_{NaOH} (in 1) + n_{NaOH} (in 2) = n_{NaOH} (final, after mixing)

molarity $= \frac{n_{NaOH}}{V_{soln.}}$ $\qquad C_1 = \frac{n'_{NaOH}}{V_1}$



$$\boxed{\frac{C-C_2}{C_1-C} = \frac{V_1}{V_2}}$$

$$C_1V_1 + C_2V_2 = C(V_1+V_2)$$

$$(C_2-C)V_2 = (C-C_1)V_1$$

$$\therefore \frac{C_2-C}{C-C_1} = \frac{V_1}{V_2} = \frac{C-C_2}{C_1-C}$$

Solution

2.5 litre of 1 M NaOH + 3 litre of 0.5 M NaOH = 5.5 litre of ‘x’ M NaOH

$$\frac{x - 0.5}{1 - x} = \frac{V_1}{V_2} = \frac{2.5}{3}$$

On solving we will get , 3x - 1.5 = 2.5 - 2.5x

$$x = \frac{4}{5.5} = \mathbf{0.73\ M}$$

Hence, option (c) is the correct answer.

Mole fraction of glycerine ($C_3H_8O_3$) in a solution of 36 g of water and 46 g of glycerine is:

BOARDS

Solution

Given:
Mass of glycerine = 46 g
Molar mass of glycerine = 92 g mol^{-1}
Mass of water = 36 g
Molar mass of water = 18 g mol^{-1}

$$\text{Moles of glycerine} = \frac{46}{92}$$

$$\text{Moles of water} = \frac{36}{18}$$

Number of moles of glycerine = 0.5 mol

Number of moles of water = 2 mol

Total number of moles = 2 + 0.5 = 2.5 moles

$$\text{Mole fraction of glycerine} = \frac{\text{Moles of glycerine}}{\text{Total moles}} = \frac{0.5}{2.5} = 0.2$$

Hence, mole fraction of glycerine = 0.2

Density of a 2.05 M solution of acetic acid in water is 1.02 g/mL. The molality of the solution is:

BOARDS

(a) 3.28 mol kg^{-1}

(b) 2.28 mol kg^{-1}

(c) 0.44 mol kg^{-1}

(d) 1.14 mol kg^{-1}

Solution

Given:
Density of the solution = 1.02 g mL^{-1}
Let the volume of the solution = 1 L = 1000 mL

$$\text{Density} = \frac{\text{Mass (g)}}{\text{Volume (mL)}}$$

$$1.02 \text{ g mL}^{-1} = \frac{\text{Mass (g)}}{1000 \text{ mL}}$$

Thus, mass of solution = 1020 g

Given:
Molarity of acetic acid = 2.05

$$\textbf{Molarity} = \frac{n}{\text{Volume (L)}} = \frac{\text{Mass of acetic acid}}{\text{Molar mass of acetic acid x V}}$$

$$\textbf{2.05} = \frac{\text{Mass of acetic acid}}{60 \times 1}$$

Thus, mass of acetic acid = 123 g

Mass of solvent = Mass of solution – Mass of solute = 1020 - 123 = 897 g
Mass of solvent = 897 g = 0.897 kg

$$\textbf{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$$

$$\textbf{Molality} = \frac{2.05}{0.897} = 2.28\ \text{m}$$

Therefore, molality of acetic acid solution = 2.28 molal

Hence, option (b) is the correct answer.

The mole fraction of benzene in solution containing 30% by mass in carbon tetrachloride is:

Solution

Let the mass of solution = 100 g
Molar mass of benzene = 78 g mol^{-1}
Molar mass of CCl_4 = 154 g mol^{-1}

30% by mass means that 30 g of solute is present in 100 g of solution.
Therefore, mass of benzene = 30 g
Mass of CCl_4 = 100 – 30 = 70 g

Moles of benzene $= \frac{30}{78} = 0.385$

Moles of CCl_4 $= \frac{70}{154} = 0.454$

Total moles = 0.385 + 0.454 = 0.8395

$$\text{Mole fraction of benzene} = \frac{\text{Moles of benzene}}{\text{Total moles}}$$

$$\text{Mole fraction of benzene} = \frac{0.385}{0.839}$$

Mole fraction of benzene = 0.459

Hence, mole fraction of benzene is 0.459.

Calculate the resultant molarity of following:

✓ a) 200 mL 1 M HCl + 300 mL water

✓ b) 200 mL 1 M HCl + 100 mL 0.5 M H_2SO_4

$1 H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

$0.5M \rightarrow 0.5 \times 2 = 1M$

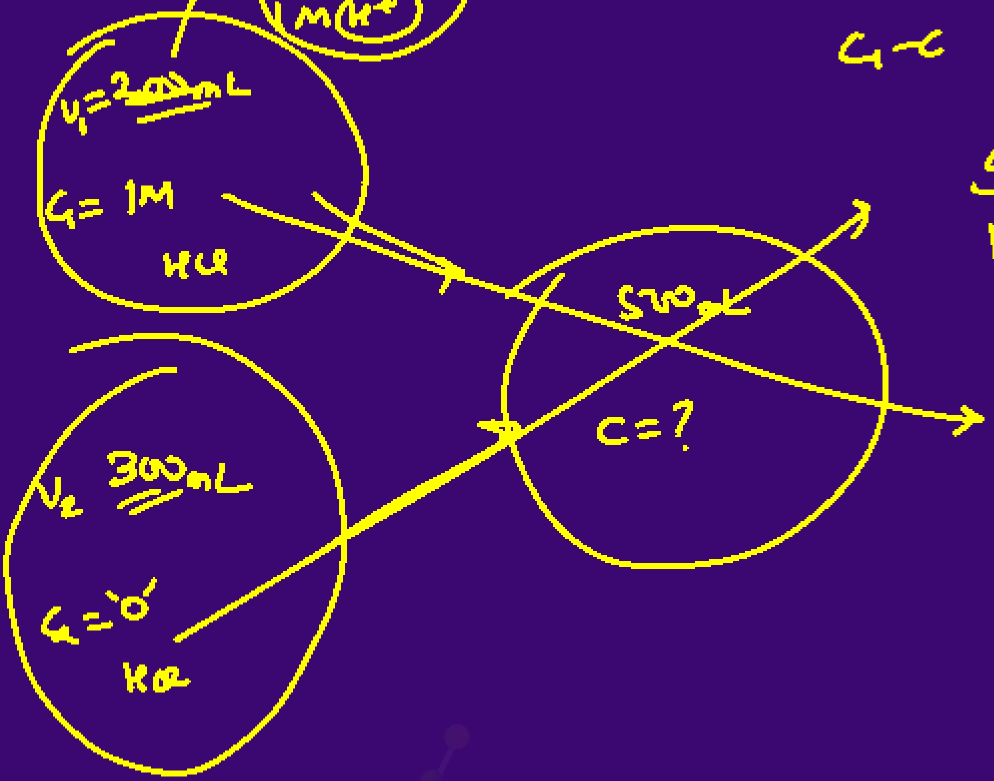
$C_{final} = 0.4M$

$\boxed{1M}$

$1M\ \boxed{H^+}$

$\boxed{1M\ HCl}$

ⓐ



$$\frac{C - C_2}{C_1 - C} = \frac{V_1}{V_2}$$

$$\frac{C}{1 - C} = \frac{200}{300}$$

$$C = \frac{200}{500}$$

$$= 0.4$$

(b) $M_1 = 1$ M

$V_1 = 200$ ml

$M_2 = 0.5$

$V_2 = 100$ ml

$n = 2$

Using the formula

$M_1 V_1 + nM_2 V_2 = \mathbf{MV}$

Where M is the resultant molarity and V is the total volume

$(1 \times 200) + (2 \times 0.5 \times 300) = M \times 500$

$M = 1$ M

Resultant molarity is 1 M.

Volume of water needed to mix with 10 mL, 10 N HNO_3 to get 0.1 N HNO_3 is:

BOARDS NEET

Solution

10 mL
10 N
HNO_3

v mL
0 N
HNO_3

(10+v) mL
0.1 M

$$\frac{0.1-0}{10-0.1} = \frac{10}{v}$$

$$\therefore v = \frac{99}{0.1} = 990\,mL$$

a 1000 mL

[✓ b] 990 mL

c 1010 mL

d 10 mL

Hence, option (b) is the correct answer.

Concentrated aqueous sulphuric acid is 98% H_2SO_4 by mass and has a density of 1.8 g mL^{-1}. Volume of acid required to make one litre of 0.1 M H_2SO_4 solution is:

Solution

pure $m_{H_2SO_4} = 9.8g$

98% H_2SO_4

$\rho = 1.8$ g/mL

$v = ?$

$\xrightarrow{H_2O}$ 1L

0.1M

H_2SO_4

$\cancel{1L} \times 0.1 \frac{mol}{\cancel{L}}$

$= 0.1 mol$

$m_{soln} = \rho.V = 1.8v$ grams

$m_{pure\ H_2SO_4} = m_{soln} \times \frac{98}{100}$

$m_{H_2SO_4}(ini.) = m_{H_2SO_4}(final)$

$= 0.1 mol \quad \therefore m^{ini}_{H_2SO_4} = 0.1 mol \times 98 \frac{g}{mol}$

$= 9.8g$

$\overset{10}{\cancel{9.8}}g = 1.8v \times 0.\cancel{98}$

$v = \frac{10}{1.8}$

$= 5.55$ ml

a 16.65 mL

b 22.20 mL

[✓ c 5.55 mL]

d 11.10 mL

Hence, option (c) is the correct answer.

A solution is prepared by dissolving 24.5 g of sodium hydroxide in distilled water to give 1 L solution. The molarity of NaOH in the solution is:

BOARDS

Solution

$m_{NaOH} = 24.5\,g$; $n_{NaOH} = \frac{24.5\,g}{40\,g/mol} = \frac{24.5}{40}\,mol$

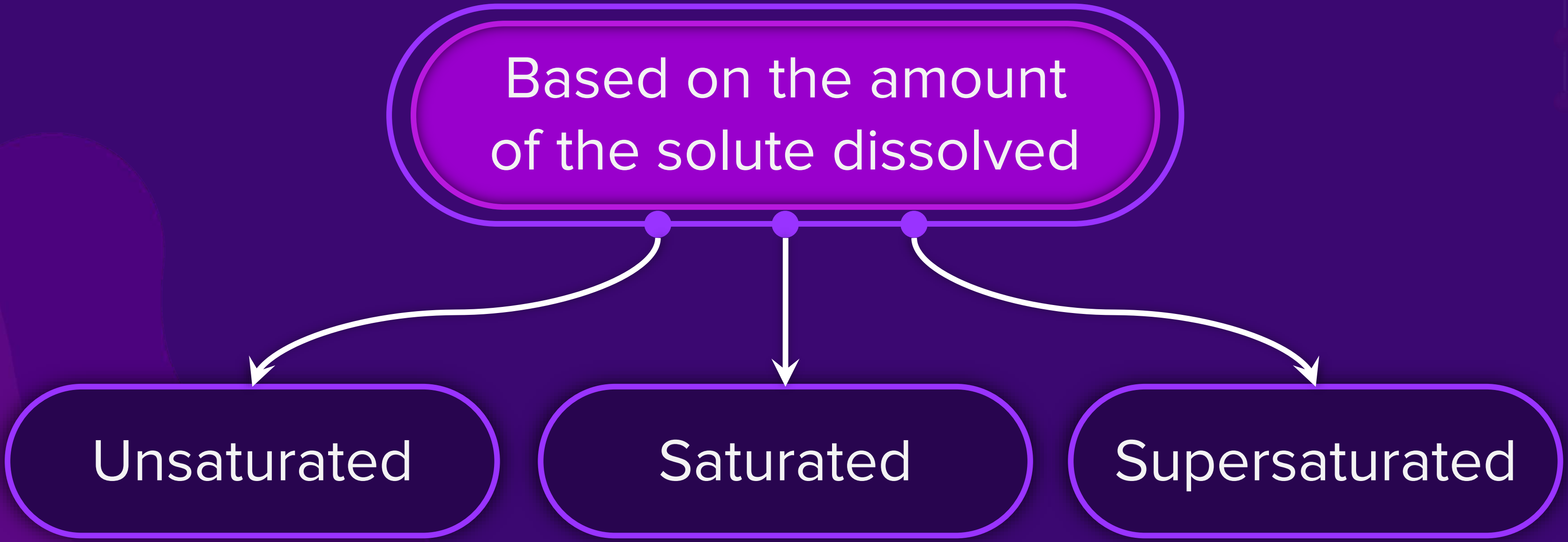
$V_{soln} = 1L$

$$M = \frac{\left(\frac{24.5}{40}\right) mol}{1L}$$

$$= 0.6125\,M$$

- a 0.2450 M
- [✓ b] 0.6125 M
- c 0.9800 M
- d 1.6326 mL

Hence, option (b) is the correct answer.



Based on the amount of the solute dissolved

Unsaturated

Saturated

Supersaturated

Unsaturated Solution

It is a solution in which **more amount** of solute **can be dissolved** at a particular temperature.

Saturated Solution

A solution in which no more solute **can be dissolved** at a particular temperature.

Supersaturated Solution

A solution which contains **more amount** of the solute **than in the saturated solution** at a particular temperature and pressure.

Solubility

Solubility of a substance is its **maximum amount that can be dissolved** in a specified amount of solvent (generally **100 g** of solvent) at a specified temperature to form a **saturated** solution.

Solubility

Solubility of one substance into another depends on

1. **Nature** of the solute and solvent
2. **Temperature**
3. **Pressure**

Factors-Affecting Solubility

There are three major factors that affect the solubility of a gas in a liquid:

- Nature of the solvent and gas
- Temperature
- Pressure

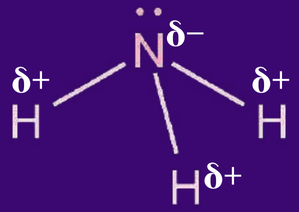
Factors-Affecting Solubility

Nature of the solvent and gas

Like dissolves like

Polar gases dissolve in polar solvents and non-polar gases dissolve in non-polar solvents.

- Polar gases : SO_2, NH_3 etc.
- Polar solvents : H_2O and alcohol etc.
- NH_3 is highly soluble in water, because both are polar and due to H-bonding.



(Ammonia)

Polar

(Water)

Polar

- Non polar gases : O_2 , N_2 and CO_2 etc.
- Non polar solvents : CCl_4, benzene etc.

Factors-Affecting Solubility

Question : O_2 in water is :
(a) Highly soluble
(b) Moderately soluble
(c) Poorly soluble

Answer : O_2 in water is poorly soluble. Because O_2 is non-polar in nature whereas water is a polar solvent.

When a gas undergoes **ionisation** in a solvent, then it is **highly soluble** in that solvent.

E.g. **HCl** is highly soluble in **water.**

- When HCl (g) reacts completely with H_2O and gives H_3O^+ and Cl^- and it becomes highly soluble because equilibrium shifts forward, increasing the solubility of HCl.
- If the solvent reacts with the gas. It can dissolve greater amounts of that gas.

Factors-Affecting Solubility

Temperature

Generally,

Exothermic

Temperature ↑

Solubility ↓

For exothermic reaction as temperature increases solubility of gases in liquid decreases according to Le chatelier's principle.

Consider any gas is dissolved in water.

$$\text{Gas} + \text{aq} \rightleftharpoons \text{Gas (aq)} + \text{Heat}$$

Solubility of gas increases in the forward direction if the temperature is lowered. According to Le chatelier's principle, if the reaction is exothermic in forward direction and temperature of reaction is increased, the reaction shifts to the backward direction.

Factors-Affecting Solubility

Application:

Oxygen dissolves only to a **small extent** in water. It is this dissolved oxygen which sustains all **aquatic life.**

Solubility of gases **increases** with **decrease of temperature**. It is due to this reason that aquatic species are more comfortable in cold water rather than in warm water.

Factors-Affecting Solubility

- Nature of the solvent & gas
- Temperature
- **Pressure**

Henry’s Law

Henry gave a **quantitative relation** between pressure and solubility of a gas in a solvent which is known as **Henry’s law**.

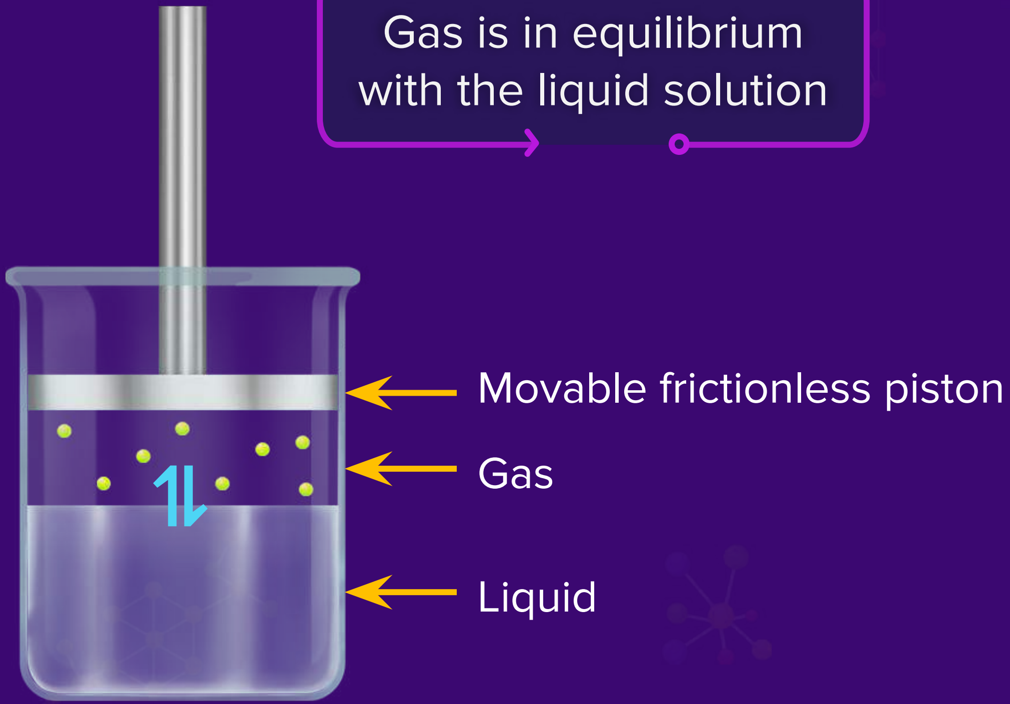
The **solubility of a gas** in a liquid at a given temperature is **directly proportional** to its **partial pressure,** at which, it is dissolved.

Henry's Law

$\chi \propto P$

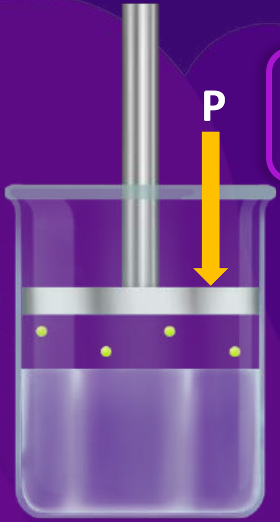
$P = K_H \chi$

- **P** is the **partial pressure** of gas in equilibrium with the solution.
- **K_H** is Henry's law **constant.**
- χ is the **mole fraction** of the unreacted gas in the solution.

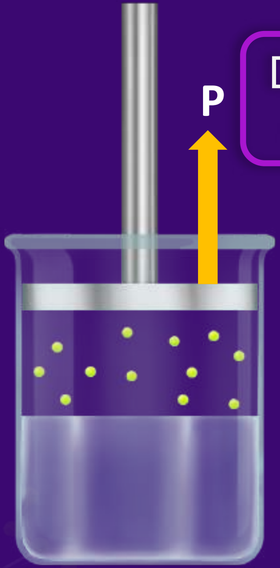


Gas is in equilibrium with the liquid solution

Henry’s Law



Increase in pressure increases solubility

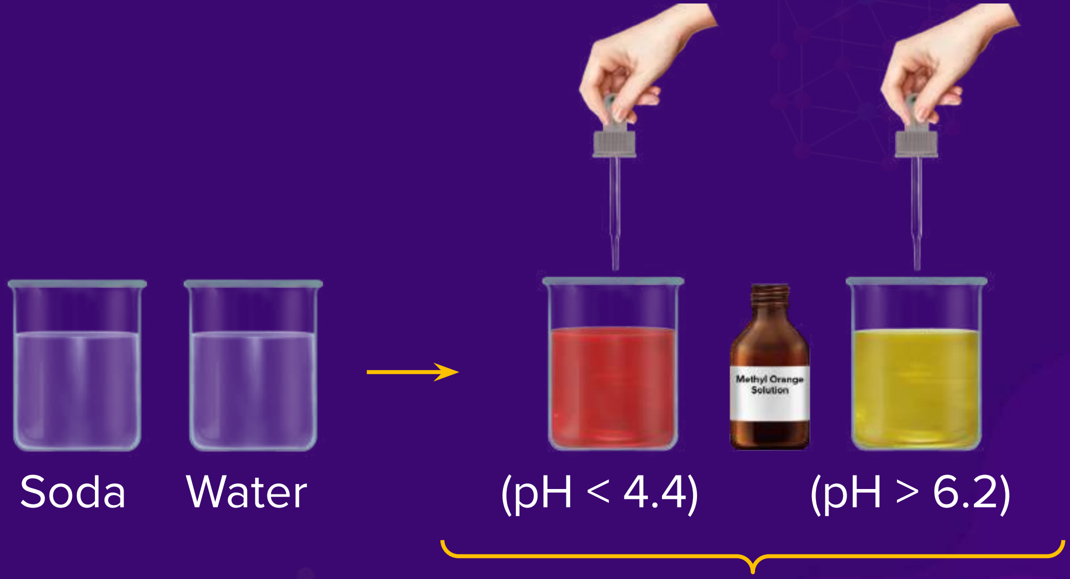


Decrease in pressure decreases solubility

Henry's Law

Example:

Let us take two beakers. The first beaker has soda and the second has water. As we add some drops of methyl red indicator in both the beakers, the first beaker (that has soda) has a red colour (pH < 4.4) and the second beaker has a yellow colour (pH > 6.2).



Addition of methyl red indicator

Henry’s Law

Now, some coloured soda solution is taken in a syringe and its nozzle is closed. If we pull the syringe up, then the colour of the solution changes from red to yellow. This means that the amount of CO_2 dissolved in the soda water decreases. However, when we press the syringe again, the colour of the solution changes back to red. This means that the amount of CO_2 dissolved in the soda water increases.

Soda
Red coloured solution

Water
Yellow coloured solution

Decrease in dissolved $CO_2(g)$ in the soda

Increase in dissolved $CO_2(g)$ in the soda

Characteristics of Henry’s Law Constant

a. K_H has same unit as that of **pressure**: **torr or bar**

b. Different gases have **different K_H** for the **same solvent**.

c. **K_H** value of gas is **different** in **different solvents.**

d. K_H value **increases** with increase in **temperature**.

e. **Higher** the value of K_H of a gas, **lower** will be its **solubility.**

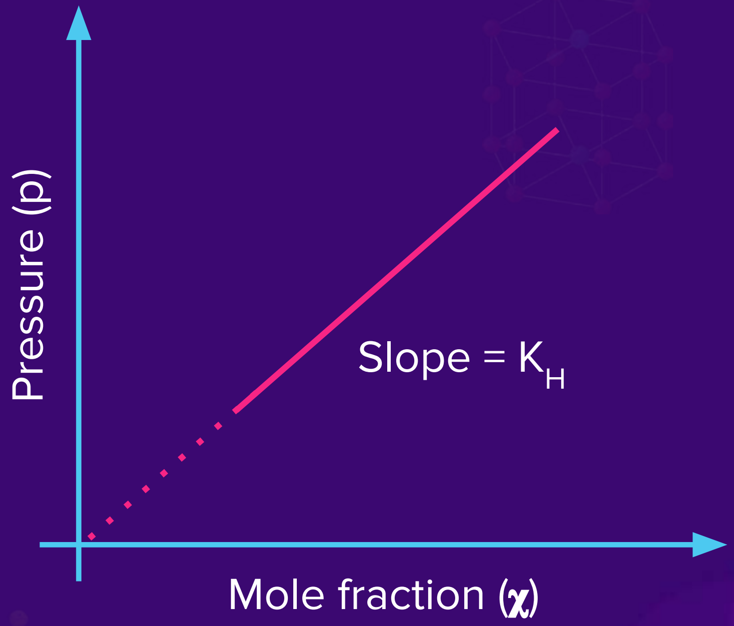
Since, $\chi = \dfrac{P}{K_H}$

Graphical Analysis (Henry's Isotherm)

The Henry's law expression is given as follows:

$$p = K_H \times \chi$$

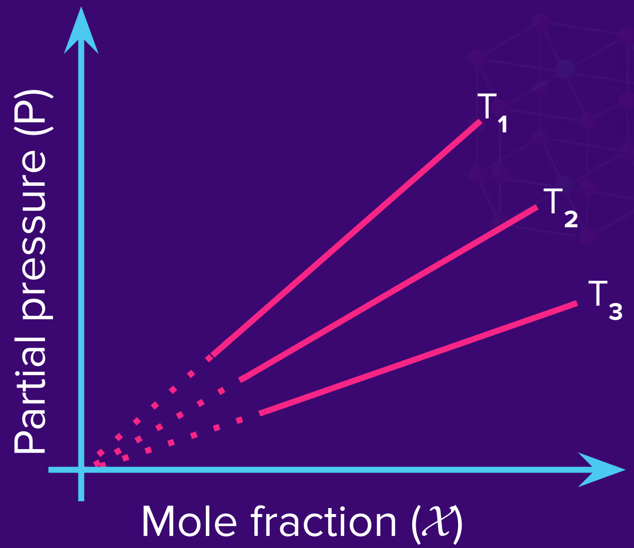
By comparing it with the straight line equation;
($y = mx + c$), we get,
Intercept (c) = 0
Slope (m) = K_H



Plot of P vs χ is a straight line passing through the origin with **slope equal to K_H**.

Graphical Analysis (Henry’s Isotherm)

- Let us assume that $T_1 > T_2 > T_3$, temperature T_1 has the highest slope, and T_3 has the lowest. Thus, for the fixed value of the partial pressure, as the temperature increases, the slope of the graph also increases.
- However, to maintain the fixed value of the partial pressure of any gas, the mole fraction of the gas in the solution decreases. This indicates that the solubility of the gas also decreases as the temperature increases. This occurs only in case of exothermic reactions.



$T_1 > T_2 > T_3$

$K_{H_1} > K_{H_2} > K_{H_3}$

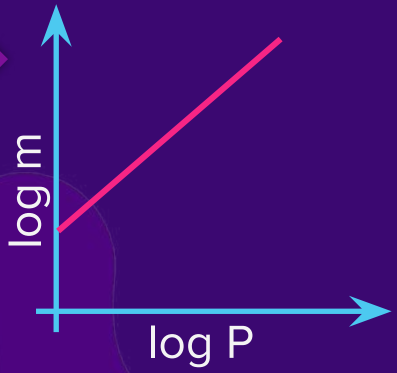
Graphical Analysis (Henry’s Isotherm)

If a mixture of gases is brought in contact with a solvent, each constituent gas dissolves in **proportion to its partial pressure**.

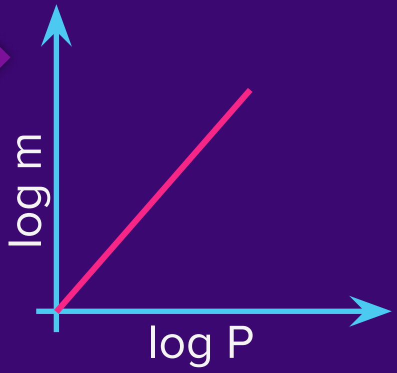
Henry's law applies to each gas **independent** of the pressure of other gas.

Which of the following curves represents the Henry's law?

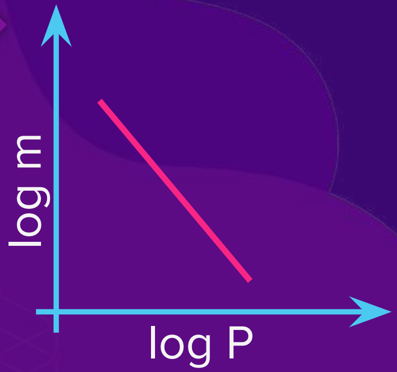
a)



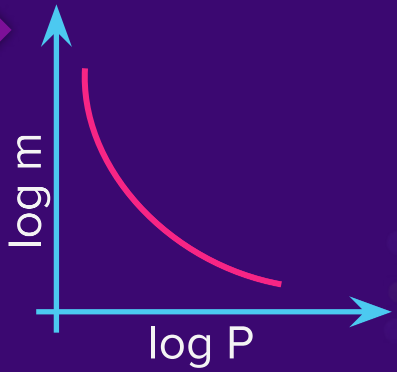
b)



c)



d)



Solution

If concentration of solution is expressed in terms of mass of gas dissolved than **Henry's law** states that **"Mass of the gas dissolved in a given volume of a liquid at a given temperature is proportional to the partial pressure of the gas in equilibrium with the liquid, provided the gas does not undergo any chemical change during the formation of solution."**

$$m = K_H P$$

Taking log on both sides;

$$\log m = \log (K_H P)$$

$$\log m = \log K_H + \log P$$

By comparing it with the straight line equation;

(y = mx + c), we get,

Intercept (c) = $\log K_H$

Hence, option (a) is the correct answer.

Applications of Henry’s Law

- To increase the solubility of CO_2 in soft drinks and soda water, the bottle is sealed under high pressure. Therefore, when the bottle is opened, the pressure decreases, so the solubility of CO_2 also decreases. As a result, the CO_2 escapes from the bottle.

Bottle is sealed under high pressure

Fizzes as CO_2 comes out

Variation of solubility of gas with the pressure

Application of Henry’s Law

- At high altitudes, the partial pressure of oxygen is less than that at the ground level because the population density of the oxygen particles is less at the higher altitudes as compared to that of on the ground level. At heights low blood oxygen level causes anoxia.

Applications of Henry’s Law

- **Scuba diving tanks** are diluted with helium.

Scuba divers must cope with **high concentration** of dissolved gases while, breathing air at high pressure under water.

Increased pressure, increases solubility of atmosphere gases in the blood which are released when the diver comes towards the surface; as the pressure decreases and results in **formation of nitrogen bubbles** in blood.

The bubble **blocks the capillaries** and creates a medical condition known as **bends** that are painful and dangerous to life.

To avoid bends as well as toxic effects of high concentration of nitrogen in blood, tanks used by scuba divers are filled with air **diluted with helium**.

Limitations of Henry’s Law

Henry’s law is **valid** only under the following conditions:

1. **Pressure of the gas is not too high.**
2. **Temperature is not too low.**
3. The gas should **not undergo** any **chemical reaction** with the solvent.
4. The gas should **not undergo dissociation** in solution.

8×10^{-x} moles of gas A is dissolved in 36 *mL* of water when pressure of gaseous mixture above water is 4 *atm*. Mole percentage of gas A in mixture is 25. Henry's constant for gas A in water is 2.5×10^3 *atm*. Find x.

BOARDS

Solution

$n_A = 8 \times 10^{-X}$, $V_w = 36\ mL$, $P = 4\ atm$, $\chi(A) = 0.25$

$p(A) = \chi(A) \times P = 0.25 \times 4 = 1\ atm$

According to Henry's Law:

$p(A) = K_H \times \chi(A)$

$K_H = 2.5 \times 10^3\ atm$

$p(A) = K_H \times \chi(A)$

$$1 = 2.5 \times 10^3 \times \frac{n_A}{n_A + n_{H_2O}}$$

Mass of water = V_w × density of water

$$= 36\ mL \times 1\ g\ mL^{-1} = 36\ g$$

$$n_{H_2O} = \frac{36}{18} = 2 \text{ moles}$$

$$1 = 2.5 \times 10^3 \times \frac{n_A}{n_A + n_{H_2O}}$$

$$1 = 2.5 \times 10^3 \times \frac{n_A}{n_A + 2} \Rightarrow \frac{n_A}{n_A + 2} = 4 \times 10^{-4}$$

$$\frac{n_A}{2} = 4 \times 10^{-4},\ n_A = 8 \times 10^{-4} = 8 \times 10^{-x}$$

$x = 4$

Evaporation and Condensation

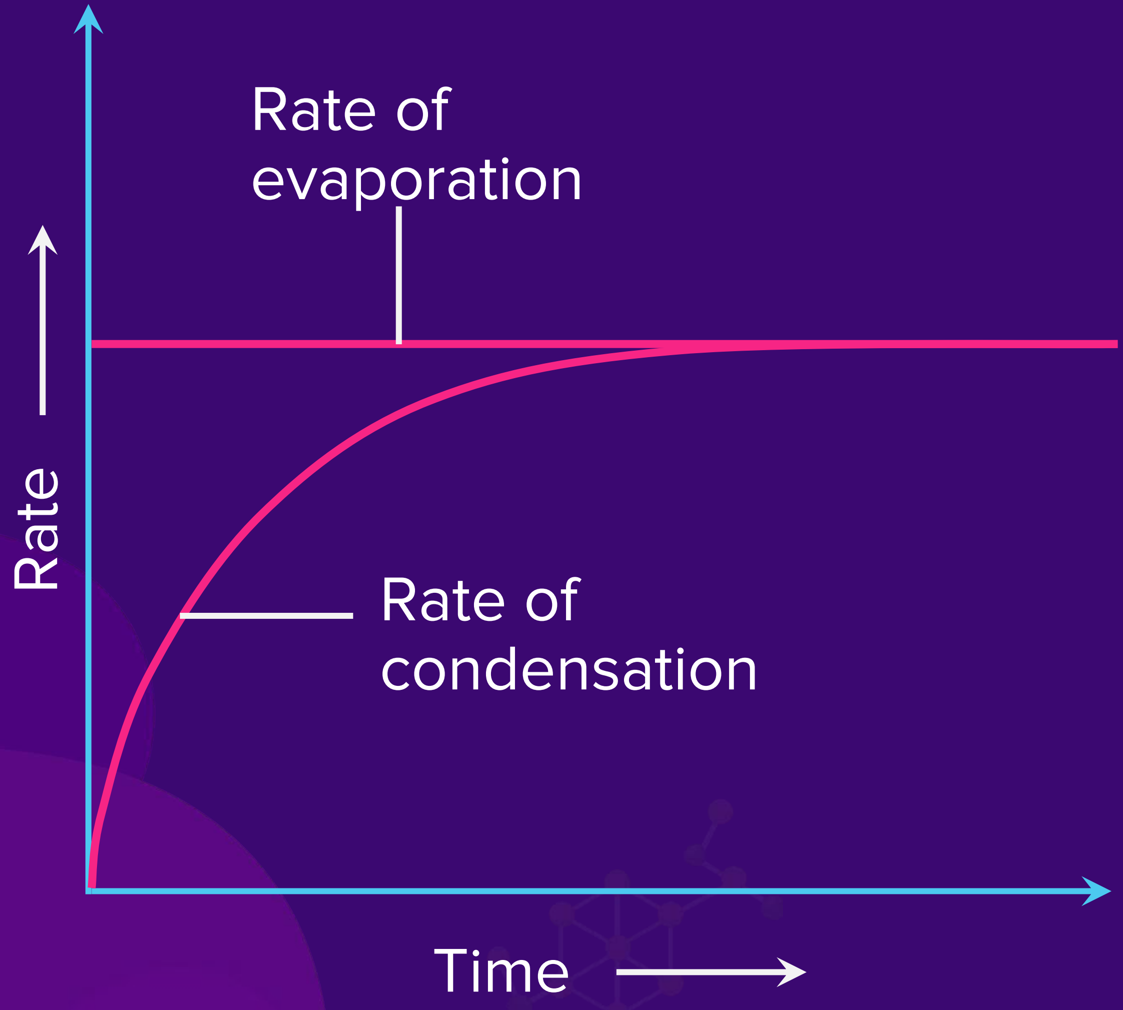
- Some of the more **energetic particles** on the **surface** of the liquid move fast enough to **escape** from the attractive forces holding the liquid together.
- As the gaseous particles bounce around, some of them will **hit the surface** of the liquid again, and be **trapped** there.
- At this stage, partial pressure of vapours is called vapour pressure of the liquid.

There will rapidly be an **equilibrium set up** in which, the number of particles leaving the surface is **exactly balanced** by the number rejoining it.

At equilibrium

Rate of evaporation = Rate of condensation

Evaporation and Condensation



Vapour Pressure (V.P.)

Pressure exerted by the vapours over the liquid surface **at equilibrium**.

Partial pressure of vapours which are in equilibrium with their pure solid or liquid is called vapour pressure of that substance.

$$H_2O(l) \rightleftharpoons H_2O(g)$$

$$K_p = P_{eq}(H_2O\ (g))$$

$$K_p = \text{V.P.}$$

The solubility of $N_2(g)$ in water exposed to the atmosphere, when its partial pressure is 593 *mm Hg* is 5.3×10^{-4} *M*. Its solubility at 760 *mm Hg* and at the same temperature is:

BOARDS

(a) 4.1×10^{-4} M **(b)** 6.8×10^{-4} M **(c)** 1500 M **(d)** 2400 M

Solution

$p_1 = 593$ *mm of Hg* , $s_1 = 5.3 \times 10^{-4}$ *M*, $p_2 = 760$ *mm of Hg* , $s_2 = ?$

$p = K_H \times \chi$, for dilute solutions p is directly proportional to $K_H \times M$

$$\frac{p_1}{p_2} = \frac{s_1}{s_2}$$

$$s_2 = \frac{p_2}{p_1} \times s_1 \Rightarrow \frac{760}{593} \times 5.3 \times 10^{-4} = \mathbf{6.79 \times 10^{-4}}$$

Hence, option (b) is the correct answer.

Henry's law constant for O_2 dissolved in water is 4.34×10^4 *atm* at 25 °C. If the partial pressure of oxygen in air is 0.4 *atm*. The concentration (M) of dissolved oxygen in water in equilibrium with air at 25 °C is 5×10^{-P}. What is the value of P?

BOARDS

(a) 5 **(b)** 6 **(c)** 4 **(d)** 8

Solution

$$K_{H(O_2)} = 4.34 \times 10^4 \, atm$$

$$p_{(O_2)} = 0.4 \, atm$$

According to Henry's law:

$$p_{(O_2)} = K_{H(O_2)} \times \chi_{(O_2)}$$

$$0.4 \, atm = 4.34 \times 10^4 \, atm \times \chi_{(O_2)}$$

$$\Rightarrow \chi_{(O_2)} = 0.092 \times 10^{-4}$$

Consider total one moles of solution,

$n_{O_2} + n_{H_2O} = 1$

$n_{H_2O} = 1 - 0.092 \times 10^{-4} \approx 1\ mol$

Volume of solution will be almost equal to volume of 1 mole of water i.e., 18 *mL*.

$$\text{Molarity of } O_2 = \frac{\text{Moles of } O_2}{\text{Volumes of solution in } L}$$

$$\text{Molarity of } O_2 = \frac{0.092 \times 10^{-4}}{18 \times 10^{-3}}$$

$$\text{Molarity of } O_2 = 5.11 \times 10^{-4}$$

Hence, value of P = 4.

Hence, options (c) is the correct answer.

According to Henry’s law, the solubility of a gas in a given volume of liquid increases with increase in:

BOARDS

(a) Temperature
(b) Pressure
(c) Both (a) and (b)
(d) None of the above

Solution

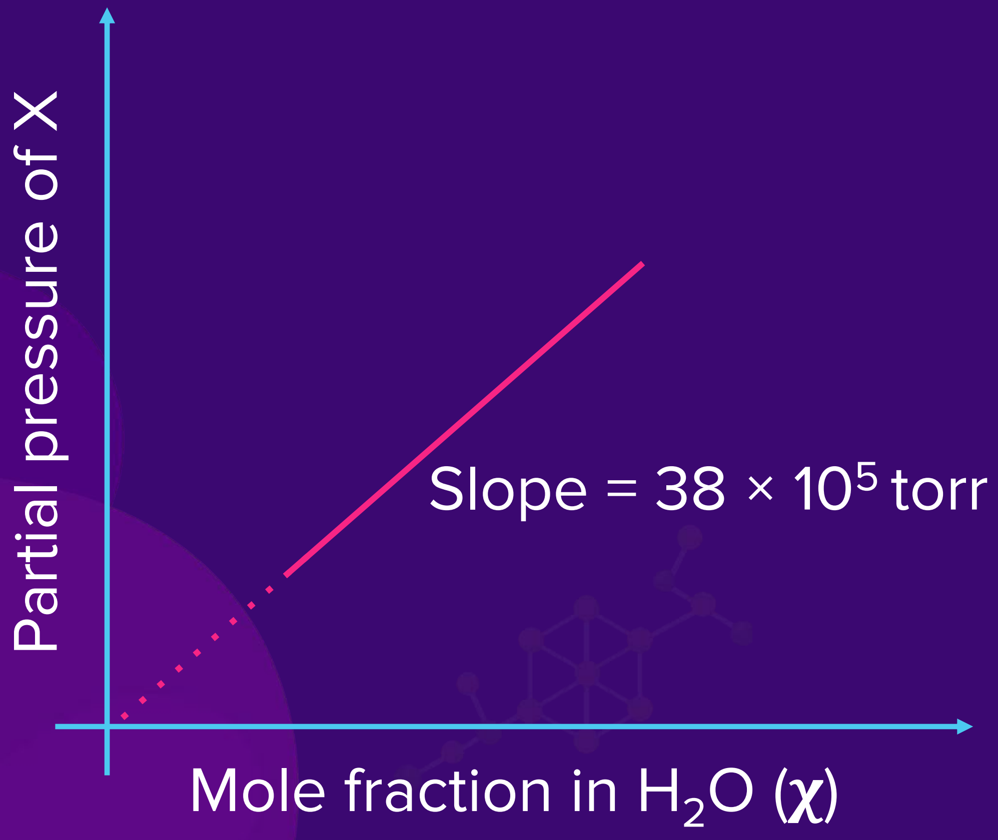
Henry’s law states that, “The **solubility of a gas** in a liquid at a given temperature is **directly proportional** to its **partial pressure,** at which, it is dissolved.”

Hence, option (b) is the correct answer.

A gas 'X' is present with saturated water vapour over water liquid at a total pressure of 1.5 *atm*. Vapour pressure of H_2O at the same temperature is 0.5 *atm*. The solubility of gas 'X' in terms of moles in 10 moles $H_2O(l)$ is given by $P \times 10^{-3}$. What is the value of P?

BOARDS

(a) 2 **(b)** 5 **(c)** 4 **(d)** 3



Solution

$p^{o}_{(X)} + p^{o}_{H_2O} = 1.5\ atm$

$p^{o}_{(X)} + 0.5\ atm = 1.5\ atm$

$p^{o}_{(X)} = 1.5 - 0.5 = 1\ atm = 760\ torr$

According to Henry's Law:

$p^{o}_{(X)} = K_H \times \chi_{(X)}$

$K_H = 38 \times 10^5$ torr (from graph)

$$\chi_{(X)} = \frac{760}{38 \times 10^5} = 2 \times 10^{-4}$$

1 *mol* of solution has 2×10^{-4} *mol* of X

X is almost negligible hence 1 *mol* of water is only present.

That means 1 *mol* of water has 2×10^{-4} *mol* of X

10 moles of water will have 2×10^{-3} *mol* of X

Hence, the value of P is 2.

Hence, option (a) is the correct answer.

The solubility of gases in liquids:

BOARDS

(a) Increases with increase in pressure and temperature
(b) Decreases with increase in pressure and temperature
(c) Increases with increase in pressure and decrease in increase in temperature
(d) Decreases with increase in pressure and increase in temperature

Solution

According to Henry’s law, the solubility of gases in liquids increases with increase in pressure while solubility decreases with increase in temperature.

Hence, option (c) is the correct answer.

Solubility of one substance into another depends on:

BOARDS

a Nature of the solvent & solute

b Temperature

c Pressure

d All of these

Solution

Solubility of one substance into another depends on the nature of the solvent and solute, temperature and pressure.

Hence, option (d) is the correct answer.

Which one of the following statements regarding Henry's law is **not correct**?

NEET BOARDS

(a) Different gases have different K_H (Henry’s law constant) values at the same temperature.

(b) The value of K_H increases with increase of temperature and K_H is function of the nature of the gas.

(c) The partial pressure of the gas in vapour phase is proportional to the mole fraction of the gas in the solution.

(d) Higher the value of K_H at a given pressure, higher is the solubility of the gas in the liquids.

Solution

(a) Different gases have different K_H (Henry’s law constant) values at the same temperature. This is true statement.

(b) The value of K_H increases with increase of temperature and K_H is a function of the nature of the gas. This is also a true statement as temperature increases, solubility decreases so, K_H increases.

(c) The partial pressure of the gas in the vapour phase is proportional to the mole fraction of the gas in the solution. This is also true as this is Henry's law.

(d) Higher the value of K_H at a given pressure, higher is the solubility of the gas in the liquids. This statement is not true as at higher values of K_H the solubility will be less.

Hence, option (d) is correct answer.

If the solubility of H_2S gas in water at STP is 0.2 m, then Henry's law constant for H_2S in water at 273 K is:

1atm

BOARDS

a 3.6 × 10^8 Pa

b 5 × 10^8 Pa

c 5 × 10^5 Pa

d 2.77 × 10^7 Pa

$\frac{1}{.3333} = 3$

$\frac{1}{.36} < 3$

0.2 molal

1kg H_2O → 0.2 mol H_2S

1000 g H_2O

$\frac{1000}{18}$ mol H_2O → 0.2 mol H_2S

$x_{H_2S} = \frac{0.2 \text{ mol}}{\frac{1000}{18} \text{ mol}}$

$= 3.6 \times 10^{-3}$

$p_{H_2S} = K_H \cdot x_{H_2S}$

$K_H = \frac{1 \text{atm}}{3.6 \times 10^{-3}}$

$= \frac{10^3}{3.6} = \frac{10^2}{0.36}$ atm

$\frac{10^2}{.36} \times 1.013 \times 10^5$ Pa

$\frac{1}{.36} \times 10^7$ = 2.77 × 10^7

Hence, option (d) is correct answer.

Features of Vapour Pressure (V.P.)

(a) Vapour pressure of a liquid **does not** depend on

1. **Amount** of liquid taken
2. **Surface area** of the liquid
3. **Volume** or shape of the container

Features of Vapour Pressure (V.P.)

(b) **Nature** of the liquid

Intermolecular attractive forces ↓ → Vapour pressure ↑

Loosely held molecules escape more easily into the vapour phase.

Example : In O_2 gas, the bond is nonpolar which means the attraction is very low and in H_2O, the bond is polar and there will be more attraction between H and O in H_2O. So at same temperature the vapour pressure of O_2 is high because of less attractive force.

Features of Vapour Pressure

Why?

Vapour pressure ↑

Boiling point ↓

As vapour pressure increases, the tendency to go in vapour state increases and hence boiling point decreases.

Features of Vapour Pressure

Boiling point

- Temperature at which the **vapour pressure** of a liquid is **equal** to the **external pressure**.
- On increasing the temperature, liquid starts boiling and there comes a point where on heating, temperature remains constant and heat absorbed, which is called latent heat of vapourisation. After crossing that point liquid converts into vapour. The maximum temperature of liquid is boiling point.
- Boiling point is dependent on pressure.
- On increasing the external pressure, boiling point increases.
- In a pressure cooker the pressure increases and hence boiling point of water increases. For instance, now the boiling point of water is around 110 °C. Now rice has the opportunity to stay at that high temperature and hence they will cook faster as compared to the open container.
- Temperature at which the vapour pressure of a liquid is equal to the external pressure; $P_{(liq)} = P_{ext}$, T = Boiling point
- At normal boiling point, the vapour pressure of the pure liquid = 1 *atm*

Features of Vapour Pressure

(c) **Temperature**

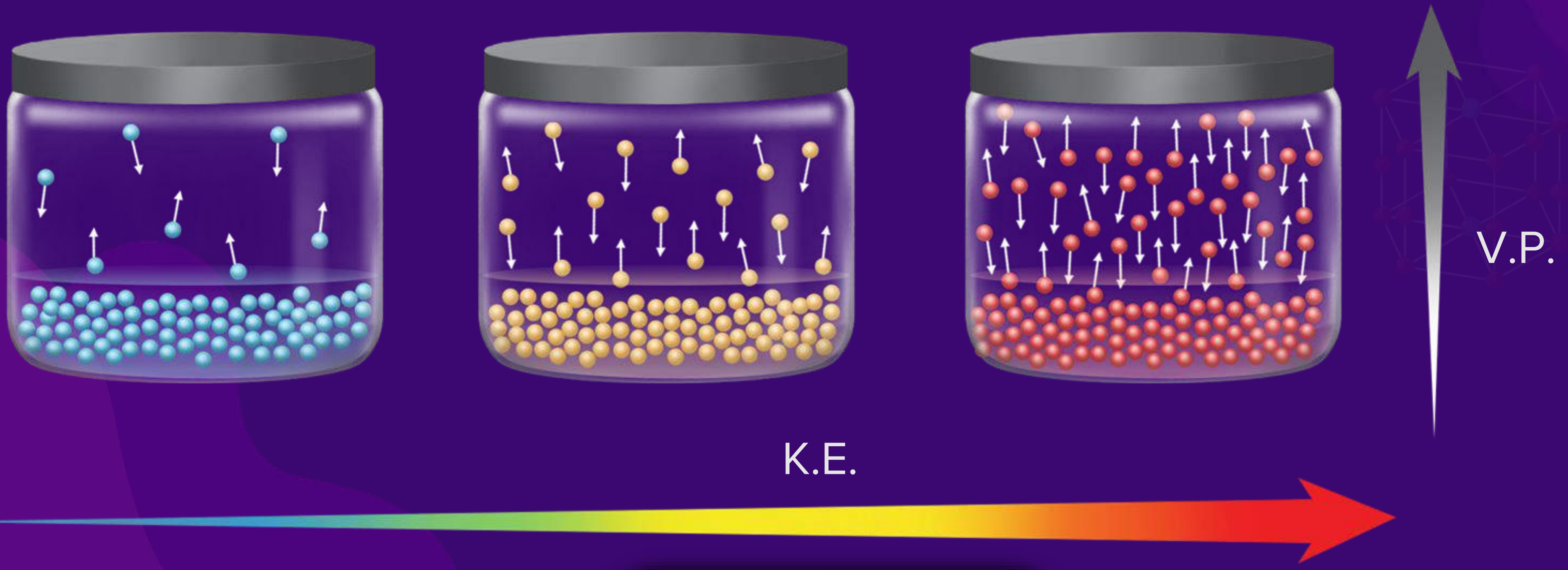
Temperature ↑ Vapour pressure ↑

More molecules from the liquid have **enough K.E.** to escape from the surface of the liquid.

$H_2O\,(l) \rightleftharpoons H_2O\,(g)$ $\quad \Delta H > 0$

According to **Le Chatelier's principle**, increasing the temperature of a system in a dynamic equilibrium, favours the **endothermic** change.

Features of Vapour Pressure



Temperature ↑

K.E. of particles ↑

Raoult’s Law for a Mixture of Two Volatile Miscible Liquids

Raoult's Law

In the solution of volatile liquids, the partial vapour pressure of each component is directly proportional to its **mole fraction**.

Raoult's Law

$P_A \propto x_A$ $\quad$ $P_A = x_A P_A^0$

P_A Partial vapour pressure of component 'A'

x_A Mole fraction of component 'A' in solution

P_A^0 Vapour pressure of pure component 'A' at a given temperature

Raoult's Law

In the solution of volatile liquids, the partial vapour pressure of each component is directly proportional to its **mole fraction**.

$P_A \propto x_A$ $\qquad$ $P_A = x_A P_A^0$

P_A	Partial vapour pressure of component 'A'
x_A	Mole fraction of component 'A' in solution
P_A^0	Vapour pressure of pure component 'A' at a given temperature

Raoult's Law For Binary Solutions of A & B

$$P_A = x_A P_A^0 \qquad P_B = x_B P_B^0$$

$$P_T = P_A + P_B$$

$$P_T = x_A P_A^0 + x_B P_B^0$$

P_T in Terms of Composition of Vapour Phase

y_A	Mole fraction of A in vapour phase above the solution
y_B	Mole fraction of B in vapour phase above the solution

P_T in Terms of Composition of Vapour Phase

$P_A = y_A P_T$ Dalton's Law

$P_A = x_A P_A^0$ Raoult's Law

Similarly,

$P_B = y_B P_T = x_B P_B^0$

$x_A + x_B = 1$

$$\frac{y_A P_T}{P_A^0} + \frac{y_B P_T}{P_B^0} = 1$$

$$\Rightarrow \frac{1}{P_T} = \frac{y_A}{P_A^0} + \frac{y_B}{P_B^0}$$

P_T in Terms of Composition of Vapour Phase

Solutions which obey **Raoult's law** over the entire range of concentration are known as **ideal solutions.**

Graphical Representation of Raoult's Law

$x_A + x_B = 1$

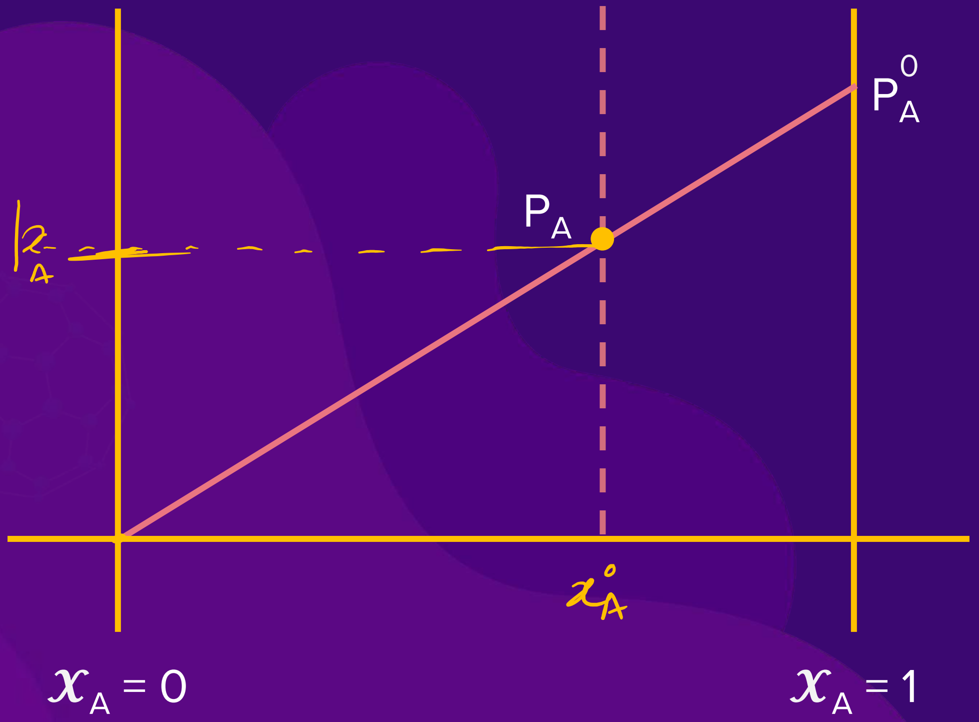
Relation between **total pressure vs mole fraction** in liquid phase.

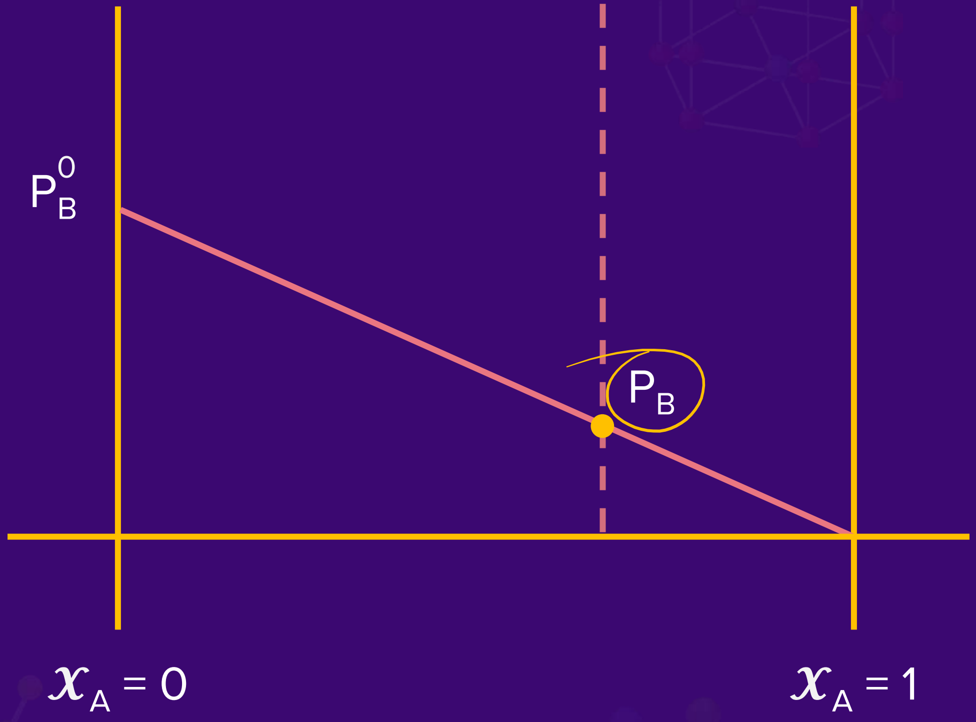
$P_T = (P_A^0 - P_B^0)\,x_A + P_B^0$

$P_T = x_A P_A^0 + x_B P_B^0$

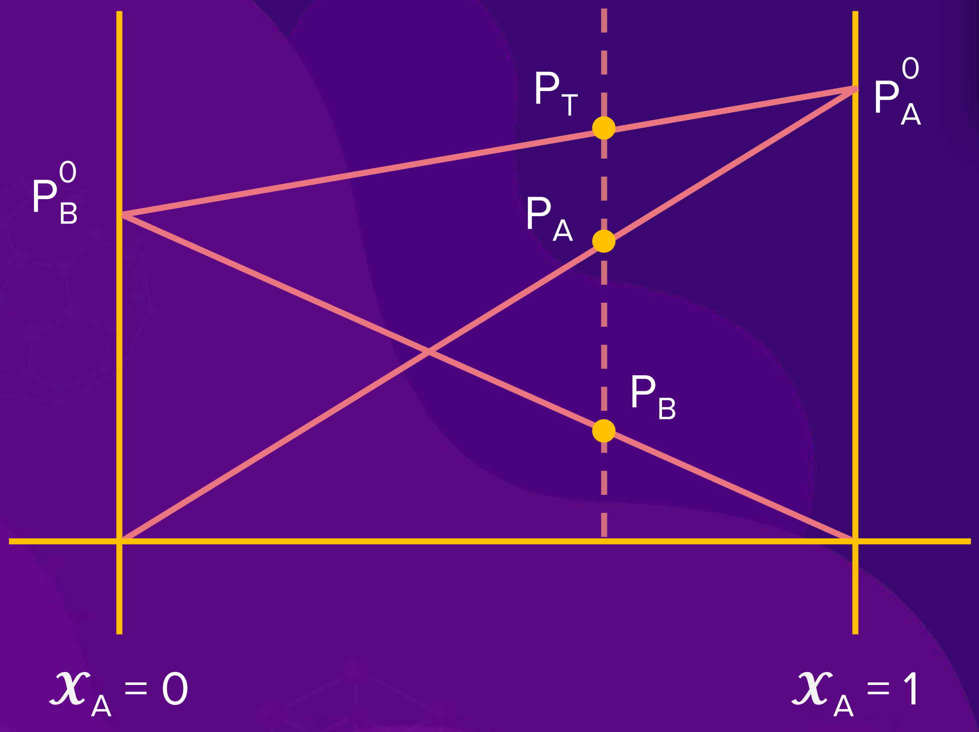
This represents equation of a **straight line** of P_T vs x_A

Graphical Representation of Raoult's Law





Graphical Representation of Raoult's Law



If $P_A^0 > P_B^0$

A is more volatile than B

B.P. of A < B.P. of B

Graphical Representation of Raoult's Law

Did You Know?

Compositions of the liquid and vapour that are in **mutual equilibrium** are [redacted] **the same**.

The vapour will be **richer** in the **more volatile** component, if the mole fraction of components in liquid phase is comparable.

A solution has a 1 : 4 mole ratio of pentane to hexane. The vapour pressures of the pure hydrocarbons at 20 *degree C* are 440 *mm Hg* for pentane and 120 *mm Hg* for hexane. The mole fraction of pentane in the vapour phase would be :

Solution

$\left(\frac{n_p}{n_h}\right)_{liq} = \frac{1}{4} \Rightarrow \boxed{\frac{x_p}{x_h} = \frac{1}{4}}$

$P_p^\circ = 440$ mmHg

$P_h^\circ = 120$ mmHg

$y_p = ?$

$\frac{y_p \cancel{P} =}{y_n \cancel{P} =} \left(\frac{x_p}{x_h}\right)\frac{P_p^\circ}{P_h^\circ}$

$\frac{y_p}{y_h} = \left(\frac{1}{4}\right)\frac{44\cancel{0}}{12\cancel{0}} = \frac{11}{12}$

$\frac{y_p}{y_p+y_n} = \frac{11}{11+12} = \frac{11}{23}$

$1 = \boxed{y_p+y_n}$

$\therefore y_p = \frac{11}{\boxed{23}} \lesssim 0.5$

$\boxed{\frac{11}{22}} = \frac{1}{2} = 0.5$

$\frac{11}{23}$

- a 0.200
- b 0.549
- c 0.786
- [✓ d 0.478]

Hence, option (d) is the correct answer.

The vapour pressure of two liquids P and Q are 80 and 60 *torr*, respectively. The total vapour pressure of solution obtained by mixing 3 moles of P and 2 moles of Q would be :

Solution

$P_P^\circ = 80$ torr
$P_Q^\circ = 60$,,

$P = ?$
$n_P = 3$
$n_Q = 2$

$x_P = \frac{3}{3+2} = 0.6$
$x_Q = \frac{2}{2+3} = 0.4$

$P = x_P P_P^\circ + x_Q P_Q^\circ$
$= 0.6 \times 80 + 0.4 \times 60$
$= 48 + 24 = 72$

- **a** 72 torr
- **b** 140 torr
- **c** 68 torr
- **d** 20 torr

Hence, option (a) is the correct answer.

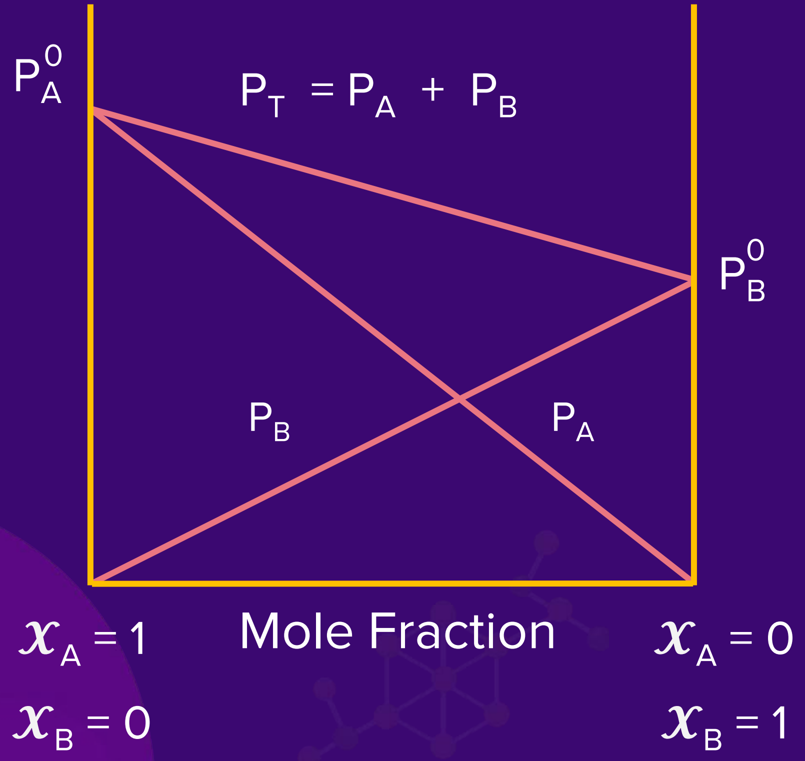
Ideal Solution

Solutions that **obey Raoult's law** over the entire range of concentration

If the forces of attraction between **A—A, B—B is similar to A—B**, then A and B will form ideal solution.

Characteristics of Ideal Solution

(i) Raoult's law is **obeyed**.



Characteristics of Ideal Solution

(ii)

$\Delta_{mix}H = 0$, i.e., there should not be an enthalpy change when components of ideal solutions are mixed.

(iii)

$\Delta_{mix}V = 0$, (1L + 1L = 2L) i.e., there should not be a change in volume on mixing.

Characteristics of Ideal Solution

ideal sol^n

$\Delta P_{mix} = 0$ $\Delta H_{mix} = 0$

$\Delta V_{mix} = 0$

(iv) $(\Delta_{mix}S)_{sys} > 0 \Rightarrow$ randomness increases on mixing.

$(\Delta S)_{sys} > 0$

Probability of picking red ball is less as it has more randomness

Characteristics of Ideal Solution

(v)

$$(\Delta_{mix}S)_{surr} = 0 \Rightarrow (\Delta_{mix}S)_{surr} = \frac{-q_{sys}}{T} = 0$$

(vi)

$(\Delta_{mix}S)_{univ} > 0 \Rightarrow$ Mixing is a spontaneous process

Ideal Solution

Examples

- n-Hexane and n-Heptane
- Ethyl bromide and Ethyl iodide
- Benzene and Toluene
- Chlorobenzene and Bromobenzene

Non-Ideal Solution

Solutions that **do not obey Raoult's law** over the entire range of concentration

Non-Ideal Solution

If the forces of attraction between **A—A, B—B is different from A—B**, then A and B will form non-ideal solution.

Characteristics of Non-Ideal Solution

1 Raoult's law is **not obeyed**.

2 $\Delta H_{mix} \neq 0$

3 $\Delta V_{mix} \neq 0$

Non-Ideal Solution

- Non-ideal Solutions
  - Positive deviation from Raoult's law
  - Negative deviation from Raoult's law

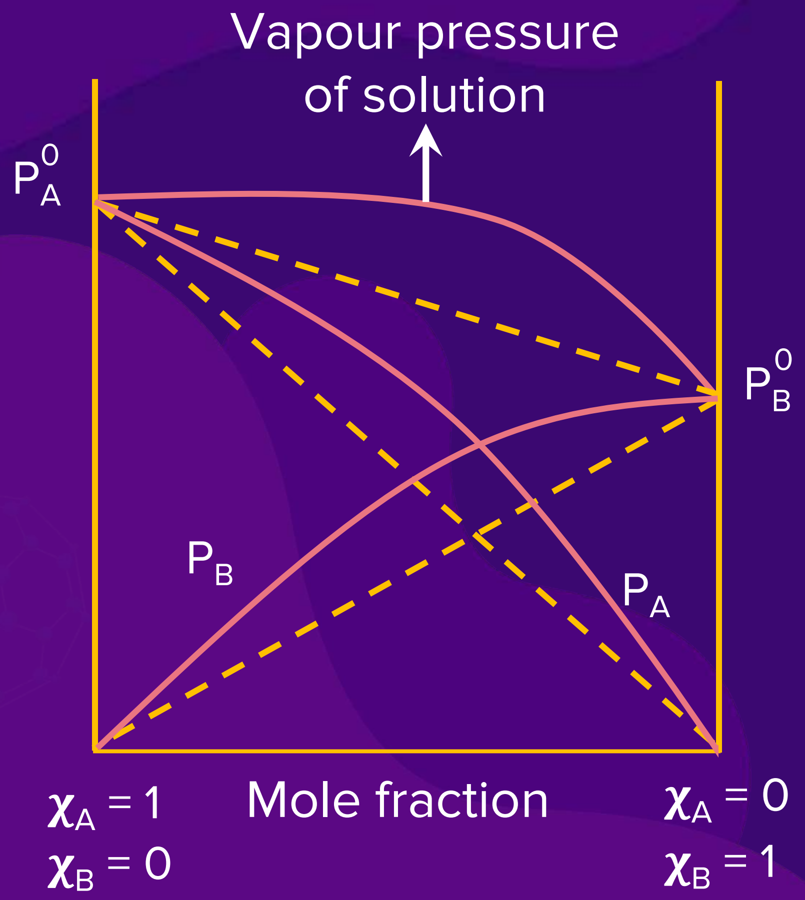
Positive Deviation

If the forces of attraction between **A—A, B—B is stronger than A—B**

Partial pressure of each component A and B is **higher** than that calculated from Raoult's law.

Hence, the **total pressure** over the solution is **also higher** than the solutions, if were ideal.

Positive Deviation



Dashed lines represent vapour pressures and total pressure corresponding to **ideal solution.**

V.P. ↑ → B.P. ↓

B.P. of solution < B.P. of both A and B

Characteristics of Positive Deviation

1 Raoult's law is **not obeyed**.

2 $\Delta V_{mix} > 0$ (1L + 1L > 2L)

Intermolecular forces between molecules of A and B are **weaker** than those in the pure liquids.

3 $\Delta H_{mix} > 0$

When the liquids are mixed, **less heat is evolved** when the new attractions are set up than was absorbed to break the original ones.

Characteristics of Positive Deviation

4 $(\Delta_{mix}S)_{sys} > 0 \Rightarrow$ randomness increases on mixing.

5 $(\Delta_{mix}S)_{surr} = 0 \Rightarrow (\Delta_{mix}S)_{surr} = \frac{-q_{sys}}{T} < 0$

6 $(\Delta_{mix}S)_{univ} > 0 \Rightarrow$ Mixing is a spontaneous process

Positive Deviation

Examples

- Water and ethanol
- Chloroform and water
- Ethanol and acetone
- Methanol and chloroform

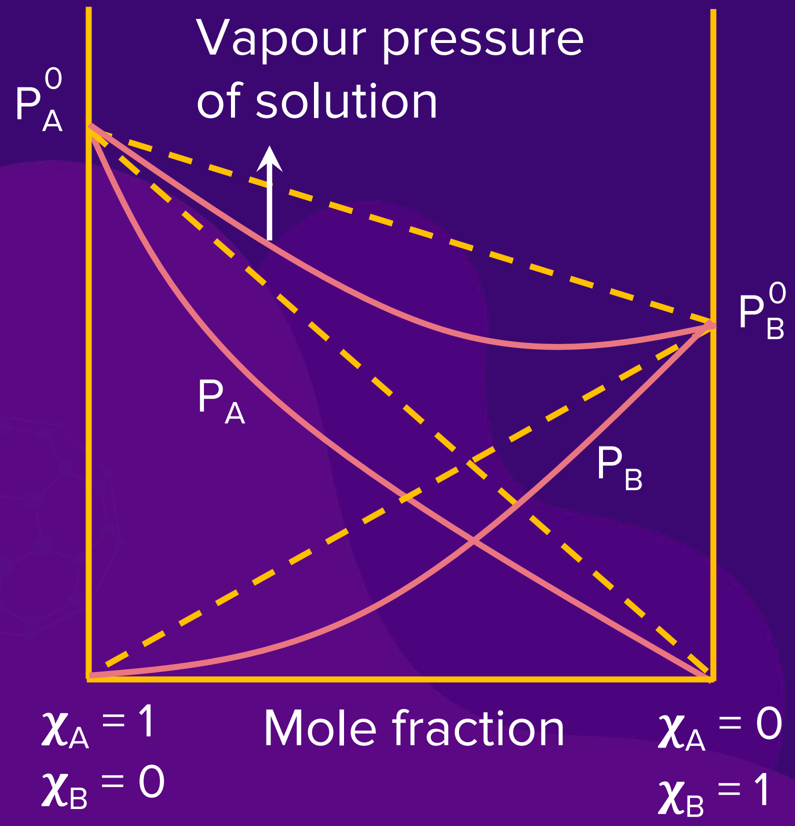
Negative Deviation

If the forces of attraction between **A—A, B—B are weaker than A—B**

Partial pressure of each component A and B is **lower** than that calculated from Raoult's law.

Hence, the **total pressure** over the solution is also **lower** than the solutions, if were ideal.

Negative Deviation



Dashed lines represent vapour pressures and total pressure corresponding to **ideal solution**.

V.P. ↓ B.P. ↑

B.P. of solution > B.P. of both A and B

Characteristics of Negatie Deviation

1 Raoult's law is **not obeyed**.

2 $\Delta V_{mix} < 0$ (1L + 1L < 2L)

Intermolecular forces between molecules of A and B are **stronger** than those in the pure liquids.

3 $\Delta H_{mix} < 0$

When the liquids are mixed, **more heat is evolved** when the new attractions are set up than was absorbed to break the original ones.

Characteristics of Negatie Deviation

4 $(\Delta_{mix}S)_{sys} > 0 \Rightarrow$ randomness increases on mixing.

5 $(\Delta_{mix}S)_{surr} = 0 \Rightarrow (\Delta_{mix}S)_{surr} = \frac{-q_{sys}}{T} > 0$

6 $(\Delta_{mix}S)_{univ} > 0 \Rightarrow$ Mixing is a spontaneous process

Examples of Negative deviation

- Chloroform and acetone
- Chloroform and methyl acetate
- H_2O and HCl
- H_2O and HNO_3
- Acetic acid and pyridine
- Phenol and aniline

Comparing Ideal & Non-Ideal Solutions

Ideal solutions	Non-ideal solutions	
	Positive deviation	Negative deviation
$P_T = \chi_A P_A^0 + \chi_B P_B^0$	$P_T > \chi_A P_A^0 + \chi_B P_B^0$	$P_T < \chi_A P_A^0 + \chi_B P_B^0$
A-A & B-B molecular interaction are similar as A-B	A-A & B-B molecular interaction are stronger than A-B	A-A & B-B molecular interaction are weaker than A-B
$\Delta_{mix}H = 0$	$\Delta_{mix}H > 0$	$\Delta_{mix}H < 0$
$\Delta_{mix}V = 0$	$\Delta_{mix}V > 0$	$\Delta_{mix}V < 0$

Comparing Ideal & Non-Ideal Solutions

Ideal solutions	**Non-ideal solutions**	
	Positive deviation	**Negative deviation**
$(\Delta_{mix}S)_{sys} > 0$	$(\Delta_{mix}S)_{sys} > 0$	$(\Delta_{mix}S)_{sys} > 0$
$(\Delta_{mix}S)_{surr} = 0$	$(\Delta_{mix}S)_{surr} < 0$	$(\Delta_{mix}S)_{surr} > 0$
$(\Delta_{mix}S)_{univ} > 0$	$(\Delta_{mix}S)_{univ} > 0$	$(\Delta_{mix}S)_{univ} > 0$
$(\Delta_{mix}G)_{sys} < 0$	$(\Delta_{mix}G)_{sys} < 0$	$(\Delta_{mix}G)_{sys} < 0$

The vapour pressure of the solution of two liquids A ($P^\circ = 80\ mm\ Hg$) and B ($P^\circ = 120\ mm\ Hg$) is found to be 100 mm Hg when $\chi_A = 0.4$. The result shows that:

BOARDS

(a) Solution exhibits ideal behavior

(b) Solution shows positive deviations

(c) Solution shows negative deviations

(d) Solution will show positive deviations for lower concentration and negative deviations for higher concentrations

Solution

Vapor pressure of solution as per Raoult's law should have been 104 *mm Hg* for it to become an ideal solution but it is 100 *mm Hg* as given in the question. Thus, solution shows negative deviation

Hence, option (c) is the correct answer.

P_A and P_B are the vapour pressures of pure liquid components, A and B, respectively of an ideal binary solution. If χ_A represents the mole fraction of component A, the total pressure of the solution will be:

BOARDS

(a) $P_A + \chi_A(P_B - P_A)$

(b) $P_A + \chi_A(P_A - P_B)$

(c) $P_B + \chi_A(P_B - P_A)$

(d) $P_B + \chi_A(P_A - P_B)$

Solution

$P_T = \chi_A P_A + \chi_B P_B$

$\chi_A + \chi_B = 1$

$\Rightarrow \chi_B = 1 - \chi_A$

$P_T = \chi_A P_A + (1 - \chi_A)P_B$

$P_T = \chi_A P_A + P_B - \chi_A P_B$

$\Rightarrow \mathbf{P_T = P_B + \chi_A(P_A - P_B)}$

Hence, option (d) is the correct answer.

The vapour pressure of water **depends upon**:

BOARDS

(a) Surface area of container

(b) Volume of container

(c) Temperature

(d) All of the above

Solution

The vapour pressure of water depends upon temperature.

Hence, option (c) is the correct answer.

The mixture which shows positive deviation from Raoult's law is:

BOARDS

(a) Ethanol + Acetone

(b) Benzene + Toluene

(c) Acetone + Chloroform

(d) Chloroethane + Bromoethane

Solution

Benzene + Toluene is ideal solution.

Acetone + Chloroform shows negative deviation.

Chloroethane + Bromoethane is ideal solution.

Hence, option (a) is the correct answer.

For an ideal solution, the correct option is:

BOARDS

(a) $\Delta_{mix}G = 0$ at constant T and P

(b) $\Delta_{mix}S = 0$ at constant T and P

(c) $\Delta_{mix}V \neq 0$ at constant T and P

(d) $\Delta_{mix}H = 0$ at constant T and P

Solution

For ideal solution :

$\Delta_{mix}G < 0$ at constant T and P

$\Delta_{mix}S > 0$ at constant T and P

$\Delta_{mix}V = 0$ at constant T and P

$\Delta_{mix}H = 0$ at constant T and P

Hence, option (d) is the correct answer.

Which of the following statements is correct regarding a solution of two components A and B exhibiting positive deviation from ideal behaviour?

BOARDS

(a) Intermolecular attractive forces between A-A and B-B are stronger than those between A-B

(b) $\Delta_{mix}H = 0$ at constant T and P

(c) $\Delta_{mix}V = 0$ at constant T and P

(d) Intermolecular attractive forces between A-A and B-B are equal to those between A-B

Solution

For a solution of two components A and B exhibiting positive deviation from ideal behaviour intermolecular attractive forces between A-A and B-B are stronger than those between A-B.

Hence, option (a) is the correct answer.

Which condition is not satisfied by an ideal solution?

BOARDS

(a) $\Delta_{mix}V = 0$

(b) $\Delta_{mix}S = 0$

(c) Obey the Raoult's law

(d) $\Delta_{mix}H = 0$

Solution

For ideal solution :

$\Delta_{mix}G < 0$ at constant T and P

$\Delta_{mix}S > 0$ at constant T and P

$\Delta_{mix}V = 0$ at constant T and P

$\Delta_{mix}H = 0$ at constant T and P

Hence, option (b) is the correct answer.

Azeotropic Mixtures

Very large deviations from ideality lead to a special class of mixtures known as **azeotropes, azeotropic mixtures, or constant-boiling mixtures**.

Azeotropic Mixtures

- Only non ideal solutions form azeotropes. Solutions do not have a constant boiling point.
- Composition of liquid A and vapour A becomes the same. Composition of liquid B and vapour B also becomes the same. So now the composition of liquid won't change even if it evaporates. Hence boiling point of liquid won't change.
- A liquid solution which evaporates without its change in composition is called azeotrope and are non ideal.

Azeotropes

Liquid mixtures which boils at a constant temperature and can be distilled **without any change** in the composition

A boiling liquid mixture at the azeotropic composition produces **vapours of exactly the same composition** as that of the liquid.

Composition of liquid mixtures at which **distillation cannot separate** the two liquids because the condensate has the **same composition** as that of the azeotropic liquid.

Azeotropes

Types of azeotropes

1. Minimum boiling azeotropes
2. Maximum boiling azeotropes

Minimum Boiling Azeotropes

Non-ideal solutions showing **large positive deviation** from Raoult's law form minimum boiling azeotropes that boil at a temperature **lower than the boiling points** of its components 'A' and 'B'.

Examples:

(i) Ethanol - water mixture containing ≈ 95% by volume of ethanol;

(ii) Chloroform - methanol mixture containing 87.4% chloroform and 12.6% methanol by mass

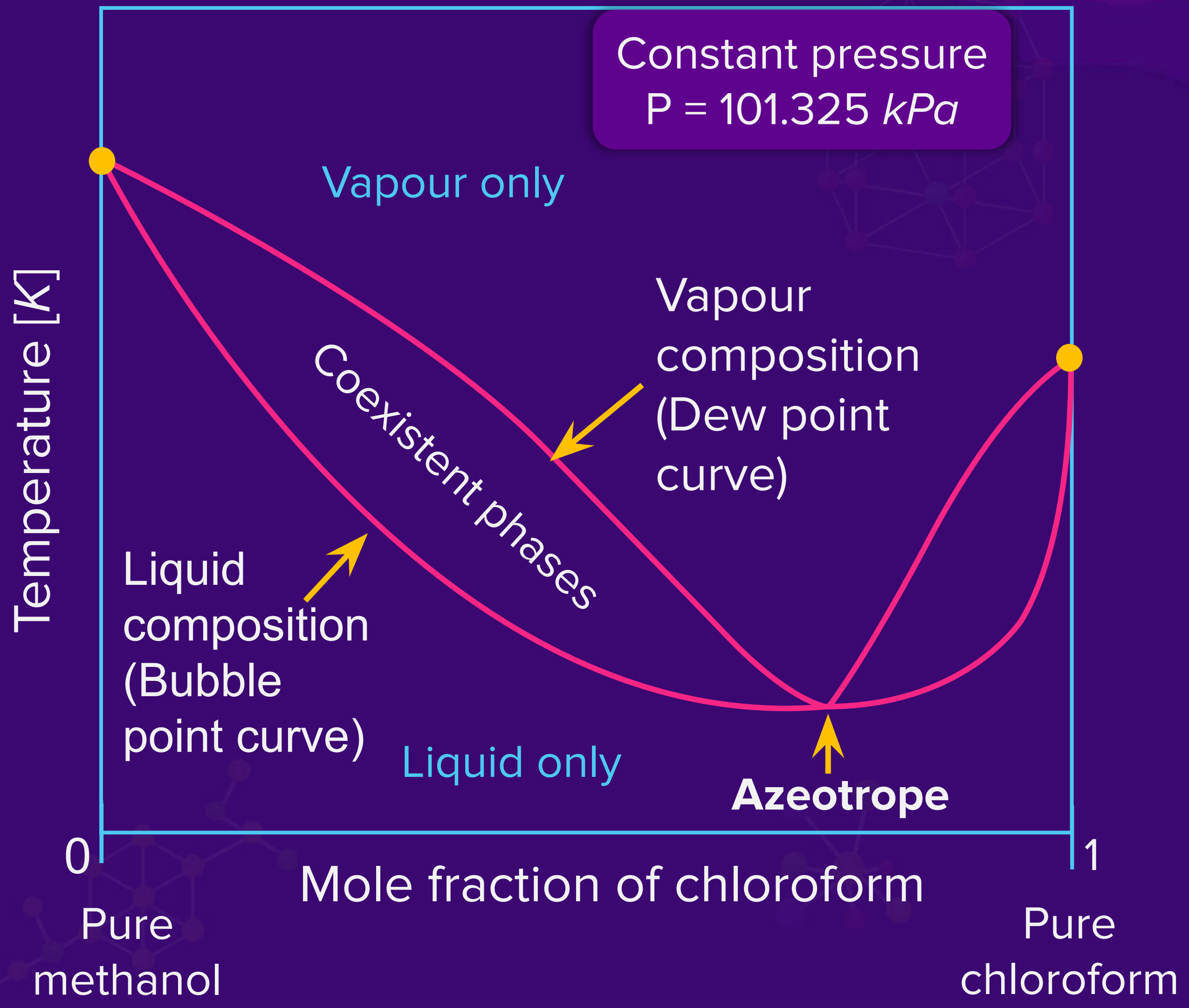
Minimum Boiling Azeotropes

Temperature vs composition graph

The composition of the point where dew point and bubble point curves intersect is azeotropic composition.

At this point the composition of the components of the solution is same in both liquid and vapour phase.

Boiling point of the solution at this composition is lesser than the boiling points of either of the components.



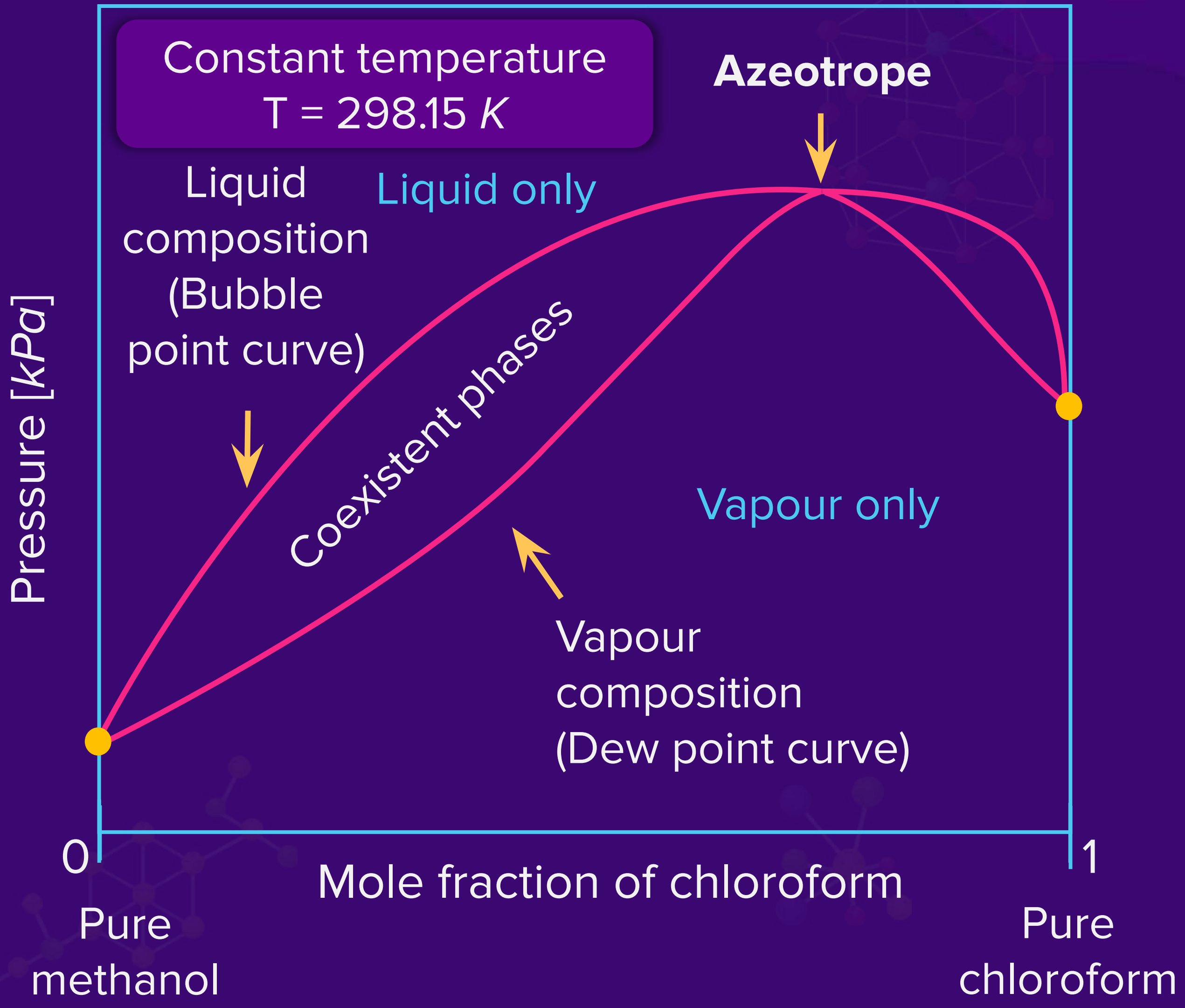
Minimum Boiling Azeotropes

Pressure vs composition graph

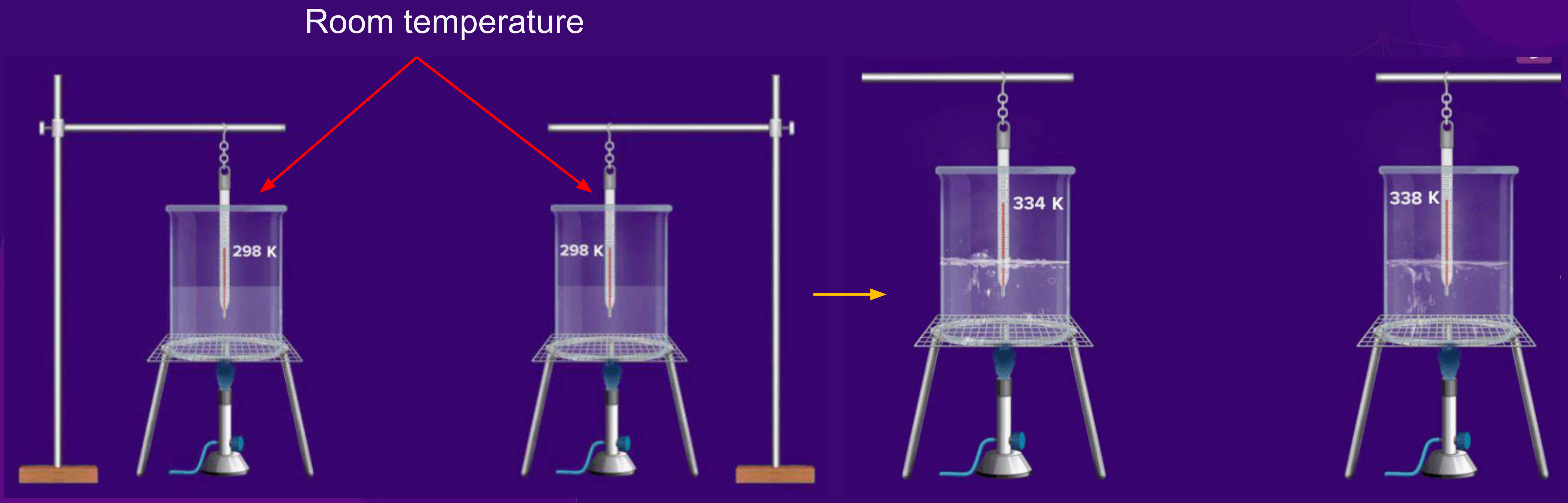
The composition of the point where dew point and bubble point curves intersect is azeotropic composition.

At this point the composition of the components of the solution is same in both liquid and vapour phase.

Vapour pressure of the solution at this composition is greater than the vapour pressures of the either of the components.



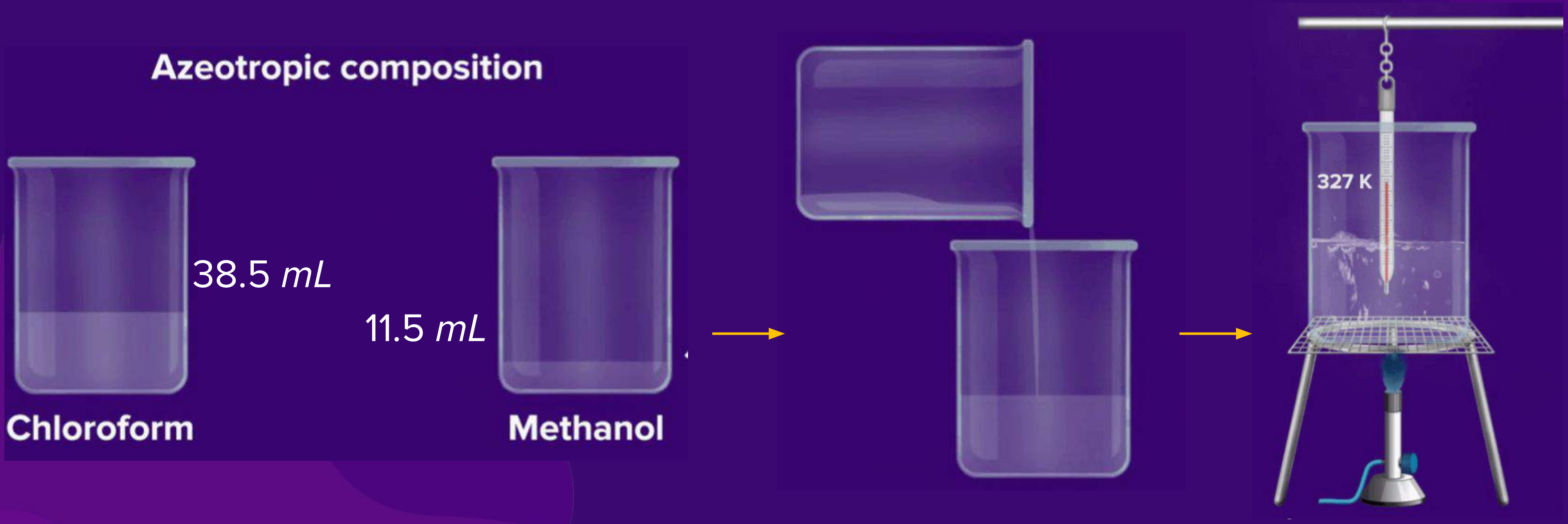
Minimum Boiling Azeotropes



Chloroform Methanol Chloroform Methanol

Both chloroform and methanol is taken at room temperature (298 *K*) and heated to their respective boiling points. Chloroform (B.P. = 334 *K*) and methanol (B.P. = 338 *K*)

Minimum Boiling Azeotropes



Then we mix them in certain composition and then boil the mixture. We will see that this mixture boiled at 327 *K* that is at a lower temperature as compared to their individual boiling point. And this is only possible in **minimum boiling azeotropes.**

Maximum Boiling Azeotropes

Non-ideal solutions showing **large negative deviation** from Raoult's law form maximum boiling azeotropes that boil at a temperature **higher than the boiling points** of its individual components.

Examples:

(i) A mixture of HNO_3 and H_2O containing ≈ 68% HNO_3 and 32% water by mass

(ii) Formic acid - water mixture containing 77.6% formic acid (CH_2O_2) and 22.4% water by mass

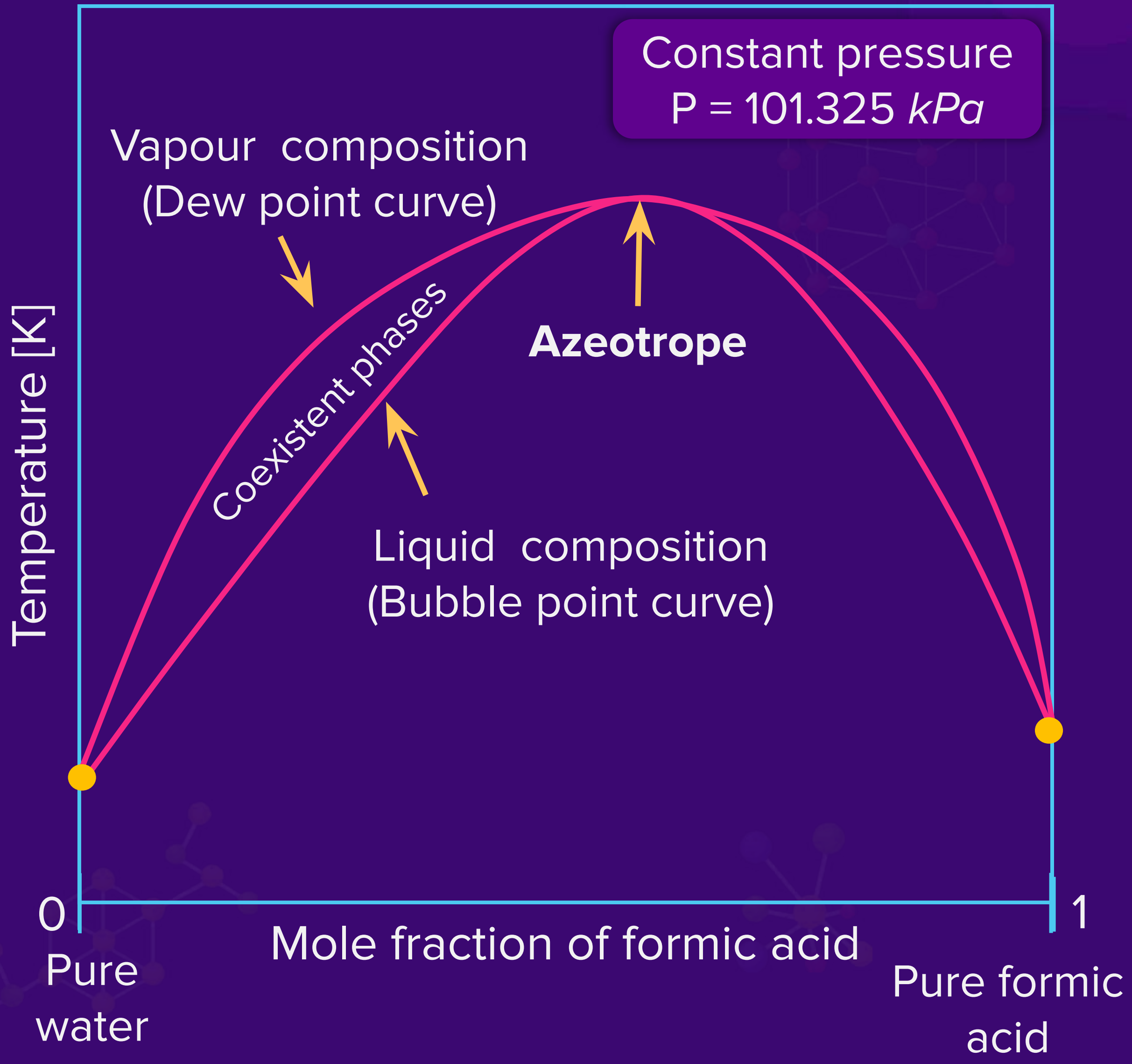
Maximum Boiling Azeotropes

Temperature vs composition graph

The composition of the point where dew point and bubble point curves intersect is azeotropic composition.

At this point the composition of the components of the solution is same in both liquid and vapour phase.

Boiling point of the solution at this composition is higher than the boiling points of either of the components.



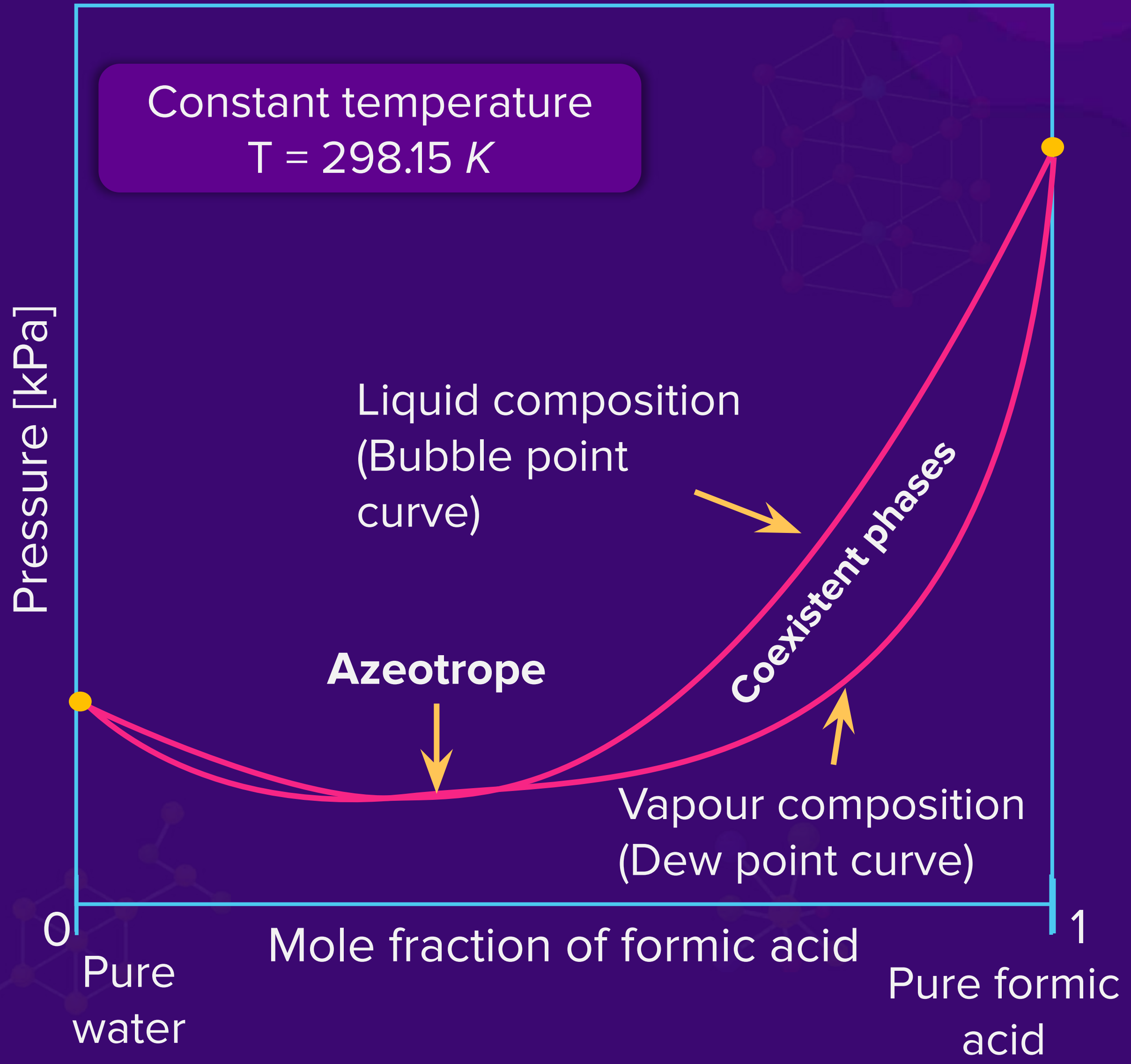
Maximum Boiling Azeotropes

Pressure vs composition graph

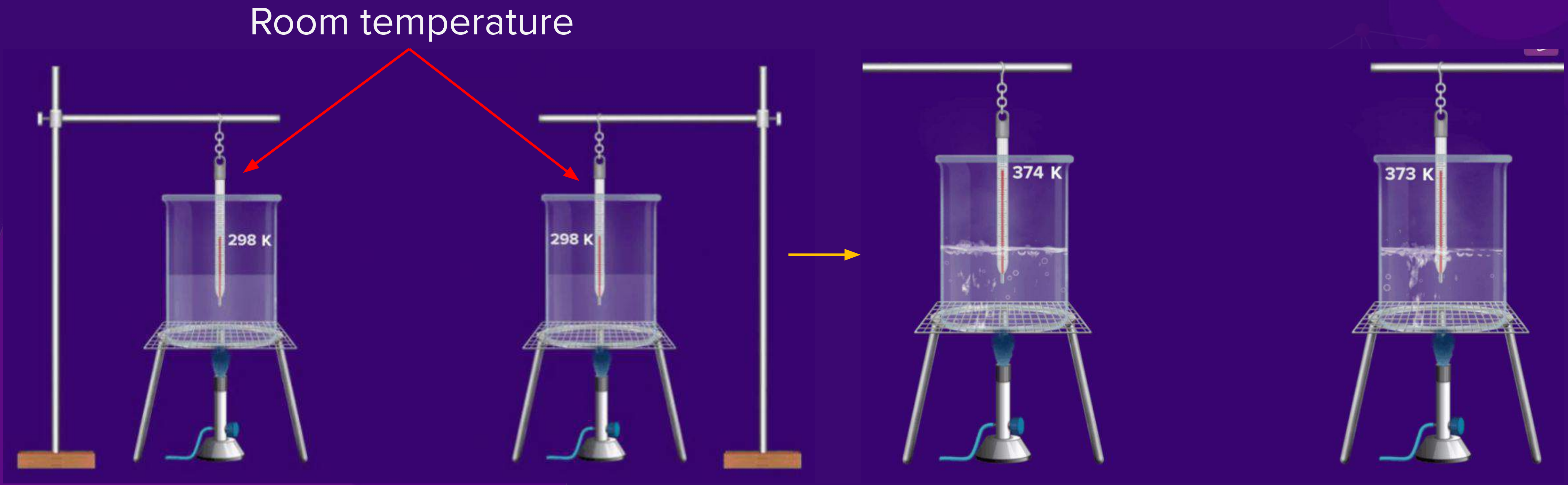
The composition of the point where dew point and bubble point curves intersect is azeotropic composition.

Vapour pressure of the solution at this composition is lesser than the vapour pressures of the either of the components.

An ideal solution does not boil at a constant temperature, whereas azeotrope does.



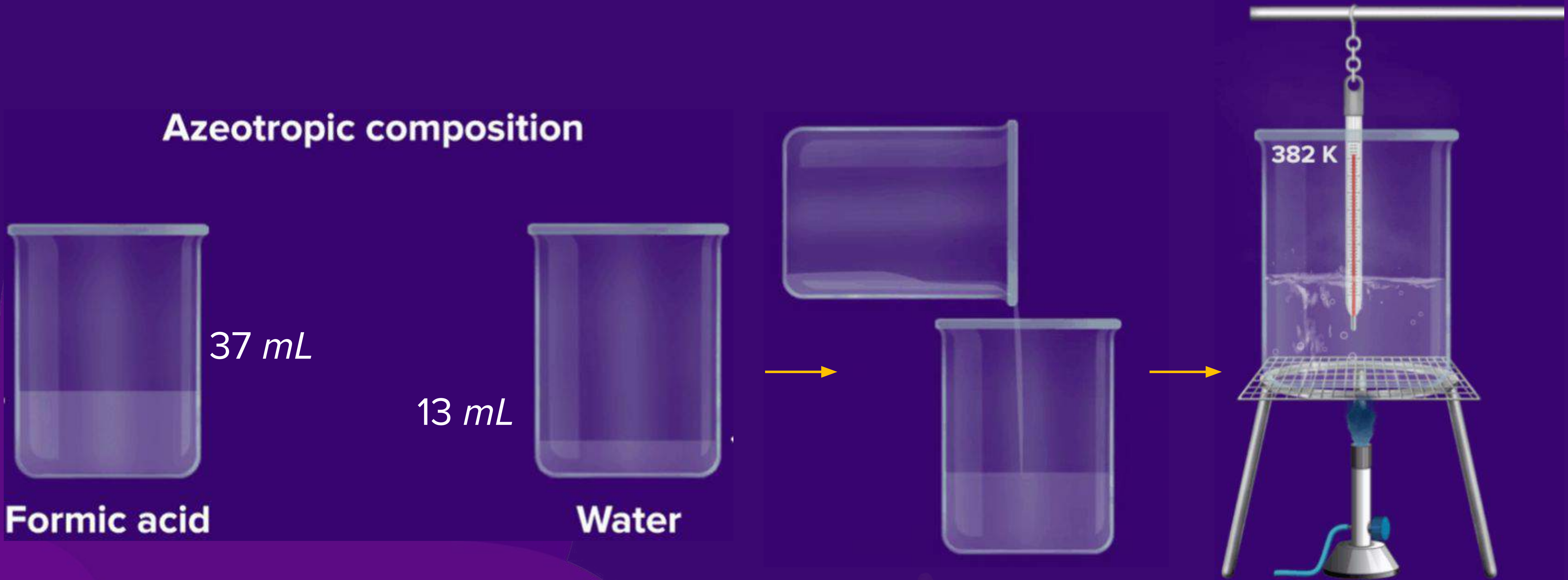
Maximum Boiling Azeotropes



Formic acid Water Formic acid Water

Both water and formic acid is taken at room temperature (298 *K*) and heated to their respective boiling points.
Formic acid (B.P. = 374 *K*) and water (B.P. = 373 *K*)

Maximum Boiling Azeotropes



Then we mix them in a certain composition and then boil the mixture. We will see that this mixture boiled at 382 K that is at a higher temperature as compared to their individual boiling point. And this is only possible in **maximum boiling azeotropes.**

An ideal solution does not boil at a constant temperature, whereas azeotrope does.

Which of the following **azeotropic solutions** has the boiling point **more** than the **boiling point** of the constituents **A** and **B**?

BOARDS

(a) CH_3CH_2OH and CH_3COCH_3

(b) CS_2 and CH_3COCH_3

(c) $CHCl_3$ and CH_3COCH_3

(d) CH_3CHO and CS_2

Solution

The one which shows negative deviation will have more boiling point which is a mixture of chloroform and acetone.

Hence, option (c) is the correct answer.

How many facts related to $H_2O + HCl$ solution are correct?

BOARDS

(a) It shows negative deviation

(b) It forms maximum boiling azeotropic mixture

(c) $\Delta S > 0$

(d) All of these

Solution

(a) As $H_2O + HCl$ will have H-bonding so it will show negative deviation.

(b) All solutions showing negative deviation forms maximum boiling azeotropic mixture.

(c) On mixing, entropy always increases, hence $\Delta S > 0$.

Hence, option (d) is the correct answer.

Which is/are correct statement(s)?

BOARDS

(a) When a mixture is more volatile than both pure components, there is positive deviation from Raoult’s law

(b) When the mixture is more volatile than both pure components, there is negative deviation from Raoult’s law

(c) Ethanol and water form ideal solution

(d) $CHCl_3$ and water form ideal solution

Solution

If the mixture is more volatile than both pure components then their vapour pressure will be higher than what we obtained from Raoult's law i.e. mixture will show positive deviation.
Ethanol and water do not form ideal solution but form positive deviation.
$CHCl_3$ and water do not form ideal solution.
Hence, option (a) is the correct answer.

Benzene and Toluene form an ideal solution at room temperature. For this process, the **false statement** is:

BOARDS

(a) $\Delta H = 0$

(b) ΔS_{system} is positive

(c) $\Delta S_{surroundings} = 0$

(d) ΔG is positive

Solution

- As they form ideal solution so $\Delta H = 0$.
- ΔS_{system} is positive.
- There is no effect on surrounding so $\Delta S_{surroundings} = 0$.
- ΔG for this process is negative as mixing is a spontaneous process.

Hence, option (d) is the correct answer.

A binary liquid solution is prepared by mixing n-heptane and ethanol. Which one of the following statement is correct regarding the behaviour of the solution?

BOARDS

a The solution is non-ideal, showing +ve deviation from Raoult's law.

b The solution in non-ideal, showing –ve deviation from Raoult's law.

A binary liquid solution is prepared by mixing n-heptane and ethanol. Which one of the following statement is correct regarding the behaviour of the solution?

c n-heptane shows +ve deviation while ethanol shows –ve deviation from Raoult's Law.

d The solution formed is an ideal solution.

Hence, option (a) is the correct answer.

For an ideal solution of two components A and B, which of the following is true?

BOARDS

(a) $\Delta_{mix}H < 0$

(b) $\Delta_{mix}H > 0$

(c) A–B interaction is stronger than A–A and B–B interactions

(d) A–A, B–B, and A–B interactions are identical

Solution

For an ideal solution, $\Delta_{mix}H = 0$

In ideal solutions, all the interactions are identical and therefore **option (d)** A–A, B–B, and A–B interactions are identical is the correct answer.

Hence, option (d) is the correct answer.

The mixture that forms **maximum boiling azeotrope** is:

BOARDS

(a) Heptane + octane

(b) Water + nitric acid

(c) Ethanol + water

(d) Acetone + carbon disulphide

Solution

Maximum boiling azeotropes are those whose boiling point is more than their expected boiling point and show negative deviation.
(a) Heptane + octane, are alkanes and almost form ideal solutions.
(b) Water + nitric acid, shows negative deviation. Solutions of water and acid always shows negative deviation.
(c) Ethanol + water, shows positive deviation. Solutions of water and alcohols except methanol always shows positive deviation.
(d) Acetone + carbon disulphide, shows positive deviation because carbon disulphide is non-polar and acetone is polar.

Hence, option (b) is the correct answer.

The concentration unit, **independent of temperature** would be:

BOARDS

(a) Normality

(b) Weight volume percent

(c) Molality

(d) Molarity

Solution

The concentration unit which is independent of temperature will also be volume independent.

(a) Normality is equivalents per liters of solution, so this is volume dependent.

(b) Weight volume percent is also volume dependent.

(c) Molality is moles of solutes per kg of solvent and this is volume independent and therefore will also be independent of temperature.

(d) Molarity is moles of solute per liters of solution, so this is volume dependent.

Hence, option (c) is the correct answer.

Assertion: One molal aqueous solution of glucose contains 180 *g* of glucose in 1 *kg* water.

Reason: Solution containing one mole of solute in 1000 *g* of solvent is called one molal solution.

BOARDS

(a) If both Assertion and Reason are correct and the Reason is a correct explanation of the Assertion

(b) If both Assertion and Reason are correct but Reason is not a correct explanation of the Assertion

(c) If the Assertion is correct but Reason is incorrect

(d) If both the Assertion and Reason are incorrect

Solution

Assertion: Molar mass of Glucose($C_6H_{12}O_6$) = 180 $g\ mol^{-1}$

Given, mass of glucose = 180 g

Moles of glucose = $\frac{180}{180}$ = 1 *mole*

Molality = Moles of solute/solvent in kg = $\frac{1}{1}$ = 1 *molal*

Therefore, Assertion is **true.**

Reason: Solution containing one mole of solute in 1000 g of solvent is called one molal solution and this is also **true.**

Hence, option (a) is the correct answer.

Which one of the following is incorrect for ideal solution?

BOARDS

Solution

a $\Delta H_{mix} = 0$

b $\Delta V_{mix} = 0$

c $\Delta P = P_{obs} - P_{\text{calculated by Raoult's law}} = 0$

d $\Delta G_{mix} = 0$

Hence, option (d) is the correct answer.

Which of the following statements about the composition of the vapour over an ideal 1 : 1 molar mixture of benzene and toluene is correct? Assume that the temperature is constant at 25 °C. (Given, vapour pressure data at 25 °C, benzene = 12.8 *kPa*, toluene = 3.85 *kPa*)

BOARDS

(a) The vapour will contain equal amounts of benzene and toluene

(b) Not enough information is given to make a prediction

(c) The vapour will contain a higher percentage of benzene

(d) The vapour will contain a higher percentage of toluene

Solution

Moles of Benzene (n_B) = Moles of Toluene (n_T) = 1
Mole fraction of Benzene (x_B) = 0.5 and Mole fraction of Toluene (x_T) = 0.5
$P^0_B = 12.8\ kPa$ and $P^0_T = 3.85\ kPa$
Vapour pressure of benzene is more which means benzene is more volatile than toluene therefore vapour will contain more amounts of benzene. The vapour will contain a higher percentage of benzene because vapour pressure of benzene is more.

Hence, option (c) is the correct answer.

A solution of acetone in ethanol:

BOARDS

(a) Obeys Raoult’s law

(b) Shows a negative deviation from Raoult’s law

(c) Shows a positive deviation from Raoult’s law

(d) Behaves like a near ideal solution

Solution

Ethanol forms hydrogen bonds and acetone will interrupt the hydrogen bond between ethanol molecule and the extent of hydrogen bonding will decrease. Therefore, the solution will show positive deviation from Raoult's law.

Hence, option (c) is the correct answer.

A solution containing components A and B follows Raoult's law:

BOARDS

(a) A–B attraction force is greater than A–A and B–B

(b) A–B attraction force is less than A–A and B–B

(c) A–B attraction force remains same as A–A and B–B

(d) Volume of solution is different from sum of volume of solute and solvent

Solution

In ideal solutions, all the interactions are identical and therefore option (c) A–A, B–B, and A–B interactions are identical is the correct answer.

Hence, option (c) is the correct answer.

All form ideal solution except:

BOARDS

(a) C_6H_6 & $C_6H_5CH_3$

(b) n-Hexane & n-heptane

(c) C_6H_5Cl & C_6H_5Br

(d) Phenol & aniline

Solution

(a) C_6H_6 & $C_6H_5CH_3$, both are nonpolar and hence they will form ideal solution.

(b) n-Hexane & n-heptane, both are nonpolar and hence they will form ideal solution.

(c) C_6H_5Cl & C_6H_5Br are almost of similar kind and hence they will form ideal solution.

(d) Phenol & aniline will form hydrogen bonding and will not form ideal solution and will show negative deviation.

Hence, option (d) is the correct answer.

A liquid is kept in a closed vessel. If a glass plate (negligible mass) with a small hole is kept on top of the liquid surface, then the vapour pressure of the liquid in the vessel is:

BOARDS

(a) More than what would be, if the glass plate were removed

(b) Same as what would be, if the glass plate were removed

(c) Less than what would be, if the glass plate were removed

(d) Cannot be predicted

Solution

Here vapour pressure of liquid depends only on temperature so here vapour pressure will remain same.

Hence, option (b) is the correct answer.

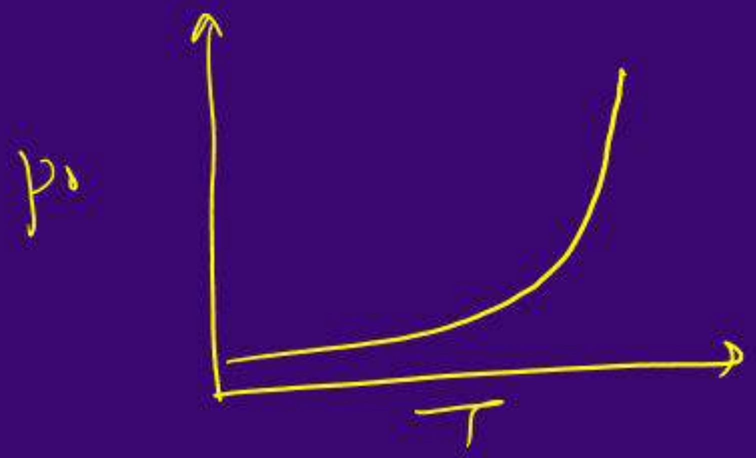
Choose the correct statement with respect to the vapour pressure of a liquid among the following:

- **(a)** Increases linearly with increasing temperature
- **(b)** Increase non-linearly with increasing temperature
- **(c)** Decreases linearly with increasing temperature
- **(d)** Decreases non-linearly with increasing temperature

Solution

Vapour pressure of a liquid increases exponentially with increasing temperature.

Hence, option (b) is the correct answer.



Low concentration of oxygen in blood and tissues of people living at higher altitude is due to:

BOARDS

(a) Low temperature

(b) Low atmospheric pressure

(c) High atmospheric pressure

(d) Both low temperature & high pressure

Solution

At higher altitude the pressure is low but the human body temperature is same. As the partial pressure of oxygen is decrease at higher altitude the solubility of oxygen in blood will decrease.

Hence, option (b) is the correct answer.

Solubility of Solid in Liquid

Dissolution

When a solid solute is added to the solvent, some **solute dissolves** and its **concentration increases** in the solution.

Crystallisation

Some solute particles in the solution **collide with the other solid solute** particles and get **separated** out of solution.

Dissolution & Crystallisation

At equilibrium

Rate of dissolution = Rate of crystallisation

$$\text{Solute} + \text{Solvent} \rightleftharpoons \text{Solution}$$

Equilibrium

At this stage, the concentration of solute in solution will **remain constant** under the given temperature and pressure conditions

Such a solution is said to be **saturated** with the given solute.

Factors Affecting Solubility

Nature of the solvent & the solute

Temperature ✓

Pressure ✓

Nature of Solvent & Solute

Like dissolves like

Polar solutes dissolve in **polar solvents** and **non-polar solute** dissolve in **non-polar solvents**.

Example

Sodium chloride **dissolves readily in water**, whereas naphthalene and anthracene do not.

Naphthalene and anthracene **dissolve readily in benzene** but sodium chloride do not.

Temperature

$$\text{Solute} + \text{Solvent} \rightleftharpoons \text{Solution}$$

If dissolution is **exothermic**

$\Delta H < 0$

By Le Chatelier's principle,

T ↑ Solubility ↓

Temperature

Solute + Solvent $\rightleftharpoons$ Solution

If dissolution is **endothermic** | **$\Delta H > 0$**

By Le Chatelier's principle,

T ↑ » Solubility ↑

Pressure

Pressure **does not have any significant effect on solubility** of solids in liquids.

Solids and liquids are **highly incompressible** and practically remain unaffected by changes in the pressure.

Colligative Properties

The properties of the solution that are **dependent only on the total number of solute particles** relative to solvent/solution

They are **not dependent on the nature of particle** i.e., shape, size, charge, etc.

Colligative Properties

of solutions

- Relative lowering of vapour pressure
- Elevation in boiling point
- Depression in freezing point
- Osmotic pressure

Abnormal Colligative Property

If solute gets associated or dissociated in solution, then experimental/observed/actual value of colligative property will be **different** from theoretically predicted value.

Abnormal Colligative Property

For electrolytic solutes, the number of particles would be **different** from the number of particles actually added, due to **dissociation or association** of solute.

Abnormal Colligative Property

$$\mathrm{NaCl}\ (s) + \mathrm{H_2O}\ (l) \longrightarrow \mathrm{Na^+}\ (aq) + \mathrm{Cl^-}\ (aq)$$

Example: On adding **1 mole of NaCl** in excess water gives 1 mole of Na^+ and 1 mole of Cl^- ions i.e. 2 moles of solute.

Abnormal Colligative Property

Abnormality in colligative property can be calculated in terms of **van't-Hoff factor**. (i)

van’t–Hoff Factor

The actual extent of **dissociation/ association** can be expressed with a **correction factor** known as van’t Hoff factor (i).

van’t–Hoff Factor (i)

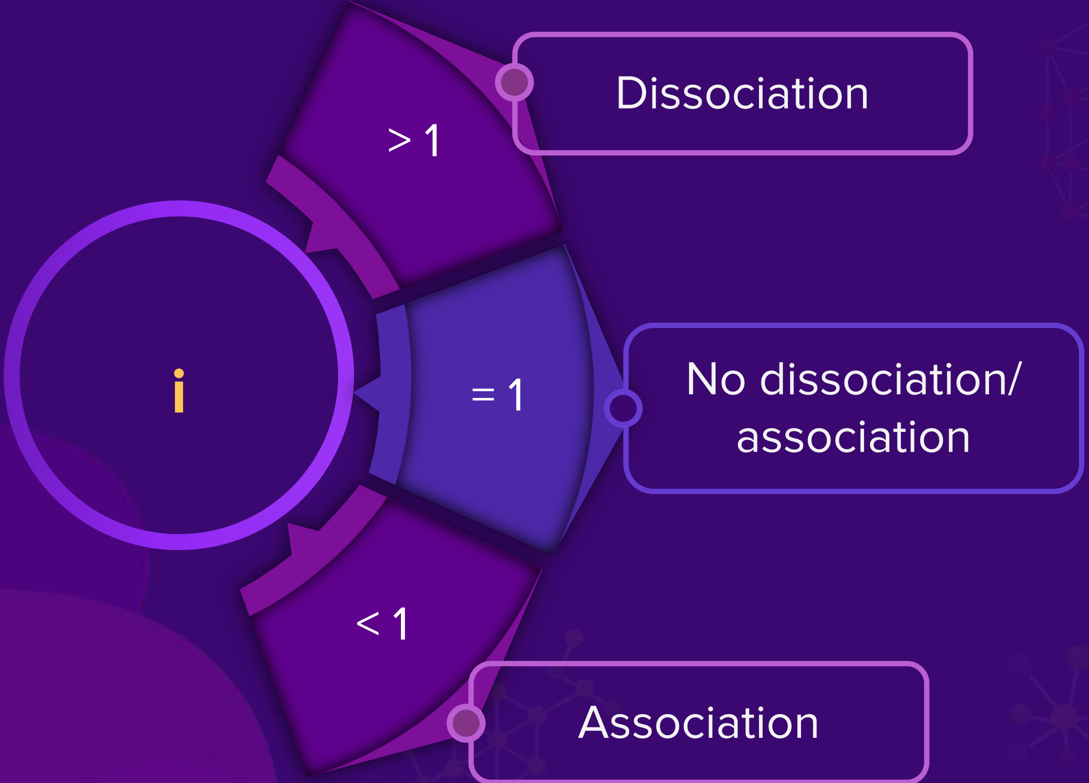
$$i = \frac{\overset{\text{Experimental / Actual}}{\text{Observed colligative property}}}{\left(\underset{\text{Theoretical}}{\cancel{\text{Calculated}}}\text{ colligative property assuming no assoc}^{n}\text{/dissoc}^{n}\right)}$$

$$= \frac{\left(\overset{\text{assuming no assoc}^{n}\text{/dissoc}^{n}}{\text{Theoretical molar mass of substance}}\right)}{\underset{\text{or observed}}{\text{Experimental molar mass of the substance}}}$$

van’t–Hoff Factor (i)

$$i = \frac{\text{Moles of solute particles in solution after dissociation/ association}}{\text{Moles of solute particles before association/dissociation}}$$

van’t–Hoff Factor (i)



i

> 1

Dissociation

= 1

No dissociation/ association

< 1

Association

i for Dissociation of Electrolyte

$$\boxed{i = \frac{C}{C_0} = 1 + (x+y-1)\alpha}$$

$$\alpha = \frac{n_{dissociated}/V}{n_{initial}/V} = \frac{C_{dissociated}}{C_0}$$

$$\underline{A_xB_y\,(aq)} \rightleftharpoons \underline{xA^{y+}} + \underline{yB^{x-}}$$

t = 0	C_0		0		0
Δ	$-C_0\alpha$		$+xC_0\alpha$	+	$yC_0\alpha$
t = t_{eq}	$C_0(1 - \alpha)$	+	$xC_0\alpha$	+	$yC_0\alpha$

$$= C_{final} = C_0\left[1 - \alpha + x\alpha + y\alpha\right] = C_0\left[1 + (x+y-1)\alpha\right]$$

i for Dissociation of Electrolyte

At equilibrium, **Net concentration** $= C - C\alpha + xC\alpha + yC\alpha$

$= C[1 + (x + y - 1)\alpha]$

$= C[1 + (n - 1)\alpha]$

n Number of particles in which 1 molecule of electrolyte dissociates

i for Dissociation of Electrolyte

$$i = \frac{C\,[1 + (n - 1)\,\alpha\,]}{C}$$

$$i = 1 + (\overset{x+y}{n} - 1)\,\alpha$$

Examples

NaCl (100% ionised) i = 2

$BaCl_2$ (100% ionised) i = 3

Determine i (van't-Hoff factor) for:

1. $K_4[Fe(CN)_6]$ (75% ionisation)
2. $KCl.MgCl_2.6H_2O$(100% ionisation)

Solution

Answer- 1. i=4
Answer- 2. i=5

$\alpha = 75/100 = 0.75$

$i = \frac{n_{final}}{n_0} = 1+4\alpha$

$= 1+4\times 0.75$

$= 4$

$i = 5/1 = \underline{\underline{5}}$

$\underline{\underline{i=4}}$

$\underline{\underline{i=5}}$

$\xrightarrow{aq} 1K^+(aq) + 1Cl^-(aq) + 1Mg^{2+}(aq) + 2Cl^-(aq)$

$1+1+1+2 = \underline{\underline{5}}$

$K_4[Fe(CN)_6] \longrightarrow 4K^+ + [Fe(CN)_6]^{4-}$

ini $\quad n_0 \quad 0 \quad 0$

D $\quad -n_0\alpha \quad +4n_0\alpha \quad n_0\alpha$

at eq. $\quad n_0(1-\alpha) + 4n_0\alpha + n_0\alpha = n_f$

$n_{final} = n_0[1-\cancel{\alpha}+4\alpha+\cancel{\alpha}] = n_0[1+4\alpha]$

i for Association of Electrolyte

$nA \rightleftharpoons A_n$

$nA \rightleftharpoons 1A_n$

$t = 0$	C_0		0	
Δ	$-C_0\beta$		$+\frac{C_0\beta}{n}$	
$t = t_{eq}$	$C_0(1-\beta)$	$+$	$\frac{C_0\beta}{n}$	$= C_{final}$

$\therefore C_{final} = C_0\left[1 - \beta + \frac{\beta}{n}\right] \quad \therefore i = \frac{C_{final}}{C_0} = 1 - \beta + \frac{\beta}{n}$

i for Association of Electrolyte

At equilibrium, **Net concentration** $= C_0 - \underline{C_0\beta + \frac{C_0\beta}{n}}$

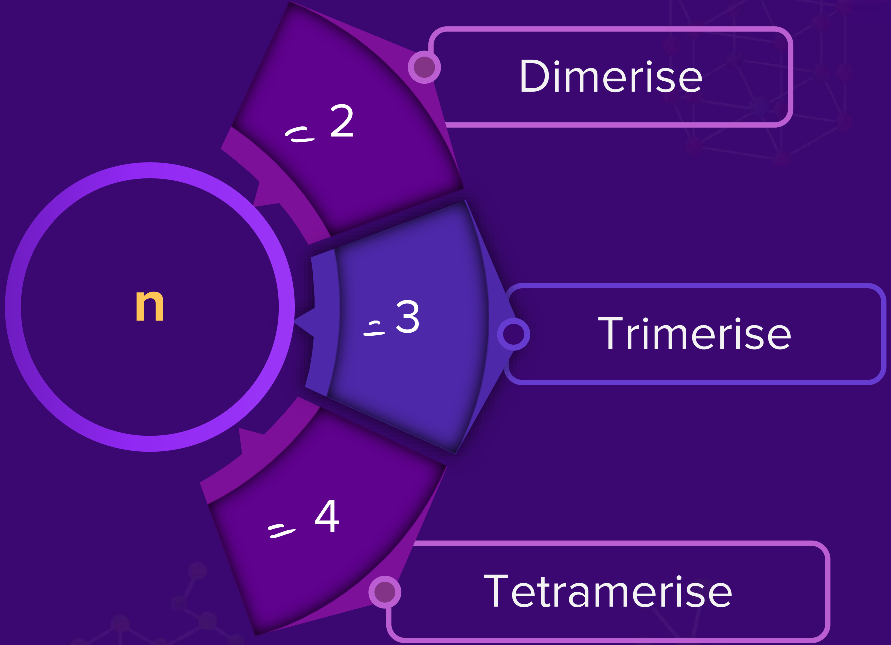
$$= C_0\left[1 + \left(\frac{1}{n} - 1\right)\beta\right]$$

i for Association of Electrolyte

$$i = \frac{C\left[1 + \left(\frac{1}{n} - 1\right)\beta\right]}{C}$$

$$i = 1 + \left(\frac{1}{n} - 1\right)\beta$$

n for Association of Electrolyte



n

≃ 2 Dimerise

≃ 3 Trimerise

≃ 4 Tetramerise

Examples

Dimerisation of CH_3COOH
in benzene (100%)

$i = 0.5$

Dimerisation of C_6H_5COOH
in benzene (100%)

$i = 0.5$

An aqueous solution of a salt MX_2 at a certain temperature has a van’t Hoff factor of 2. The degree of dissociation for this solution of the salt is:

$$MX_2 \rightleftharpoons M^{2+} + 2X^-$$

$$n_0(1-\alpha) + n_0\alpha + 2n_0\alpha = n$$

$$n = n_0[1-\alpha+\alpha+2\alpha]$$

$$= n_0[1+2\alpha]$$

$$\therefore i = \frac{n}{n_0} = 1+2\alpha = 2$$

$$\therefore 2\alpha = 2-1$$

$$= 1$$

$$\alpha = \frac{1}{2}$$

a 0.67

b 0.33

c 0.80

[✓ **d**] 0.50

Solution

Hence, option d is the correct answer.

The van't Hoff factor for 0.1 M $Ba(NO_3)_2$ solution is 2.74.

The degree of dissociation is:

$$Ba(NO_3)_2 \longrightarrow Ba^{2+} + 2NO_3^-$$

$$n_0(1-\alpha) \quad + \quad n_0\alpha + \quad 2n_0\alpha \quad = n_f$$

$$\therefore n_f = n_0[1 - \alpha + \alpha + 2\alpha] = n_0(1+2\alpha)$$

$$\therefore i = \frac{n_f}{n_0} = 1+2\alpha$$

$$2.74 = 1+2\alpha$$

$$\therefore 2\alpha = 1.74$$

$$\alpha = \frac{1.74}{2}$$

$$= 0.87$$

a 91.3%

[✓ b] 87%

c 100%

d 74%

Solution

Hence, option b is the correct answer.

van’t Hoff factor :

Colligative Properties

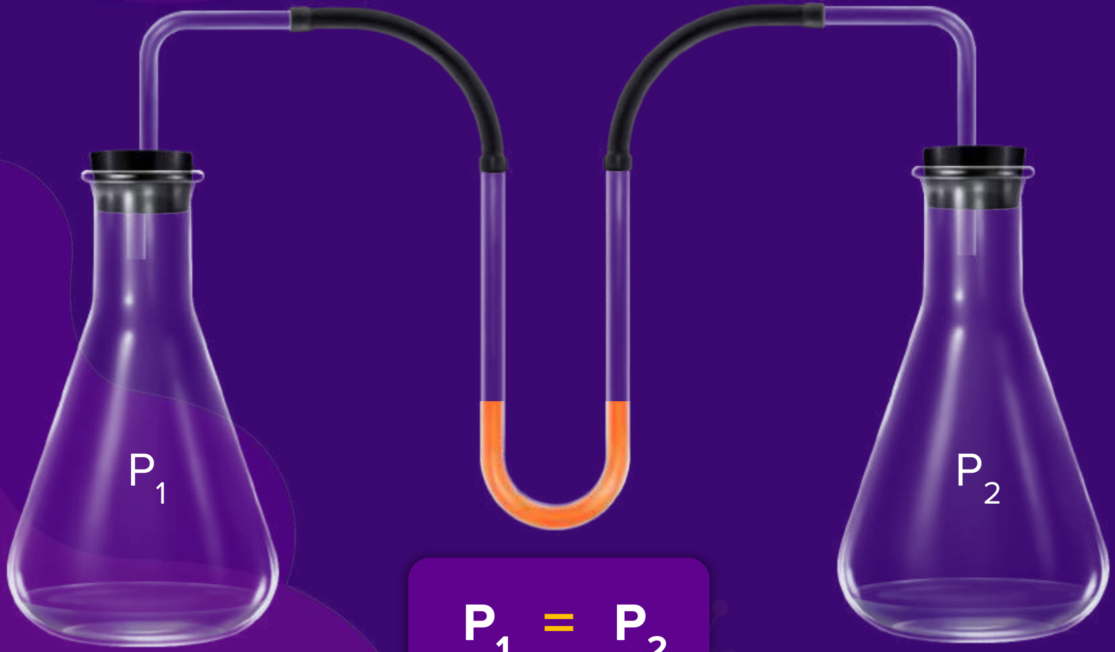
Colligative Properties

- Relative lowering of vapour pressure
- Elevation in boiling point
- Depression in freezing point
- Osmotic pressure

Relative Lowering of Vapour Pressure

V.P. of a **solution** containing a non-volatile solute (solid solute) is always found to be **less** than the V.P. of the **pure solvent**.

Relative Lowering of Vapour Pressure

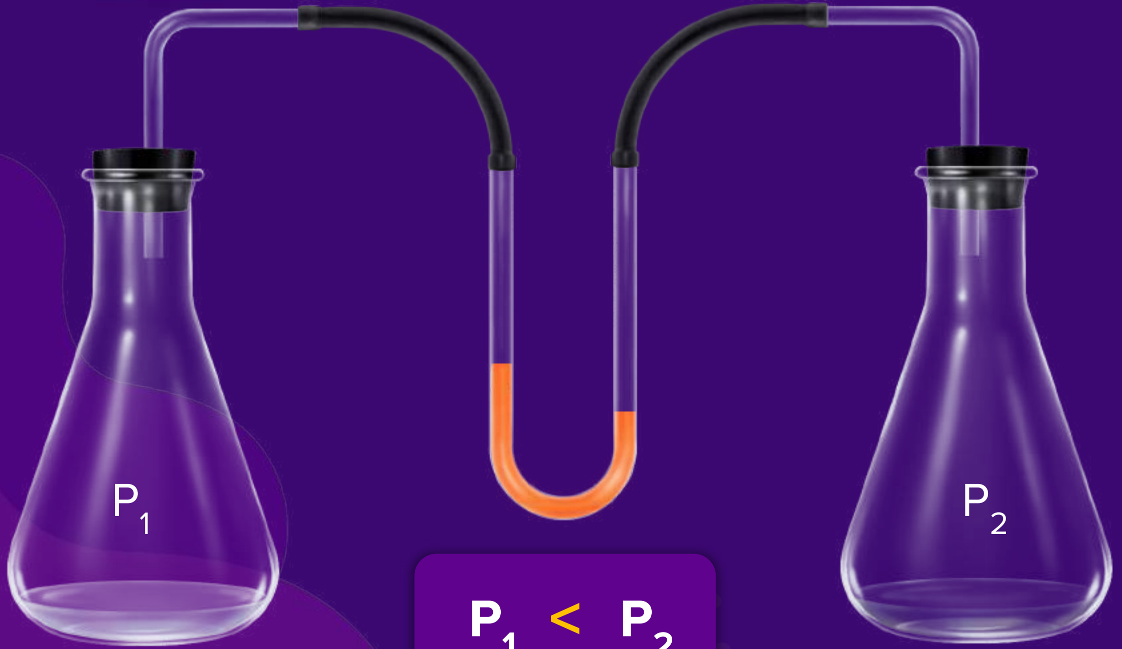


$P_1 = P_2$

Pure solvent A

Pure solvent A

Relative Lowering of Vapour Pressure



$P_1 < P_2$

Solution (non-volatile
solute in solvent A)

Pure
solvent A

Relative Lowering of Vapour Pressure

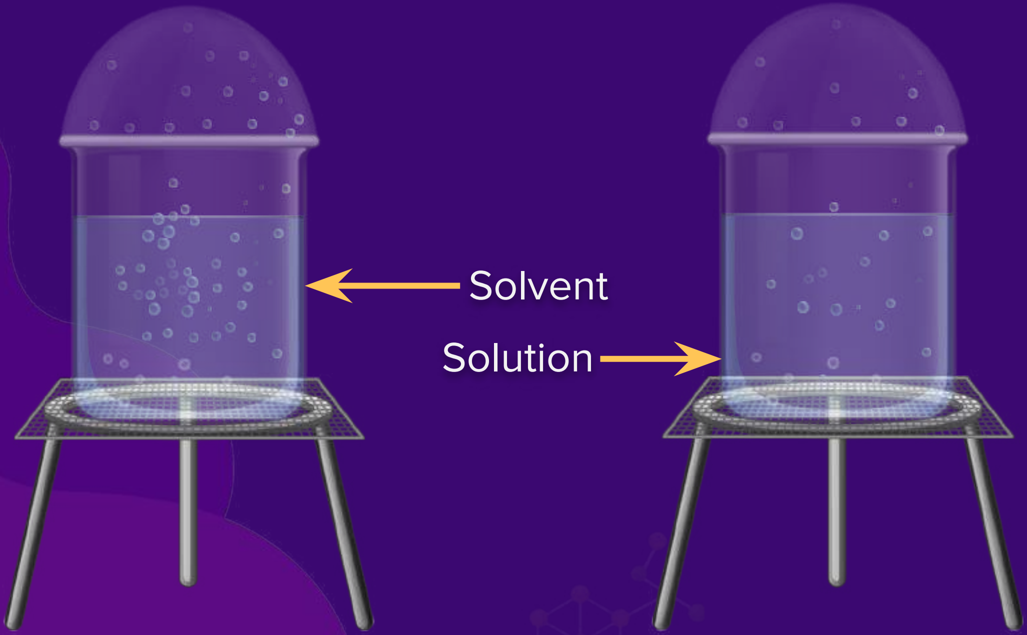
Reason

Some of the solute molecules will **occupy** some surface area of the solution.

Tendency of the solvent particles to go into the vapour phase **slightly decreases**.

Hence, $\mathbf{P^0 > P_s}$, where P^0 is V.P. of the pure solvent and P_s is the V.P. of the solution.

Relative Lowering of Vapour Pressure



Relative Lowering of Vapour Pressure

Lowering in V.P. $= P^0 - P_s = \Delta P$

Relative lowering in vapour pressure $= \dfrac{\Delta P}{P^0}$

Relative Lowering of Vapour Pressure

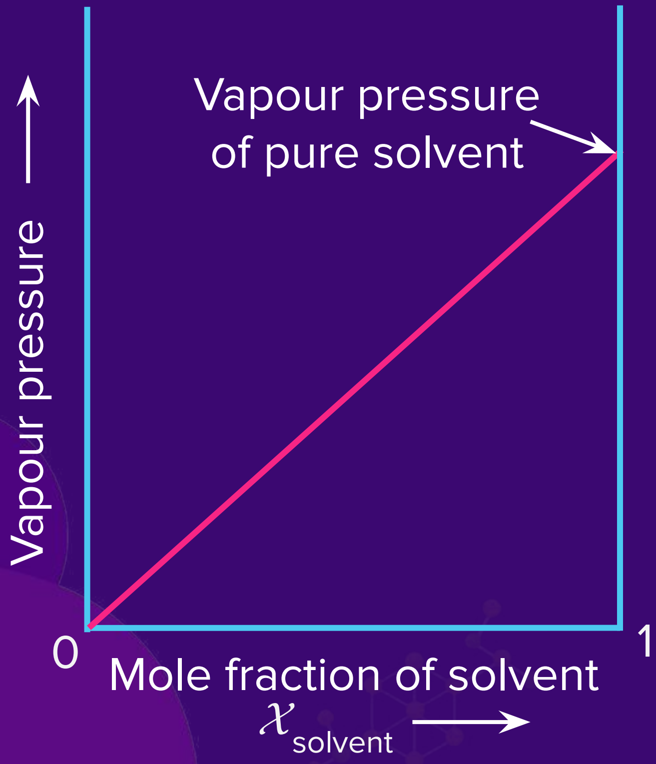
Relative lowering of vapour pressure is a **colligative property** whereas, lowering in vapour pressure is not.

Relative Lowering of Vapour Pressure

Raoult’s Law (For non–volatile solutes)

The vapour pressure of a solution of a non-volatile solute is equal to the vapour pressure of the pure solvent at that temperature multiplied by its **mole fraction**.

Relative Lowering of Vapour Pressure



Relative Lowering of Vapour Pressure

$$P_s = \chi_{solvent} P^0 = (1 - \chi_{solute}) P^0$$

$$RLVP = \frac{P^0 - P_s}{P^0} = \chi_{solute} = \frac{n}{n + N}$$

n = Number of moles of non-volatile solute
N = Number of moles of solvent in the solution

Relative Lowering of Vapour Pressure

$$\frac{P^0}{P^0 - P_s} = \frac{n + N}{n} = 1 + \frac{N}{n}$$

$$\frac{N}{n} = \frac{P^0}{P^0 - P_s} - 1 = \frac{P_s}{P^0 - P_s}$$

$$\frac{P^0 - P_s}{P_s} = \frac{n}{N}$$

Relative Lowering of Vapour Pressure

$$\frac{P^0 - P_S}{P_S} = \frac{w_{solute}}{M_{solute}} \times \frac{M_{solvent}}{w_{solvent}}$$

$$= \frac{w_{solute}}{M_{solute}} \times \frac{M_{solvent}}{w_{solvent}} \times \frac{1000}{1000}$$

w = Weight of species
M = Molar mass of species

Relative Lowering of Vapour Pressure

$$\frac{P^0 - P_s}{P_s} = \frac{w_{solute}}{M_{solute}} \times \frac{1000}{w_{solvent}} \times \frac{M_{solvent}}{1000}$$

$$\frac{P^0 - P_s}{P_s} = \text{Molality} \times \frac{M_{solvent}}{1000}$$

Relative Lowering of Vapour Pressure

If solute gets **associated or dissociated**

$$\frac{P^0 - P_s}{P^0} = \frac{i.n}{i.n + N}$$

$$\frac{P^0 - P_s}{P_s} = \frac{i.n}{N} = i \times \text{Molality} \times \frac{M_{\text{solvent}}}{1000}$$

The **weight of urea** which must be dissolved in **400 g** of water. Therefore, the final solution has vapour pressure **2%** less than the vapour pressure of pure water is:

BOARDS

Solution

- **a** 54.4 g
- **b** 27.2 g
- **c** 13.6 g
- **d** 6.8 g

Hence, option (b) is the correct answer.

Which of the following salts has the **same value** of van’t Hoff factor (i) as that of $K_3[Fe(CN)_6]$?

Solution

a Na_2SO_4

b $Al(NO_3)_3$

c $Al_2(SO_4)_3$

d $NaCl$

Hence, option (b) is the correct answer.

The **van't Hoff factor (i)** for a dilute aqueous solution of the strong electrolyte barium hydroxide is:

BOARDS

Solution

- **a** 0
- **b** 1
- **c** 2
- **d** 3

Hence, option (d) is the correct answer.

The vapour pressure of pure benzene at a certain temperature is 0.850 bar. A non-volatile, non-electrolyte solid weighing 0.5 g is added to 39.0 g of benzene (molar mass 78 g mol^{-1}). The vapour pressure of the solution then is 0.845 bar. What is the approximate value of molecular mass of the solid substance?

BOARDS

Solution

a 58 u

b 180 u

c 170 u

d 145 u

Hence, option (c) is the correct answer.

The vapour pressure of a solvent decreased by 10 *mm* of mercury when a non-volatile solute was added to the solvent. The mole fraction of the solute in the solution is 0.2 . What should be the mole fraction of the solvent if the decrease in the vapour pressure is to be 20 *mm* of mercury ?

BOARDS

$\therefore x_{v_1} = 1 - x_{u_1} = 1 - 0.2 = 0.8$

$p^\circ - p_1 = 10\,mmHg \quad x_{u_1} = 0.2 \quad Soln-1$

$p^\circ - p_2 = 20\,mmHg \quad x_{v_2} = ? \quad Soln-2$

Solution

- **a** 0.4
- **b** 0.6
- **c** 0.8
- **d** 0.2

$$\frac{\left(\frac{p^\circ - p_1}{\cancel{p^\circ}}\right)}{\left(\frac{p^\circ - p_2}{\cancel{p^\circ}}\right)} = \frac{1 - x_{v_1}}{1 - x_{v_2}}$$

$1 - x_{v_2} = 0.4 \quad \therefore x_{v_2} = 1 - 0.4 = \underline{0.6}$

Hence, option (b) is the correct answer.

The vapour pressure of CCl_4 at 25 °C is 143 mm Hg. If 0.5 g of a non-volatile solute (mol. weight = 65) is dissolved in 100 g CCl_4, the vapour pressure of the solution will be:

Solution

a	199.34 mm Hg
b	143.99 mm Hg
c	141.43 mm Hg
d	94.39 mm Hg

Hence, option (c) is the correct answer.

The relative lowering of the vapour pressure is equal to the ratio between the number of:

BOARDS

a Solute molecules to the solvent molecules.

b Solute molecules to the total molecules in the solution.

c Solvent molecules to the total molecules in the solution.

d Solvent molecules to the total number of ions of the solute.

Solution

Relative lowering of vapour pressure for dilute solutions:

$$\frac{p^{\circ}-p_s}{p^{\circ}} \boxed{\approx \frac{n_u}{n_v}} = \frac{n_u}{(n_u+n_v)}$$

So, relative lowering of the vapour pressure is equal to the ratio between the number of solute molecules to the total molecules in the solution

Hence, option (b) is the correct answer.

The vapour pressure of pure benzene, C_6H_6 at 50 °C is **268 Torr**. **How many moles** of non-volatile solute per moles of benzene is required to prepare a solution of benzene having a vapour pressure of **167 Torr** at 50 °C?

BOARDS

Solution

a 0.377

b 0.605

c 0.623

d 0.395

Hence, option (b) is the correct answer.

If P^0 and P are the vapour pressures of a solvent and its solution respectively and N_1 and N_2 are the mole fractions of the solvent and non-volatile solute respectively, then correct relation is:

Solution

a $P = P^0 N_2$

b $P = P^0 N_1$

c $P^0 = P N_1$

d $P = P^0 N_1 / N_2$

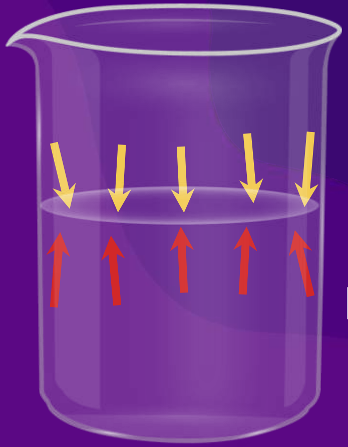
Hence, option (b) is the correct answer.

Colligative Properties

- Relative lowering of vapour pressure
- ✓ Elevation in boiling point
- Depression in freezing point
- Osmotic pressure

Elevation in Boiling Point

Temperature at which the **V.P.** of a liquid is **equal** to the **external pressure** present at the surface of the liquid.



External pressure

Vapour pressure

At **normal B.P.**, V.P. of the pure liquid = 1 atm

In a pressure cooker pressure increases , hence boiling point of water increases and now rice is cooked in more heat as water boils at higher temperature. So now rice will cook faster.

Elevation in Boiling Point

When a **non-volatile solute is added** into a volatile liquid to form solution, V.P. decreases.

The solution need to be **heated to a higher temperature** to boil it, so that V.P. becomes equal to external pressure.

Elevation in Boiling Point

For example,

Vapour pressure of an aqueous solution of **sucrose** is **less than 1.013 bar** at 373.15 K.

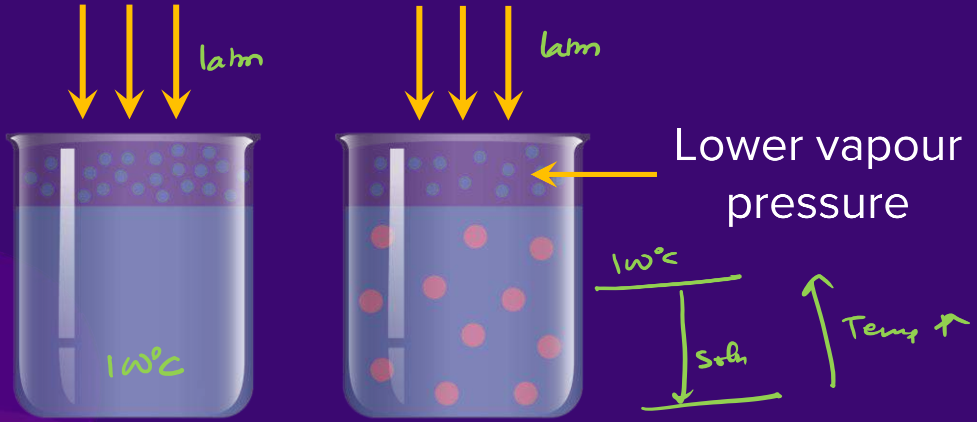
= 1 atm = 100°C

To make this solution boil, its vapour pressure must be **increased to 1.013 bar** by raising the temperature above the boiling temperature of the pure solvent (water).

Thus,

The boiling point of a solution ✓ > The boiling point of the pure solvent

Elevation in Boiling Point



Atmospheric pressure

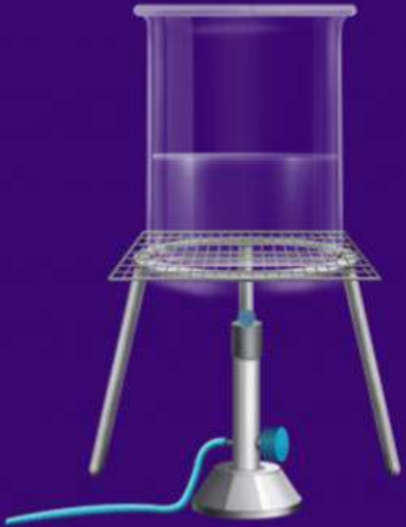
Atmospheric pressure

Lower vapour pressure

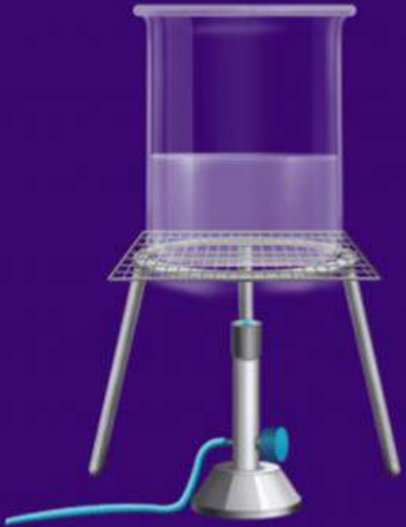
Pure solvent

Solution (with solute)

Elevation in Boiling Point



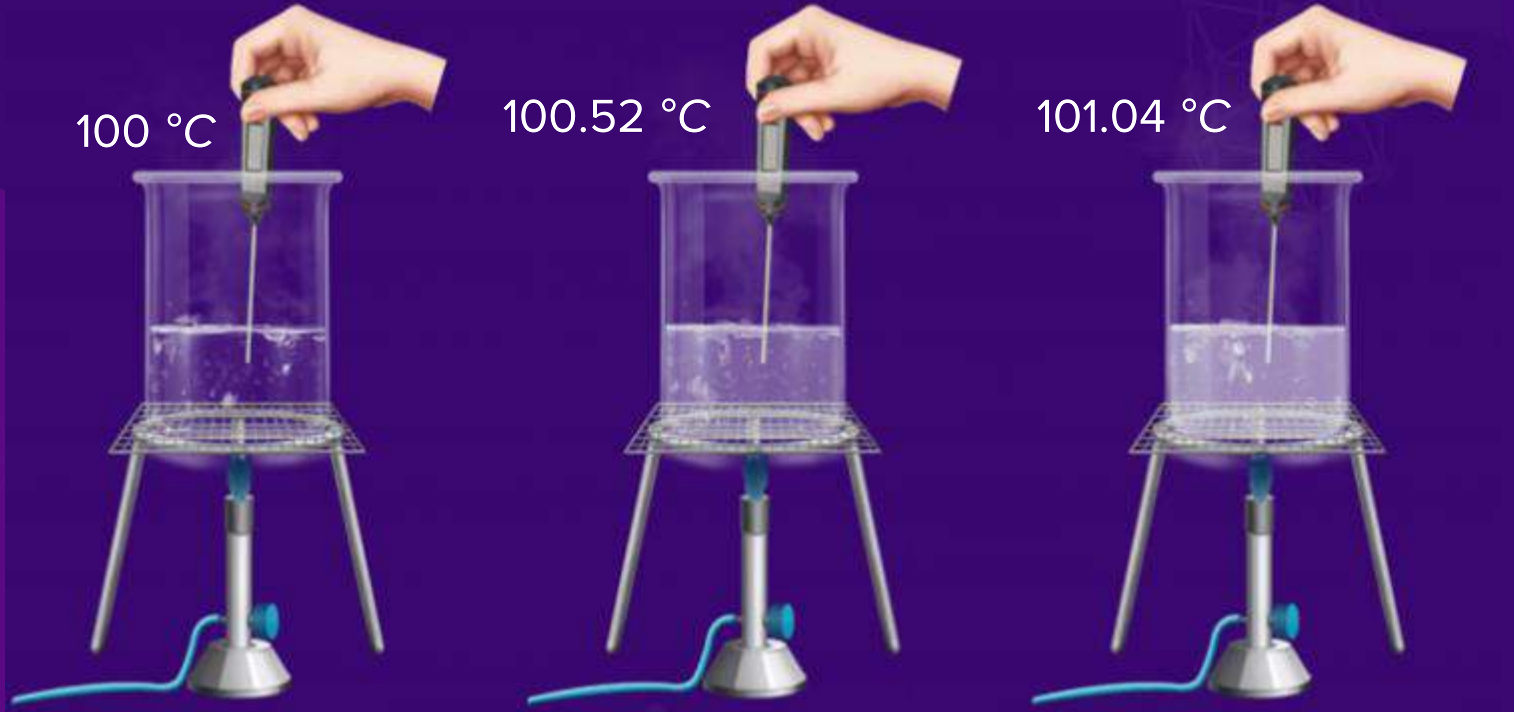
Pure water



1 mole of non-volatile solute
+
Water

2 mole of non-volatile solute
+
Water

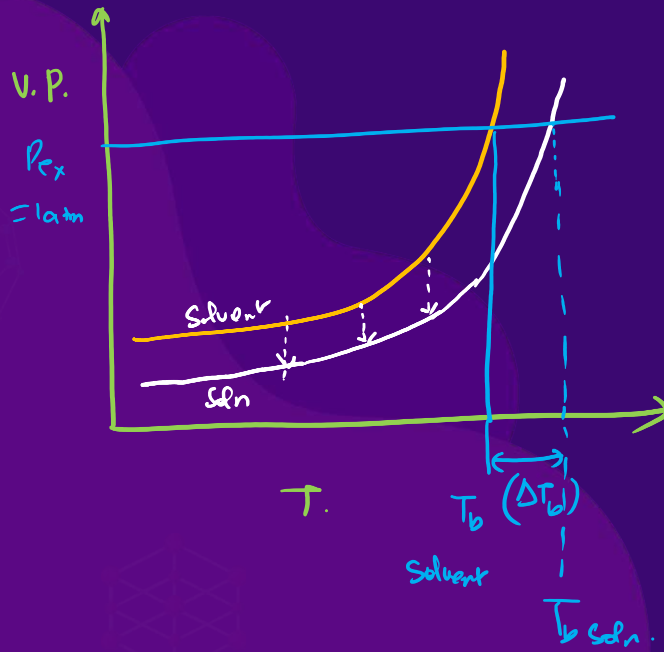
Elevation in Boiling Point

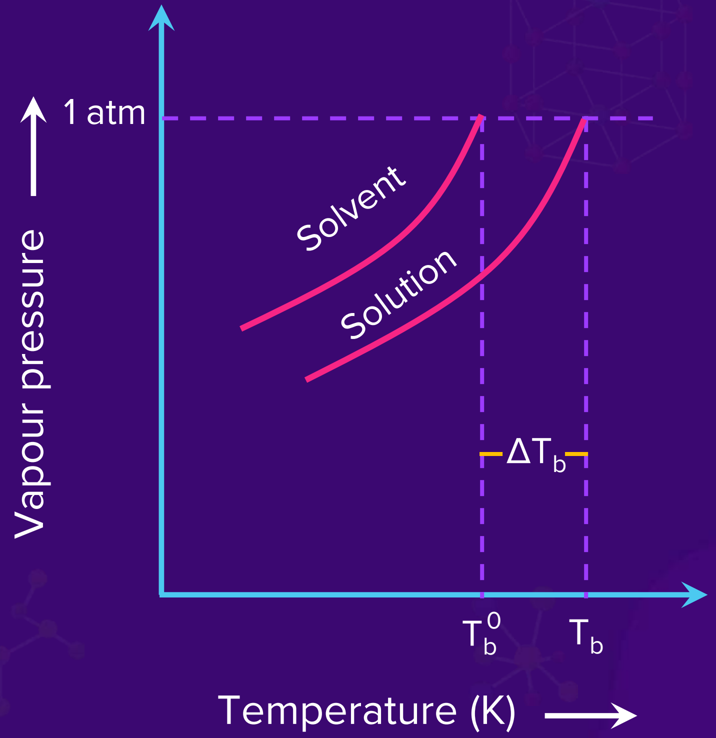


Hence, we can say that because of the addition of non-volatile solute the boiling point of solution increases

Elevation in Boiling Point

T_b = B.P. of solution (K)
T_b^0 = B.P. of pure solvent (K)





Elevation in Boiling Point

V.P. of solution < V.P. of pure solvent

Hence, to make the **V.P. equal to P_{ext}**, we have to heat the solution by a **greater amount** in comparison to pure solvent.

Elevation in Boiling Point

T_b(soln) − T_b(pure solvent)

$$\Delta T_b = T_b - T_b^0$$

$\Delta T_b \propto m \quad \therefore \quad \Delta T_b = K_b \cdot m$

$$\Delta T_b \propto m$$

m Molality

$m = m_0 \cdot i$

↓

molality calculated assuming no assoc^n/dissoc^n.

molality calculated after all possible assoc^n/ diss

$$\boxed{\Delta T_b = K_b\, m = K_b \cdot (m_0 \cdot i)}$$

K_b = B.P. elevation constant

or, Molal elevation constant

or, Ebullioscopic constant

Ebullioscopic Constant (K_b)

K_b

It is **equal to** elevation in boiling point of 1 molal solution.

Units

K/m or °C/m
or K kg mol^{-1}

Ebullioscopic Constant (K_b)

$$K_b = \frac{R{T_b^0}^2 M}{1000 \times \Delta H_{vap}}$$

- T_b^0 is B.P. of pure solvent (K)
- M is the molar mass of solvent in g/mol
- ΔH_{vap} is the molar enthalpy of vapourisation of the solvent (J/mol)
- R = 8.314 J/mol-K

Ebullioscopic Constant (K_b)

1 If solute gets associated/dissociated then **$\Delta T_b = i \times K_b \times$ Molality**

2 K_b is the property of solvent

Ebullioscopic Constant (K_b)

$\Delta T_b = K_b \cdot m$

$\Delta T_b \propto m$

$\frac{p^\circ - p}{p^\circ} \simeq m \times M_v \times 10^{-3}$

$\frac{p^\circ - p}{\boxed{p^\circ}} \propto m$

3 Elevation in B.P. is proportional to the **lowering of vapour pressure.**

$\underbrace{(p^\circ - p)}_{(\Delta P)} \propto m$

$(\Delta P) \propto m$

$(\Delta T)_b \propto m$

$$\boxed{\Delta T_b \propto \Delta P}$$

for dilute soln

At 100 °C the vapour pressure of a solution of 6.5 g of a solute in 100 g water is 732 mm. If K_b = 0.52 K kg mol^{-1}, the boiling point of this solution will be:

at 100°C P^o_w = 760 torr

Solution

$\Delta T_b = K_b \cdot m$

?

$\frac{P^o - P}{P} = m \times M_w \times 10^{-3}$

$m = \frac{760-732}{732 \times 18 \times 10^{-3}}$

$= \frac{\Delta T_b}{K_b}$

$\therefore \Delta T_b = \left(\frac{28}{732 \times 18 \times 10^{-3}} \times .52\right)$

$= \frac{\cancel{28}^{14} \times \cancel{520}^{130}}{{}_{9}\cancel{18} \times \cancel{732}_{183}}$

$= 1.1 \ldots$

$100 + 1.1$
$= 101.1.. °C$

a 102°C

b 103°C

[✓ c 101°C]

d 100°C

$\frac{\overset{4}{14} \times 130}{9 \times \overset{72}{183}}$

$= \frac{1820}{1647}$

$= \frac{1647 + 173}{1647}$

$= 1 + \frac{173}{1647}$

$= 1 + \frac{164.7 + 8.3}{1647}$

Hence, option (c) is the correct answer.

If 0.15 g of a solute, dissolved in 15 g of solvent, is boiled at a temperature higher by **0.216°C,** than that of the pure solvent. The **molecular weight** of the substance (Molal elevation constant for the solvent is 2.16 °C kg mol^{-1}) is:

$m_u = 0.15 g$ $\Delta T_b = 0.216 °c$ $M_u = ?$

$m_v = \underline{\underline{15}} g$ $K_b = 2.16 °C/molal$

Solution

$$\cancel{0.216} = \overset{10}{\cancel{2.16}} \times \frac{(0.15/M_u)}{15 \times 10^{-3}}$$

$$\therefore M_u = \frac{0.15 \times 10^{3} \times 10}{15}$$

$$= \underline{\underline{100}}$$

a 10.1

[✓ b] 100

c 1.01

d 1000

Hence, option (b) is the correct answer.

Colligative Properties

- Relative lowering of vapour pressure
- Elevation in boiling point
- Depression in freezing point
- Osmotic pressure

Freezing Point (F.P.)

Temperature at which, the vapour pressure of a solid becomes **equal to** the vapour pressure of liquid at 1 atm is called **normal freezing point**.

V.P. of solid = V.P. of liquid

Depression in Freezing Point

When a non-volatile solute is dissolved in the solvent, the **V.P. of the solvent** in the solution **decreases**.

V.P. of solid and liquid solvent will become equal at a **lower temperature**, i.e., F.P. of solution is **lower than that of a pure solvent.**

Depression in Freezing Point

ΔT_f

The **difference** between F.P. of a pure solvent **T_f^0** and F.P. of its solution **T_f**.

Depression in Freezing Point

$\Delta T_f = T_f^0 - T_f$

$\Delta T_f \propto m$

m Molality

$\Delta T_f = K_f m$

K_f = F.P. depression constant
or molal depression constant
or cryoscopic constant.

Depression in Freezing Point

K_f

It is **equal to** the depression in freezing point of 1 molal solution.

Units K/m or °C/m or K kg mol^{-1}

$$K_f = \frac{RT_f^{0\,2}M}{1000 \times \Delta H_{fus}}$$

- T_f^0 is F.P. of pure solvent (K)
- M is the molar mass of solvent in g/mol
- ΔH_{fus} is the molar enthalpy of fusion of the solvent (J/mol)
- R = 8.314 J/mol-K

Depression in Freezing Point

1. Depression in freezing point is proportional to the lowering of vapour pressure i.e. **$\Delta T_f \propto \Delta P$**

2. If solute gets associated/dissociated then **$\Delta T_f = i \times K_f \times \text{Molality}$**

3. At freezing point or below it, **only solvent molecules will freeze** & not solute molecules.

Colligative Properties

Colligative Properties

- Relative lowering in vapour pressure

$$\frac{P^{o} - P_s}{P_s} = \frac{n_u}{n_v} = m \times M_v \times 10^{-3}$$

Where, m = molality, M_v = mass of solvent

- Depression in freezing point

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

$$K_f = \frac{RT_f^{o\,2}\,M}{1000 \times \Delta H_{fus}}$$

- Elevation in boiling point

$$\Delta T_b = K_b \times m$$

$$K_b = \frac{RT_b^{o\,2}\,M}{1000 \times \Delta H_{vap}}$$

Applications of Depression in Freezing Point

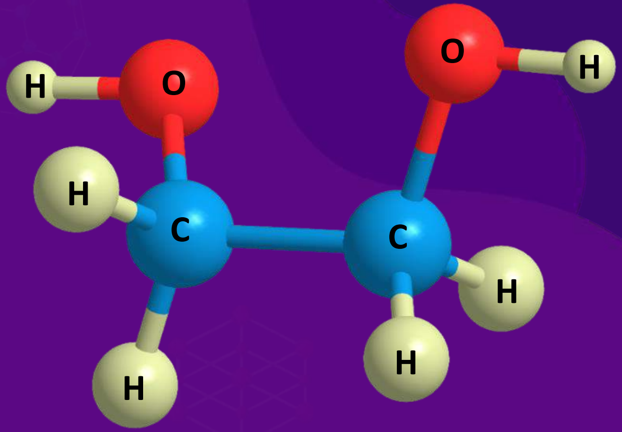
Salting roads in winter to reduce the freezing point of water.

Roads are salted in the winter to reduce the freezing point of water, so that the removal of ice from roads and footpaths becomes easy. Water normally freezes at 0°C. On adding salt (i.e., a solute), the freezing point of water is reduced to, say around, –3 °C (this value is taken for understanding purpose).

Thus, the ice formed will melt at any temperature above –3 °C. Hence, the snow on the road can be easily removed as it will start to melt when the temperature is higher than –3 °C.

Applications of Depression in Freezing Point

Ethylene glycol prevents our car's engine from freezing in the winter and acts as a coolant to reduce overheating in the summer.



Ethylene glycol is a chemical commonly used in many commercial and industrial applications including antifreeze and coolant.

The freezing point depression constant (K_f) of benzene is 5.12 K kg mol^{-1}. The freezing point depression for the solution of molality 0.078 m containing a non-electrolyte solute in benzene is (rounded off upto two decimal places).

$i = 1$

$K_f = 5.12 \ K \cdot kg \cdot mol^{-1}$

Solution

$$\boxed{-\Delta T_f} = m \cdot K_f$$

$$= 5.12 \times 0.078$$

$$\approx 5 \times 0.080 \approx 0.4$$

- a 0.2 K
- [✓ b] 0.40 K
- c 0.80 K
- d 0.60 K

Hence, option (b) is the correct answer.

If molality of the dilute solution is doubled, the value of molal depression constant (K_f) will be:

BOARDS

- a Halved
- b Tripled
- c Unchanged
- d Doubled

Solution

$\Delta T_f = K_f \times m$

If molality is doubled, K_f will remain unchanged since its a constant and it only depends on solvent.

Hence, option (c) is the correct answer.

A solution of sucrose (molar mass = 342 g mol⁻¹) has been prepared by dissolving 68.5 g of sucrose in 1000 g of water. The freezing point of the solution obtained will be:
(K_f for water = 1.86 K kg mol⁻¹)

moles of solute = $\frac{68.5\,g}{342\,g/mol}$

Solution

$-\Delta T_f = K_f\,m = 1.86 \times \frac{68.5\ \text{mol}}{342}$

$\text{K}\cdot\cancel{\text{kg}}\cdot\text{mol}^{-1}$ $1\,\text{kg}$

$-[T_f(\text{soln}) - \overset{0^\circ C}{T_{\text{solvent}}}]$

$= 1.86 \times \frac{68.5}{342}$

$\therefore T_f(\text{soln}) = -1.86 \times \frac{(68.5)}{(342)}$

$\approx -1.86 \times \frac{\frac{2}{3} \times \cancel{1000}}{\cdot 3413 \times 1000}$

$= -1.86 \times \frac{\frac{2}{\cancel{3}} \times \cancel{1000}}{(\cancel{1000}/\cancel{3})} = -.186 \times 2$

$= -.372$

- a — 0.372°C ✓
- b — 0.520°C
- c + 0.372°C
- d — 0.570°C

$\frac{1}{9} = .1111$
$\frac{1}{10} = .1$
$\frac{1}{1} = 1$
$\frac{1}{2} = 0.5$
$\frac{1}{3} = .333$
$\frac{1}{4} = .25$
$\frac{1}{5} = .20$
$\frac{1}{6} = .1666$
$\frac{1}{7} = .1414$
$\frac{1}{8} = .125$

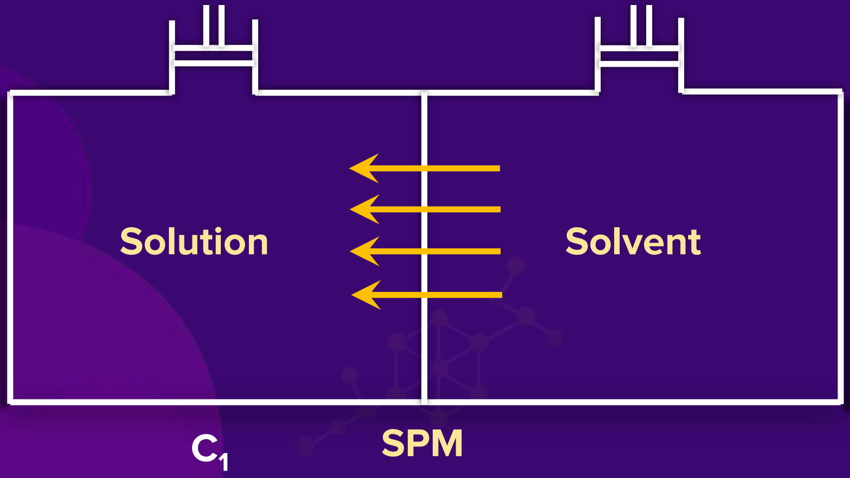
Hence, option (a) is the correct answer.

Colligative Properties

- Relative lowering of vapour pressure
- Elevation in boiling point
- Depression in freezing point
- Osmotic pressure

Osmosis

The **spontaneous flow** of solvent particles from **solvent side to solution side,** or from solution of low concentration side to solution of high concentration side through a **semi-permeable membrane (SPM)**.



Osmosis

Semipermeable membrane (SPM)

- A membrane, which allows only solvent particles to pass through it is known as a semipermeable membrane.
- The submicroscopic holes or pores present in the membrane are designed in such a way that only solvent particles can be passed through these pores.
- The transfer of bigger solute molecules through SPM is hindered.

Osmosis

Illustration of SPM:

Blue coloured particles represent solvent molecules and the red coloured particles represent the solute molecules in the figure shown. Since the solvent molecules are smaller than the solute molecules, they can pass through the holes in SPM. The solute particles being bigger in size, cannot pass through the holes in SPM.

Solvent particles can pass through a semipermeable membrane, but solute particles cannot pass through a semipermeable membrane.

Semi-permeable Membrane (SPM)

SPM

Natural

Artificial

Natural : Animal/plant cell membrane formed just below the outer skins.

Artificial SPM : $Cu_2[Fe(CN)_6]$ & Silicate of Ni, Fe, Co can act as SPM.

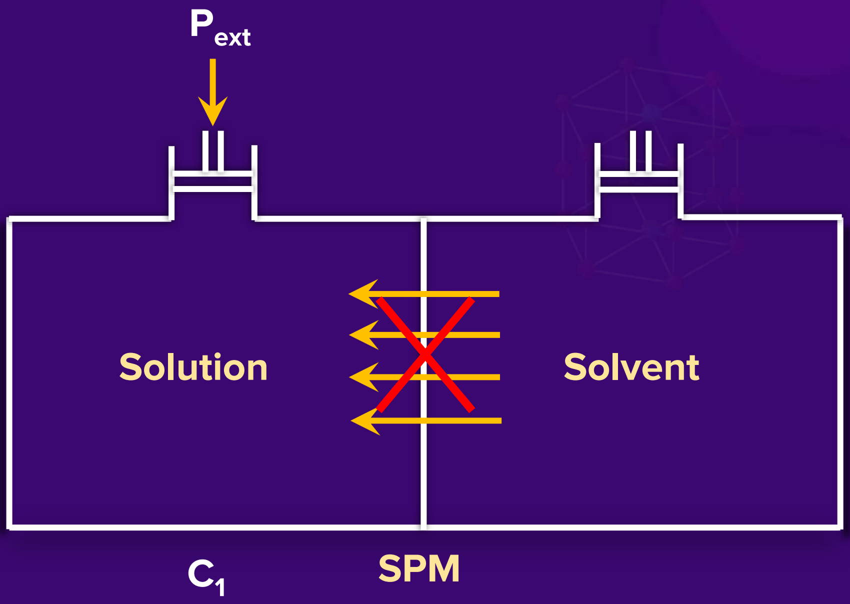
Example of Osmosis

- i A **raw mango** placed in a concentrated salt solution loses water & **shrivel** into pickle.
- ii People taking a lot of salt, experience **water retention** in tissue cells. This results in puffiness or swelling called **edema**.

Osmotic Pressure (π)

The external pressure that must be applied on the solution side to just stop the process of osmosis.

If solvent and solution are in contact by SPM then osmosis will take place. Now in order to avoid this osmosis P_{ext} is applied on the solution side. The flow of solvent molecules will not take place and hence osmosis will stop.



$$P_{ext} = \pi$$

Osmotic Pressure (π)

The extra pressure that must be applied to the solution which is in contact with its pure solvent through a semi-permeable membrane in order to just stop osmosis. This is called the **osmotic pressure (π)** of that solution.

- Osmotic pressure is a colligative property as it depends on the number of solute molecules and not on their identity or nature.
- The osmotic pressure has been found to depend on concentration of the solution and temperature.

For dilute solutions, it has been found experimentally that osmotic pressure is proportional to the molarity, C of the solution at a given temperature T.

Thus,

$$\boldsymbol{\pi \propto C \text{ (molarity) and } \pi \propto T}$$

Osmotic Pressure (π)

Hence,

$$\pi = CRT$$

Where, π = Osmotic pressure, C = Concentration ($mol\ L^{-1}$), R = Universal gas constant ($L\ atm\ mol^{-1}\ K^{-1}$) and T = Temperature (K)

Concentration can be expressed as C = n/V, where n is the number of moles of the solute and V is the volume of the solution. Substituting C = n/V in the equation for osmotic pressure,

$$\pi = \frac{n}{V}RT$$

Osmotic Pressure (π)

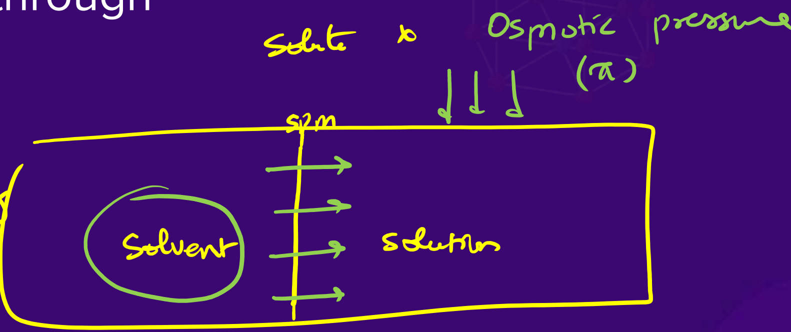
When a semi-permeable membrane is placed between a solution and solvent, solvent particles can move across the membrane but solute particles can not pass through membrane.

solvent ✓

solute ✗

Osmosis : The spontaneous flow of solvent particles from solvent side to solution side, or from solution of low concentration side to solution of high concentration side through a semipermeable membrane (SPM)

π colligative property



Osmotic Pressure (π) : The extra pressure that must be applied to the solution which is in contact with its pure solvent through a semipermeable membrane in order to just stop osmosis. This is called the **osmotic pressure (π)** of that solution

$$\pi = CRT$$

$$\pi = \frac{\text{moles of solute}}{\text{vol. of soln (in L)}} RT$$

$$\pi \propto n_{solute}$$

van't – Hoff Formula

If **more than one** type of solute particles are present.

$$C_T = C_1 + C_2 + C_3 + \ldots\ldots\ldots$$

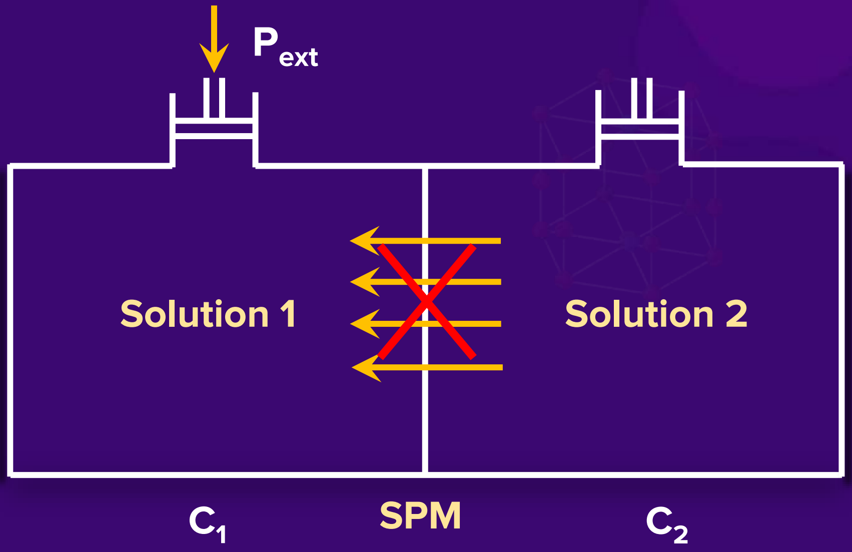
$$= \frac{n_1 + n_2 + n_3 + \ldots\ldots\ldots}{V}$$

$$\pi = C_T RT$$

Osmotic Pressure (π)

If two solutions of concentration C_1 & C_2 are kept separated by SPM, and **$C_1 > C_2$,** then the **solvent particle movement** take place from **lower to higher** concentration.

So, an **extra pressure** is applied on the higher concentration side to **stop osmosis**.



$C_1 > C_2$

$$P_{ext} = \pi_1 - \pi_2$$

Osmotic Pressure (π)

If solute gets associated or dissociated then **$\Pi = i \times CRT$**

Osmotic pressure of very dilute solutions is also **quite significant**. So, its **measurement in lab is very easy**.

During osmosis, flow of water through a semipermeable membrane is:

BOARDS

a From solution having lower concentration only.

b From solution having higher concentration only.

c From both sides of semipermeable membrane with equal flow rates.

d From both sides of semipermeable membrane with unequal flow rates.

Solution

During osmosis, flow is from both sides of semipermeable membrane with unequal flow rates.

Hence, option (d) is the correct answer.

0.450 g of urea (mol. wt. = 60 u) in 22.5 g of water show 0.170 °C of elevation in boiling point. The ~~molal~~ elevation constant of water is:

Solution

$n_{urea} = \frac{m}{M} = \frac{0.45}{60} = \frac{.03}{4}$ mol

$\Delta T_b = K_b \cdot m$

$0.17 = K_b \frac{\left[\frac{.03}{4}\right]}{22.5\times10^{-3}}$

$K_b = \frac{0.17 \times 22.5 \times 10^{-3} \times 4}{.03}$

$= \frac{0.17 \times 2.25 \times 4}{3}$

$= \frac{0.17 \times \cancel{9}\,3}{3}$

$= 0.51$

- **a** 0.051 °C kg mol^{-1}
- [✓ **b**] 0.51 °C kg mol^{-1}
- **c** 5.1 °C kg mol^{-1}
- **d** 0.83 °C kg mol^{-1}

Hence, option (b) is the correct answer.

On mixing 3 g of non-volatile solute in 200 mL of water, its boiling point (100°C) becomes 100.52°C. If K_b for water is 0.6 K kg mol^{-1} then **molar wt. of solute** is:

Solution

$\Delta T_b = 100.52 - 100 = 0.52°C = K_b \cdot m$

→ 200 mL H_2O

$\equiv 200 g\ H_2O \equiv 0.2 kg\ H_2O$

$0.52 = 0.6 \times \frac{(3/M_u)}{0.2 kg}$

$\therefore M_u = \frac{3 \times \cancel{0}.6}{0.52 \times \cancel{0}.2}$

$= \frac{18}{(1.04)}$

$\lesssim \underline{\underline{18}}$

- a 10.5 g mol^{-1}
- b 12.6 g mol^{-1}
- c 15.7 g mol^{-1}
- [✓ d 17.3 g mol^{-1}]

Hence, option (d) is the correct answer.

$C_{12}H_{22}O_{11}$

A 5% solution (by mass) of cane sugar in water has freezing point of 271 K and freezing point of pure water is 273.15 K. The **freezing point** of a 5% solution (by mass) of glucose in water is:

$\left(\frac{5g}{M_u}\right)$

5% $\quad -\Delta T_f = 2.15K \quad$ Cane sugar

5% $\quad$ glucose ?

$(-\Delta T_f)_g = \frac{342}{180} \times 2.15$

AIIMS 2006

Solution

100 g solution

Solute 5g $= (5/M_u)$

soln. 95g $= (95 \times 10^{-3}\,kg)$

$$\boxed{(-\Delta T_f) = K_f \cdot m}$$

$$(-\Delta T_f) = \frac{K_f \times (5/M_u)}{(95 \times 10^{-3})}$$

- a ~~271 K~~
- b ~~273.15 K~~
- c 269.07 K
- d ~~277.23 K~~

$(-\Delta T_f)_s = \left(\frac{K_f \cdot 5}{95 \times 10^{-3}}\right) \frac{1}{M_s}$

$(-\Delta T_f)_g = \left(\frac{K_f \cdot 5}{95 \times 10^{-3}}\right) \cdot \frac{1}{M_g}$

$\therefore \frac{(-\Delta T_f)_s}{(-\Delta T_f)_g} = \frac{M_g}{M_{cs}}$

$\frac{2.15}{(-\Delta T_f)_g} = \frac{180}{342}$

Hence, option (c) is the correct answer.

If 0.1 M solution of glucose and 0.1 M solution of urea are placed on two sides of the semipermeable membrane to equal heights, then it will be correct to say that:

a There will be no net movement across the membrane.

b Glucose will flow towards glucose solution.

c Urea will flow towards glucose solution.

d Water will flow from urea solution to glucose.

Solution

π = CRT ; Since , concentration is same (Isotonic solution), so π will be same and there will be no net movement across the membrane.

Hence, option (a) is the correct answer.

Q1

0·1M glucose solution at 27°C (300K)

Find its osmotic pressure (π).

Given : $R = 0\cdot0821 \ \frac{L\cdot atm}{mol\cdot K} \approx \frac{1}{12} \ \frac{L\cdot atm}{mol\cdot K}$

Soln

Solution

$$\pi = CRT$$

$$= 0\cdot1 \frac{\cancel{mol}}{\cancel{L}} \times \frac{1}{12} \cdot \frac{\cancel{L}\cdot atm}{\cancel{mol}\cdot \cancel{K}} \times 300\,\cancel{K}$$

$$= \frac{30}{12}\ atm$$

$$= \boxed{2\cdot5\,atm}$$

The relationship between **osmotic pressure** at **273 K** when **10 g glucose** (P_1), **10 g urea** (P_2) and **10 g sucrose** (P_3) are dissolved in **250 mL** of water is:

180

60

342

$\pi = CRT = \frac{n}{V} RT$

$n = \frac{m}{M}$

urea : $NH_2-\overset{O}{C}-NH_2$: 60

glucose : $C_6H_{12}O_6$: 180

Sucrose : $C_{12}H_{22}O_{11}$: 342

Solution

$\pi \propto n$

$\pi \propto \frac{m}{M}$

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

$P_2 > P_1 > P_3$

- **a** $P_1 > P_2 > P_3$
- **b** $P_3 > P_1 > P_2$
- [✓ **c** $P_2 > P_1 > P_3$]
- **d** $P_2 > P_3 > P_1$

Hence, option (c) is the correct answer.

Osmotic pressure of blood is 7.40 atm at 27 °C. Number of moles of glucose to be used per litre for an intravenous injection that is to have the same osmotic pressure as blood is:

Solution

✓ $\pi = 7.4\ \text{atm}$ $T = 27^\circ C \equiv 300K$ $R = .0821 \frac{L\cdot atm}{mol\cdot K}$

$= CRT$ $C_{glucose} = ?$ soln of glucose $\approx \frac{1}{12}$

$$C = \frac{\pi}{RT} = \frac{7.4\ atm}{\frac{1}{12}\frac{L\cdot atm}{mol\cdot \cancel{K}} \times 300\cancel{K}}$$

$$= \frac{\overset{1}{7.4} \times \overset{4}{\cancel{12}}}{\underset{100}{\cancel{300}}} \frac{mol}{L}$$

$$= \frac{29.6}{100} = .296\ \underline{\underline{M}}$$

≈ 0.3

a 0.3

b 0.2

c 0.1

d 0.4

Hence, option (a) is the correct answer.

The total concentration of dissolved particles inside red blood cells is approximately **0.30 M** and the membrane surrounding the cells is **semipermeable**. What would the **osmotic pressure** (in atmosphere) inside the cells become if the cells were removed from the blood plasma and placed in pure water at **298 K**?

Solution

$\pi = CRT$

RBC $0.3M$

water

$= 0.3\frac{mol}{L} \times \frac{1}{12}\frac{L\,atm}{mol\cdot K} \times 298$

$= 0.3 \times \frac{1}{12} \times 300$

$= 3 \times \frac{\overset{2.5}{\cancel{30}}}{\cancel{12}} = 7.5$

$R \approx \frac{1}{12}\frac{L\cdot atm}{mol\cdot K}$

$= .0821$

- **a** 7.34 atm
- **b** 1.78 atm
- **c** 2.34 atm
- **d** 0.74 atm

Hence, option (a) is the correct answer.

Isotonic, Hypotonic, & Hypertonic Solutions

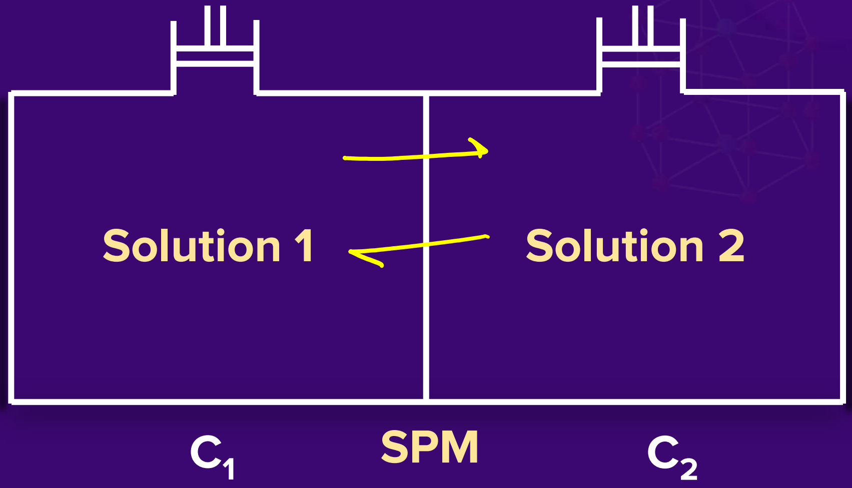
Types of Solutions

Based on the difference in osmotic pressure

1. **Isotonic** solution
2. **Hypotonic** solution
3. **Hypertonic** solution

Isotonic Solution

Two solutions having **same osmotic pressure** are considered as isotonic solution.



$\pi_1 = \pi_2$

Hypotonic & Hypertonic Solutions

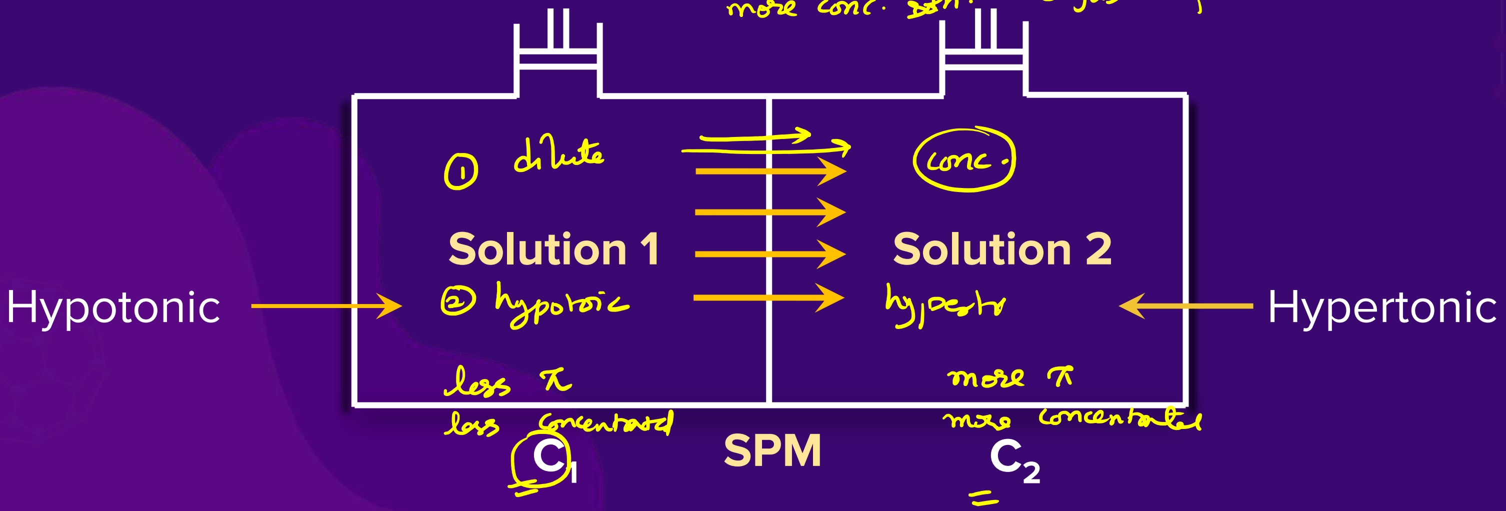
more π

less π

If two solutions ‘1’ and ‘2’ are such that **$\pi_2 > \pi_1$** , then **‘2’ is called hypertonic** solution and **‘1’ is called hypotonic** solution.

Hypotonic & Hypertonic Solutions

$\pi = (\pi_2 - \pi_1)$ must be applied on more conc. soln. to just stop osmosis.

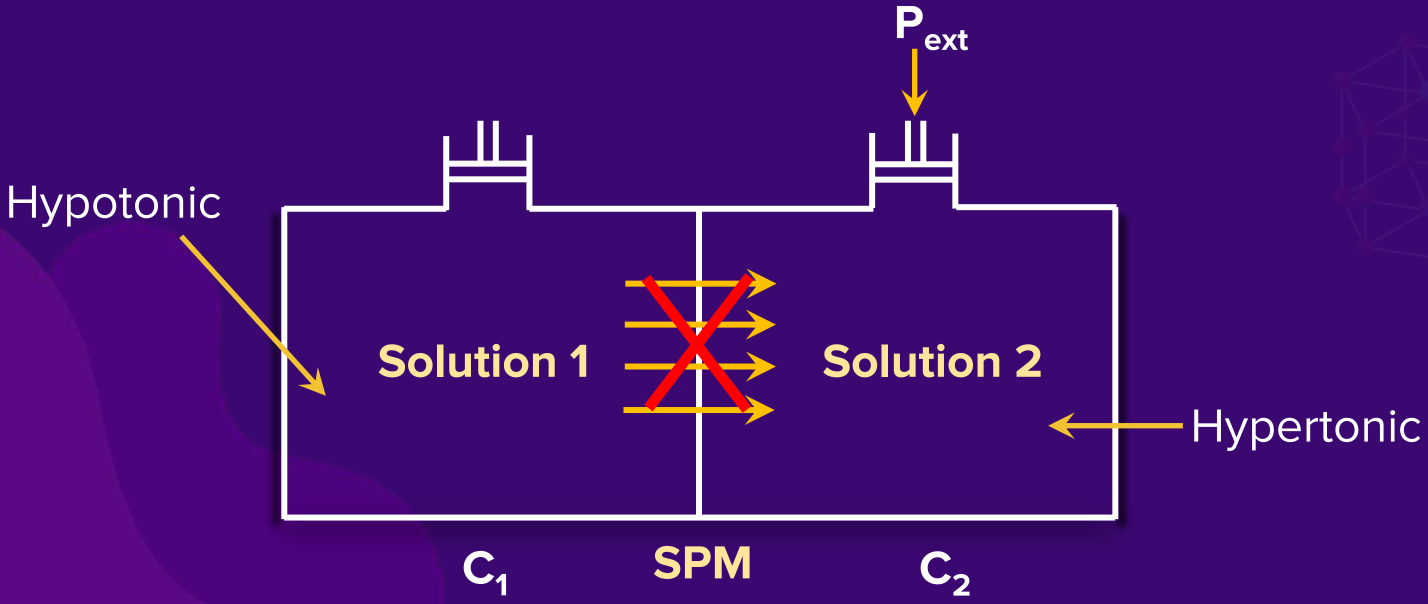


$\pi_1 < \pi_2$

Hypotonic & Hypertonic Solutions

Pressure is applied on the **hypertonic** solution to **stop the flow** of solvent particles.

Hypotonic & Hypertonic Solutions



$\pi_1 < \pi_2$ $\quad P_{ext} = \pi_2 - \pi_1$

Types of Solutions

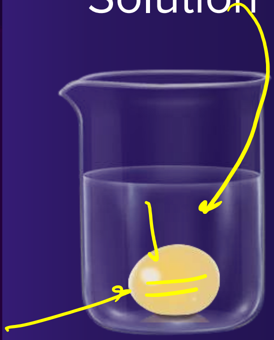
Illustration of hypotonic, hypertonic, and isotonic solutions

The white shell of an egg is carefully removed so that a semipermeable membrane is obtained on the egg.

- When this egg is placed in a solution of low concentration like corn syrup/vinegar/any solute, then the water flows into the egg and the egg swells. This solution is known as a hypotonic solution as shown in **case A**.
- When this egg is placed in a solution with equal concentration like water, no net loss or gain of water occurs. This solution is known as an isotonic solution, as shown in **case B.**
- When this egg is placed in a solution with higher concentration like boric acid, water will flow out of the egg, and the egg will shrink. Such a solution is known as a hypertonic solution as in **case C**.

Types of Solutions

Hypotonic Solution



Net water gain
Cell Swells

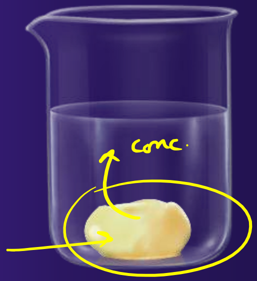
Case A

Isotonic Solution

No net
loss or gain

Case B

Hypertonic Solution



Net water loss
Cell Shrinks

Case C

Plasmolysis

When the cell is placed in a solution having its osmotic pressure greater than that of the cell sap, **water passes out of the cell** due to osmosis.

Consequently, the **cell material shrinks** gradually. The gradual shrinking of the cell material is called **plasmolysis**.

Plasmolysis

Why do mangoes **shrivel** into pickle in concentrated salt solution?

Because water from mangoes flow to concentrated salt solution because of osmosis.

Reverse Osmosis (R.O.)

If the **pressure applied** on the solution side is **more than the osmotic pressure** of the solution, then the solvent particles will move from **solution to solvent side**.

$P_{ext} > \pi$

- In R.O. filters high pressure is applied on the impure water so only pure water flows through SPM and collected.
- Used in **desalination of sea-water.**

At 27°C, a 1.2% solution (w/V) of glucose is isotonic with 4.0 g/litre of solution of solute X. Find the molar mass of X. (R = 0.082 L atm mol^{-1} K^{-1})

$M_u = 60$

100 mL soln → 1.2g glucose

1000mL → 12g glucose

1L

$C_{glucose} = \frac{(12/180)}{1L} = \frac{1}{15}$

$\pi_1 = \pi_2$

$C_1 \cancel{RT} = C_2 \cancel{RT}$

$C_1 = C_2$

$\frac{1}{15} M = \frac{\left(\frac{4}{M_u}\right)}{1L}$

$\therefore M_u = 15 \times 4$

$= 60$

Solution

- **a** 120 g/mol
- **b** 60 g/mol ✓
- **c** 180 g/mol
- **d** 90 g/mol

Hence, option (b) is the correct answer.

Application of Osmotic Pressure

a. Determination of **molecular mass** of the solute.

$$M_{solute} = \frac{wRT}{\Pi V}$$

- w is mass of solute
- V is volume of the solution
- Π is osmotic pressure of the solution
- T is temperature
- R is universal gas constant

Application of Osmotic Pressure

Widely used to **determine molar masses** of proteins and other biomolecules, as they are generally not stable at higher temperatures

This method has the advantage over other methods as pressure measurement is around the **room temperature**.

A solution contains non-volatile solute of molecular mass M_2. Which of the following can be used to calculate the **molecular mass of solute** in terms of osmotic pressure?
(m_2 = mass of solute, V = volume of solution, π = osmotic pressure)

Solution

$$M_2 = M_w = \frac{wRT}{\pi V} = \frac{m_2 RT}{\pi V}$$

a $M_2 = \left(\frac{m_2}{\pi}\right) VRT$

b $M_2 = \left(\frac{m_2}{V}\right) \frac{RT}{\pi}$

A solution contains non-volatile solute of molecular mass M_2. Which of the following can be used to calculate the **molecular mass of solute** in terms of osmotic pressure?

(m_2 = mass of solute, V = volume of solution, π = osmotic pressure)

c $M_2 = \left(\frac{m_2}{V}\right)\pi RT$

d $M_2 = \left(\frac{m_2}{V}\right)\frac{\pi}{RT}$

Hence, option (b) is the correct answer.

If **10 g** of an unknown substance (non-electrolytic) is dissolved to make **500 mL** of solution, then **osmotic pressure** at **300 K** is observed to be **1.23 atm.** Find **molecular weight** of solute.

- a 400 u ✓
- b 300 u
- c 500 u
- d 550 u

Solution

$W = 10g$

$V = 0.5L$

$\pi = 1.23atm$

$T = 300K$

$$M_w = \frac{wRT}{\pi V}$$

$$= \frac{10 \times \frac{1}{\cancel{12}} \times \overset{25}{\cancel{300}}}{1.23 \times 0.5}$$

$$= \frac{500}{1.23} \simeq 500 \times .8 = 400$$

$\frac{1}{8} = 0.125$

$\frac{1}{0.8} = 1.25$

Hence, option (a) is the correct answer.

Pure water can be obtained from sea water by:

BOARDS

- **a** Centrifugation
- **b** Plasmolysis
- **c** Reverse osmosis
- **d** Sedimentation

Solution

Reverse osmosis is used in desalination of seawater. Salts are removed by passing seawater through SPM.

Hence, option (c) is the correct answer.

From the colligative properties of solution, which one is the **best method** for the determination of molecular weight of proteins and polymers?

- a Osmotic pressure
- b Relative lowering in vapour pressure
- c Depression in freezing point
- d Elevation in boiling point

Solution

Osmotic pressure is widely used to determine molar masses of proteins and other biomolecules, as they are generally not stable at higher temperatures. This method has the advantage over other methods as pressure measurement is around the room temperature.

Hence, option (a) is the correct answer.

Application of Osmotic Pressure

b. Dialysis

Artificial kidney **removes waste** products from blood through osmosis.

- In dialysis selectively permeable membrane is used and it allows solvent as well as small solute particles but it does not allow big solute particles to pass through it such as RBC (red blood cells) and WBC (white blood cells). Dissolved impurities from blood such as urea can be passed through SPM and removed.