Solutions

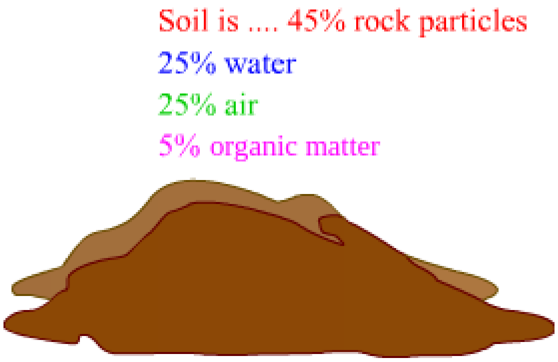
Solutions

Solutions-An Overview

Large amount of salt is dissolved in seawater. This makes it unfit for drinking directly. Can we say that the amount of salt in the sea is the same everywhere?

The air contains gases like oxygen, carbon dioxide and ozone along with various small particles like pollen grains and dust. Are the gases and the particles present in equal amounts in air?

Soil contains a lot of substances, e.g., clay, organic matter, minerals, pebbles, etc.

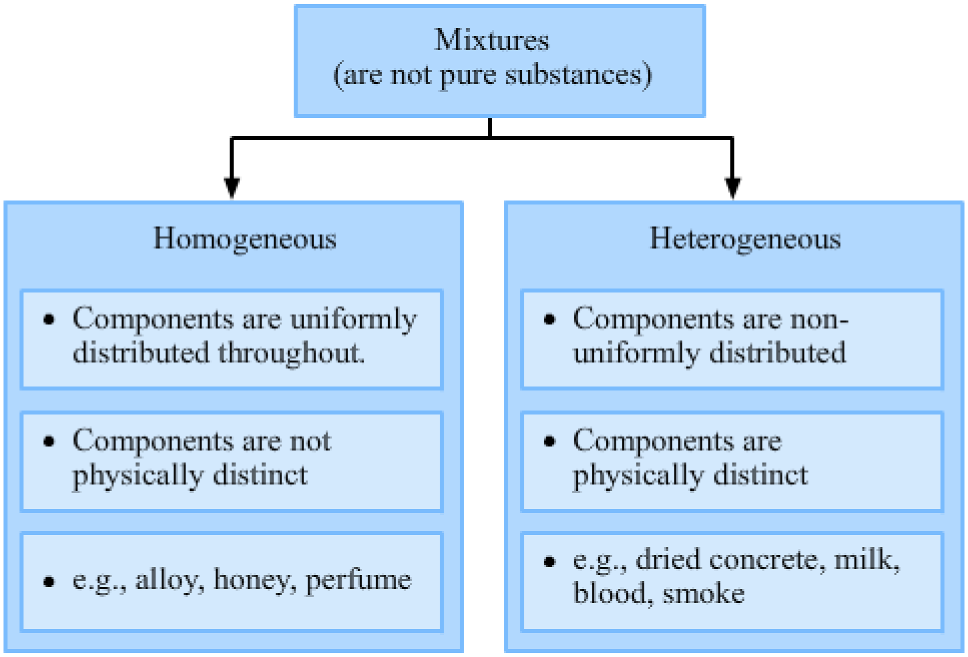


Why do the amounts of clay, organic matter, minerals, etc. in soil vary from place to place?

All of the above substances (soil, air and seawater) are examples of mixtures. Let us go through the lesson to find out what mixtures are.

Mixtures

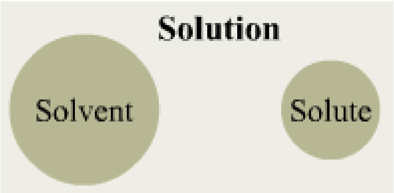
A mixture may be defined as a material having two or more types of pure forms of matter. For example, milk is a mixture as it contains a combination of water molecules, fat molecules and protein molecules. The constituents of a mixture can be separated by certain physical processes such as evaporation and boiling. Constituents of certain mixtures can also be separated manually. For example, a mixture of stones and sand can be separated manually. On the other hand, salt cannot be manually separated from saltwater. One needs to boil the mixture to separate the salt from water.



Mixtures

Solutions

Now that we know what mixtures are, let us study about solutions. Whenever we talk about solutions, we instantly think of liquids. **But is it necessary for all solutions to be liquids?**



No. A solution is simply a homogeneous mixture of two or more substances. Solutions can be solid, liquid and gaseous. **Alloy** is an example of a solid solution, while air is a gaseous solution.

A mixture is called solution when it has homogeneity at the particle level. A solution is formed when a **solute** is dissolved in a **solvent**.

Examples of solutions

Solutions	**Solvents**	**Solutes**
1. Saltwater	Water	Salt
2. Solution of iodine in alcohol	Alcohol	Iodine

3. Vinegar	Water	**Acetic acid**
4. Soda water	Water	Carbon dioxide
5. Air	Nitrogen (present in the largest amount)	Other gases (present in relatively smaller amounts)

Properties of Solutions

- They are homogeneous mixtures of solutes and solvents.
- The solute particles in a solution are extremely small in size. They are less than 1 nm (10^{-9} m) in diameter.
- The solute particles are not visible to the naked eye.
- As a result of the small size of the solute particles, a solution does not scatter a beam of light passing through it.
- Being small in size, the solute particles get dissolved in the solvent. Hence, the solute cannot be separated from the solvent by filtration.
- The solute particles do not settle down when left undisturbed.

Solved Examples

Hard

Example:

For each of the given substances, state whether or not it is a mixture. Also, if it is a mixture, then identify it as homogeneous or non-homogeneous.

Substance	Mixture	Homogeneous	Non-homogeneous
1.Solder			
2.Soap water			
3.Silver			
4.Green tea			

Solution:

Substance	Mixture	Homogeneous	Non-homogeneous
1.Solder	✓	✓	—
2.Soap water	✓	—	✓
3.Silver	✗	—	—
4.Green tea	✓	✓	—

Know More

Further Connect

Like a gas, a liquid exerts pressure of its own due to evaporation. This pressure is known as the vapour pressure of the liquid.

When a liquid solution is formed, it exerts its own vapour pressure. This results from the individual pressures of the solute and the solvent.

In 1882, a scientist named Francois-Marie Raoult established that for a solution containing volatile liquids, the partial vapour pressure of each component of the solution is proportional to its mole fraction present in solution.

$$p_1 \propto x_1$$
$$p_1 = p_1^o x_1$$

where p_1^o is the vapour pressure of pure component at the same temperature.

Did You Know?

1. The addition of solutes to solvents can lower the freezing point, elevate the boiling point and lower the vapour pressure of the solvents.

2. A solution is formed when two similar substances are mixed. For example, water and salt form a solution, but water and oil do not. This is because:

- Water and salt are polar substances (wherein the centres of positive charge and negative charge do not coincide), so they can mix with each other.
- Oil is a non-polar substance, so it does not mix with water to form a solution.

3. The solubility of a solute in a solvent is affected by temperature. Usually solubility increases with increase in temperature.

Concentration of a Solution

Concentration of a Solution

We come across many solutions in daily life, for example, saltwater, lime juice, squashes, coconut water, etc. How do we rate these solutions? We do so by classifying them as sweet, salty, etc., on the basis of our senses.

In chemistry, however, such classification of solutions is not very informative or beneficial. So, chemists usually use words such as 'saturated', 'unsaturated', 'supersaturated', 'dilute' and 'concentrated' to define the concentration of a given solution.

Uses of Concentration in Real Life

- Preparation of dyes
- Preparation of juices
- Labelling of alcoholic drinks
- Addition of antifreeze to vehicles
- Maintenance of the ionic balance in our body

Let us learn more about the concentration in this lesson.
A **solution** is a homogeneous mixture of two or more substances. The substance that is dissolved in a solution is called the **solute**, and the substance that dissolves the solute is called the **solvent**. The amount of solute in a solution may vary.

The amount of solute present in a given quantity of solution is called the concentration of that solution. The concentration of a solution helps us to determine the amount of solute present in the solution

$$\text{Concentration of solution} = \frac{\text{Amount of solute}}{\text{Amount of solution or solvent}}$$

Expression of the Concentration of a Solution

1. Concentration in terms of Mass by Mass Percentage

It is the amount of solute present in 100 gm of the solution.

$$\text{Mass percentage of solute} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

A 10% solution by weight indicates that 10 g of the solute is dissolved in 90 g of water so that the total mass of the solution is 100 g.
Mass by mass percentage can also be expressed in the following ways.

$$\text{Mass percentage of solute} = \frac{\text{Mass of solute}}{(\text{Mass of solvent} + \text{Mass of solute})} \times 100$$

or

$$\text{Mass percentage of solute} = \frac{\text{Mass of solute}}{(\text{Volume ofsolution} \times \text{density of solution})} \times 100$$

or

$$\text{Mass percentage of the solute} = \frac{\text{Mass of solute}}{(\text{Volume of solvent } \times \text{density of solvent}) + \text{Mass of solute}} \times 100$$

2. Concentration in terms of Mass by Volume Percentage

It is the mass of solute in gram present in 100 mL of the solution.

$$\text{Mass by volume percentage of solute} = \frac{\text{Mass of solute}}{\text{Volume of solution}} \times 100$$

A 10% (W/V) solution indicates that 100 mL of solution contains 10 g of solute.

3. Concentration in terms of Volume by Volume Percentage

This expression is used for the solutions in which both solvent and solute are liquids.
It is defined as the volume of solute dissolved in 100 mL of the solution to give 100 mL of a solution.

$$\text{Volume by Volume percentage of solute} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

or

$$\text{Volume by Volume percentage} = \frac{\text{Volume of Solute}}{(\text{Volume of solvent} + \text{Volume of solute})} \times 100$$

or

$$\text{Volume by Volume percentage} = \frac{\text{Mass of solute} \times \text{density of solute}}{\text{Mass of solution} \times \text{Density of solution}} \times 100$$

4. Concentration in terms of Parts per million(ppm)

This expression is used for very dilute solutions, i.e. when a very small amount of solute is present in the solution, then the concentration of the solute is expressed in terms of ppm (parts per million). For example, atmospheric pollution is expressed in ppm.
It is defined as the amount of solute in gram present in one million (10^6) grams of the solution.

$$\text{Parts per million}\left(\text{ppm}\right) = \frac{\text{Number of parts of the component}}{\text{Total number of parts of all components of the solution}} \times 10^6$$

15 ppm means that 15 g of the solute is present in 10^6 g of solution.

5. Concentration in terms of Mole fraction

It is the ratio of the moles of one component of the solution to the total number of moles of all the components present in the solution. Mole fraction is indicated by χ.
For example, a solution is composed of three components say A, B, and C

Mole fraction of A:

$$\chi_A = \frac{\text{Number of moles of A}}{\text{Number of moles of A+ Number of moles of B+Number of moles of C}}$$

Mole fraction of B:

$$\chi_B = \frac{\text{Number of moles of B}}{\text{Number of moles of A+ Number of moles of B+Number of moles of C}}$$

Mole fraction of C:

$$\chi_C = \frac{\text{Number of moles of C}}{\text{Number of moles of A+ Number of moles of B+Number of moles of C}}$$

The sum of mole fractions of all the components of a solution is equal to 1.
$\chi A + \chi B + \chi C = 1$

For binary solutions of components A and B

$$\chi_A = \frac{n_A}{n_A + n_B} = \frac{\frac{w_A}{M_A}}{\frac{w_A}{M_A} + \frac{w_B}{M_B}}$$

and

$$\chi_B = \frac{n_B}{n_A + n_b} = \frac{\frac{w_B}{M_B}}{\frac{w_A}{M_A} + \frac{w_B}{M_B}}$$

also,

$$\chi_B = 1 - \chi_A$$

For solution containing *i* number of components,

$$x_i = \frac{n_i}{n_1 + n_2 + n_3 + \ldots.. + n_i} = \frac{n_i}{\sum n_i}$$

And, $x_1 + x_2 + x_3 + \ldots. + x_i = \sum xi=1$

6. Concentration in terms of Molarity

It is defined as the number of moles of solute present in 1000 mL of the solution. Molarity is represented by M.

$$\text{Molarity } \left(M\right) = \frac{\text{Number of moles of solute}}{\text{Volume of solution in mL}} \times 1000$$

or

$$\text{M} = \frac{\text{Weight of solute}}{\text{Molar mass of solute} \times \text{Volume of solution in mL}} \times 1000$$

A 10 M solution indicates that 10 moles of solute are present in 1 L or 1000 mL of the solution.

7. Concentration in terms of Molality

It is defined as the number of moles of solute dissolved in 1000 g of the solvent. Molality is represented by m.

$$\text{Molality}\ \left(m\right)\ =\frac{\text{Number of moles of solute}}{\text{Weight of solvent in gram}}\times 1000$$

or

$$\text{m}\ =\frac{\text{Weight of solute}}{\text{Molar mass of solute}\times\text{Weight of solvent in gram}}\times 1000$$

8. Concentration in terms of Normality

Normality of a solution is defined as the number of gram equivalents of solute present per litre of the solution.

$$\text{Normality}\ =\frac{\text{Number of gram equivalents of solute}}{\text{Volumer of Solution in Liters}}$$

$$\text{Number of gram equivalents of solute}\ =\frac{\text{Mass of Solute}}{\text{Equivalent mass}}$$

$$\text{Equivalent weight}\ =\frac{\text{Molar mass}}{\text{Valency}}$$

Relation between Molarity and Mole fraction

Let us take

M = Molarity of solution

V = Volume of solution in L

ρ = density of solution

n_a and n_b be the number of moles of solvent and solute, respectively.

M_a and M_b be the molar masses of solvent and solute, respectively.

W_a = Weight of solvent, W_b = Weight of solute

$$\text{M}=\frac{n_b}{\text{V}}$$

$$n_b=\text{MV} \qquad \ldots\ldots(1)$$

$$x_b=\frac{n_b}{n_a+n_b} \qquad \ldots\ldots.(2)$$

On substituting equation (1) in equation (2), we get

$$x_b = \frac{\mathrm{MV}}{n_a + \mathrm{MV}} \qquad \ldots\ldots(\mathrm{I})$$

Now , we know $\rho = \frac{\mathrm{W}_{\mathrm{solution}}}{\mathrm{V}}$

$$\rho = \frac{\mathrm{W}_a + \mathrm{W}_b}{\mathrm{V}} \qquad \ldots..(3)$$

Also, $n_a = \frac{\mathrm{W}_a}{\mathrm{M}_a} \Rightarrow \mathrm{W}_a = n_a\mathrm{M}_a \qquad \ldots..(4)$

$$n_b = \frac{\mathrm{W}_b}{\mathrm{M}_b} \Rightarrow \mathrm{W}_b = n_b\mathrm{M}_b \qquad \ldots\ldots(5)$$

On Substituting equation (4) and equation (5) in equation (3), we get

$$\rho = \frac{n_a\mathrm{M}_a + n_b\mathrm{M}_b}{\mathrm{V}}$$

$$\rho = \frac{n_a\mathrm{M}_a}{\mathrm{V}} + \frac{n_b\mathrm{M}_b}{\mathrm{V}}$$

$$\rho = \frac{n_a\mathrm{M}_a}{\mathrm{V}} + \mathrm{M.M}_b \qquad \ldots\ldots(6) \qquad \left[\because \mathrm{M} = \frac{n_b}{\mathrm{V}}\right]$$

On rearranging equation (6), we get

$$n_a = \mathrm{V}\left(\frac{\rho - \mathrm{M.M}_b}{\mathrm{M}_a}\right) \qquad \ldots..(7)$$

Substitute equation (1) and (7) in equation (2)

$$x_b = \frac{\mathrm{MV}}{\mathrm{MV} + \mathrm{V}\left(\frac{\rho - \mathrm{M.M}_b}{\mathrm{M}_a}\right)}$$

Taking V in common, we get,

$$x_b = \frac{\mathrm{M}}{\mathrm{M} + \left(\frac{\rho - \mathrm{M.M}_b}{\mathrm{M}_a}\right)}$$

$$x_b = \frac{\mathrm{M/M}_a}{\mathrm{M.M}_a + (\rho - \mathrm{M.M}_b)}$$

$$x_b = \frac{\mathrm{M.M}_a}{\mathrm{M}(\mathrm{M}_a - \mathrm{M}_b) + \rho}$$

Relation between Molarity and Molality

Let

$\mathrm{W_B}$ = Mass of solute in gram
$\mathrm{W_A}$ = Mass of solvent in gram
V = Volume of the solution
$\mathrm{M_B}$ = Molar mass of solute
$\mathrm{M_A}$ = Molar mass of solvent

n_A = Number of moles of solute
n_B = Number of moles of solvent
ρ = Density of solution
M = Molarity
and
m = Molality

Now,

$$\frac{\text{Molarity}}{\text{Molality}} = \frac{\mathrm{M}}{\mathrm{m}} = \frac{\mathrm{n_b}}{\mathrm{V}} \times \frac{\mathrm{W_A}}{\mathrm{n_b}} = \frac{\mathrm{W_A}}{\mathrm{V}}$$

Mass of the solution = (Volume of solution)(Density of solution) = Vρ

Also

$W_B = n_B M_B$

and

$W_A = V\rho - n_B M_B$

Or

$$\frac{M}{m} = \frac{V\rho - n_B M_B}{V}$$

$$\Rightarrow \; m = \frac{MV}{V\rho - n_B M_B}$$

$$\Rightarrow \; \mathrm{m} = \frac{M}{\rho - M\mathrm{M_B}}$$

Out of molality and molarity, molality is considered to be a better way to express the concentration and thus it is prefered over molality in quantitative analysis because molality is independent of temperature. It involves mass of the solute and solvent, and we know that mass of a substance which does not change with temperature while the volume of solution may change with change in temperature.

Solubility

- Solubility of a substance is the maximum amount of the substance that can be dissolved in a specified amount of a solvent at a given temperature.

- Factors on which solubility depends:

- Nature of solute and solvent
- Temperature
- Pressure

Solubility of a Solid in a Liquid

When a solute is dissolved in a solvent, the following dynamic equilibrium is established.

Solute + Solvent $\longleftrightarrow$ Solution

This equilibrium follows Le Chatelier's Principle.

- Effect of temperature:

  According to Le Chatelier's principle, in a nearly saturated solution, if the dissolution process is −

- Endothermic ($\Delta_{sol}H > 0$), then the solubility will increase with the increase in temperature
- Exothermic ($\Delta_{sol}H < 0$), then the solubility will decrease with the increase in temperature
- Effect of pressure:
- Pressure has no significant effect on the solubility of solids in liquids.
- Reason − Solids and liquids are negligibly affected by pressure as they are highly incompressible.

Solubility of a Gas in a Liquid

- The solubility of gases in liquids increases with the increase in pressure.
- Henry's law − The solubility of a gas in a liquid is directly proportional to the pressure of the gas.

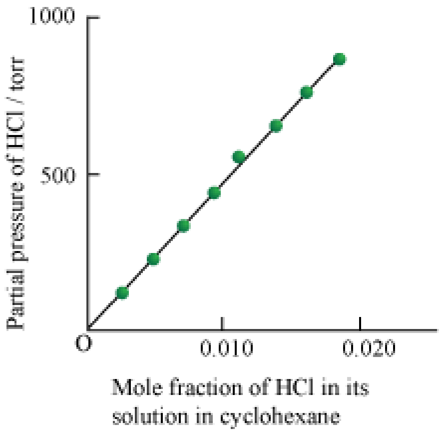
  Or

  The partial pressure of a gas in vapour phase (p) is directly proportional to the mole fraction of the gas (x) in the solution, i.e.,

  $p = K_H x$

  Where, K_H = Henry's law constant

  The plot of the partial pressure of HCl vs. its mole fraction in a solution of cyclohexane is given below.



- Some important applications of Henry's law

  (i) Bottles are sealed under high pressure to increase the solubility of CO_2 in soft drinks and soda water.

  (ii) Henry's law states that the solubility of gas increases with an increase in pressure. Therefore, when a scuba diver dives into the sea, the increased sea pressure causes the nitrogen present in air to dissolve in his blood in great amounts. As a result, when he comes back to the surface, the solubility of nitrogen again decreases and the dissolved gas is released, leading to the formation of nitrogen bubbles in the blood. This results in the blockage of capillaries and leads to a medical condition known as 'bends', which are painful and dangerous to life. Hence, the oxygen tanks used by scuba divers are filled with air and diluted with helium to avoid bends.

  (iii) The concentration of oxygen is low in the blood and tissues of climbers or people living at high altitudes. This is because at high altitudes, the partial pressure of oxygen is less than that at ground level. Low blood oxygen causes climbers to become weak and prevents them from thinking clearly. These are symptoms of 'anoxia'.

- Effect of temperature

  Solubility of gases in liquids decreases with the increase in temperature.

Vapour Pressure of Liquid Solutions

- Let p_1, p_2 = Partial vapour pressure of two volatile components 1 and 2 of a mixture

  p_1^0, p_2^0 = Vapour pressure of pure components 1 and 2

  x_1, x_2 = Mole fractions of the components 1 and 2

  p_{total} = Total vapour pressure of the mixture

Raoult's law

For a solution of volatile liquids, the partial vapour pressure of each component in the solution is directly proportional to its mole fraction.

That is, for component 1,

$p_1 \propto x_1$

And, $p_1 = p_1^0 x_1$

For component 2,

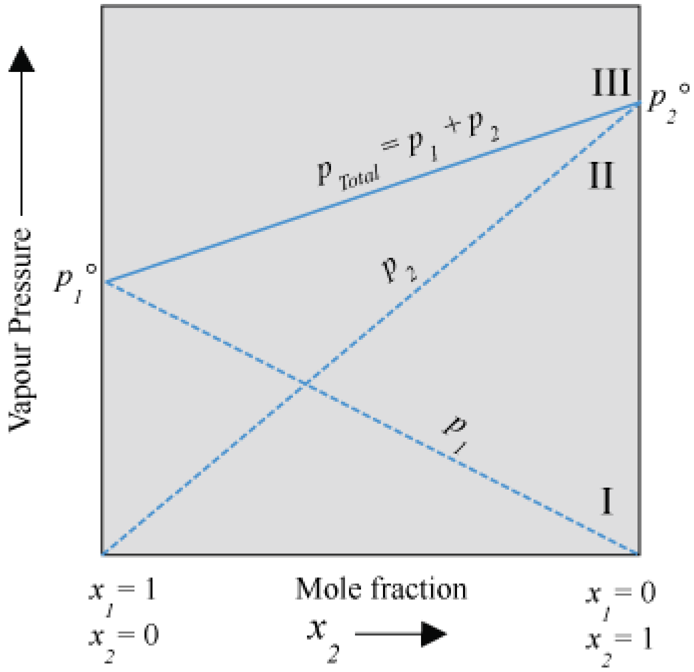
$$p_2 = p_2^0 x_2$$

According to Dalton's law of partial pressures,

$$\begin{aligned} p_{\text{total}} &= p_1 + p_2 \\ &= p_1^0 x_1 + p_2^0 x_2 \\ &= (1 - x_2) p_1^0 + x_2 p_2^0 \\ p_{\text{total}} &= p_1^0 + (p_2^0 - p_1^0) x_2 \end{aligned}$$

- From the above equation, it can be concluded that:

- p_{total} can be related to mole fraction of any one component.
- p_{total} varies linearly with x_2.
- Depending upon p_1^0 and p_2^0, p_{total} increases or decreases with the increase of x_1.
- The plot of vapour pressure and mole fraction of an ideal solution at constant temperature is shown below.



Here, maximum value of $p_{\text{total}} = p_2^0$

Minimum value of $p_{\text{total}} = p_1^0$

Here, $p_1^0 < p_2^0$ because it is assumed that component 1 is less volatile than component 2.

- The component of vapour phase in equilibrium with the solution is determined by the partial pressures of the components.

  Let y_1, y_2 = Mole fractions of the components 1 and 2 respectively in the vapour phase

  According to Dalton's law of partial pressures,

  $$p_1 = y_1 \, p_{\text{total}}$$
  $$p_2 = y_2 \, p_{\text{total}}$$

  In general, we can write

  $$p_i = y_i \, p_{\text{total}}$$

Raoult's Law as a Special Case of Henry's Law

- According to Raoult's law, the vapour pressure of a volatile component in a given solution is $p_i = x_i \, p_i^0$

- According to Henry's law, the partial vapour pressure of a gas (the component is so volatile that it exists as gas) in a liquid is

  $p = K_H x$

- It can be observed that in both the equations, the partial vapour pressure of the volatile component varies directly with its mole fraction. Only the proportionality constants K_H and p_i^0 are different. Thus, Raoult's law becomes a special case of Henry's law in which K_H is equal to p_i^0.

Vapour Pressure of Solutions of Solids in Liquids

- When a non-volatile solute is added to a solvent, the vapour pressure of the liquid decreases.

- Reason: The number of solvent molecules on the surface decreases and as a result, number of solvent molecules escaping from the surface decreases.

- Raoult's law in general form: For any solution, the partial vapour pressure of each volatile component in the solution is directly proportional to its mole fraction.
- Let us take a binary solution made by dissolving a non-volatile solute in a solvent. Since the solute is non-volatile, only the solvent molecules contribute to vapour pressure.

Let p_1' = Vapour pressure of the solvent

p_1^0 = Vapour pressure of the solvent in pure state

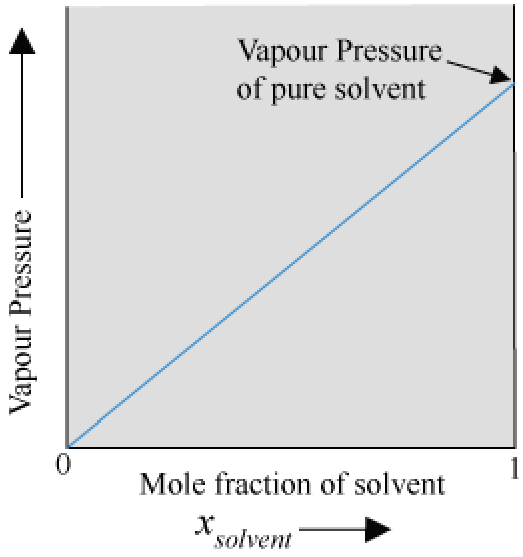
x_1 = Mole fraction of the solvent

Then, according to Raoult's law,

$$p_1 = x_1 \cdot p_1^0$$

Here, the proportionality constant is equal to the vapour pressure of the solvent in pure state.

The plot of vapour pressure vs. mole fraction of the solvent, which is linear, is shown below.



Ideal and Non-Ideal Solutions

Ideal Solution

- Solutions which obey Raoult's law over the entire range of concentrations
- For ideal solution:
- Enthalpy of mixing of the pure components to form the solution,

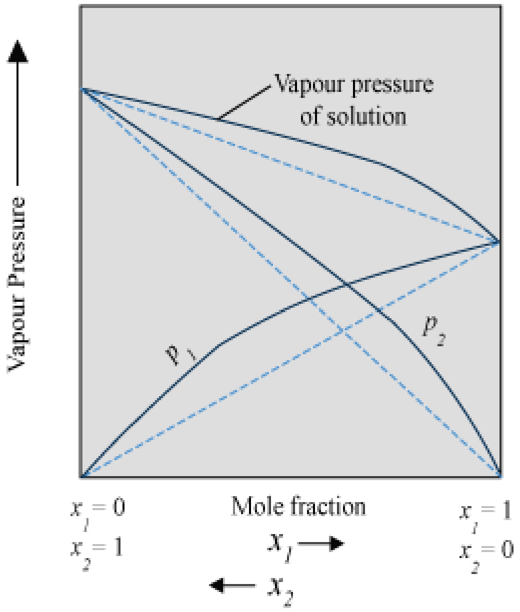
  $\Delta_{mix} H = 0$

- Volume of mixing, $\Delta_{mix} V = 0$
- An ideal solution will be formed when intermolecular forces of attraction between the molecules of solute (A – A) and those between the molecules of solvent (B – B) are nearly equal to those between solute and solvent molecules (A – B).
- Examples: *n*-Hexane and *n*-heptane, bromoethane and chloroethane, benzene and toluene

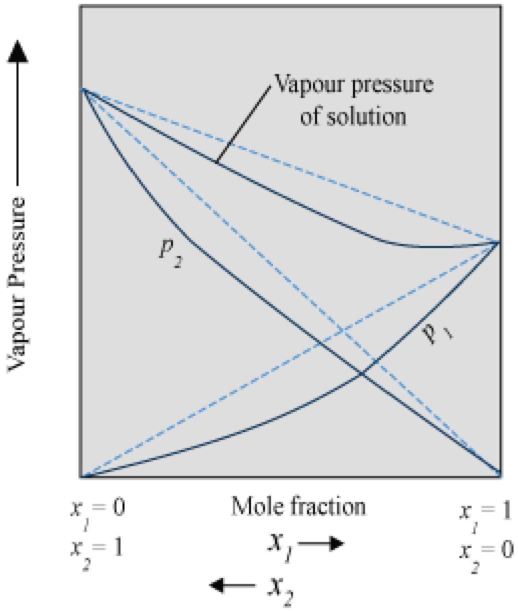
Non-Ideal Solutions

- Solutions which do not obey Raoult's law over the entire range of concentration
- The vapour pressure of a non-ideal solution is either higher or lower than that predicted by Raoult's law.
- Positive deviation from Raoult's law – When vapour pressure of solution is higher

- Negative deviation from Raoult's law – When vapour pressure of solution is lower
- Solution showing positive deviation from Raoult's law
- The plot of vapour pressure of two component solutions as a function of mole fraction is shown below.



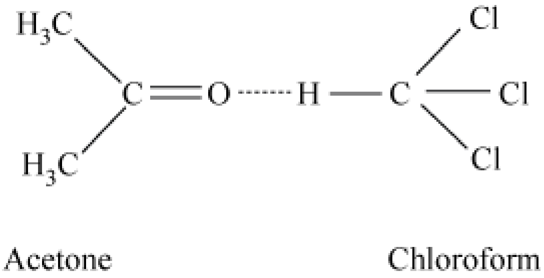
- The intermolecular forces of attraction between solute-solvent molecules are weaker than those between solute-solute molecules and solvent-solvent molecules. Therefore, solvent molecules can easily escape, resulting in increase in vapour pressure.
- Example: Ethanol and acetone mixture.
- Solution showing negative deviation from Raoult's law
- The plot of vapour pressure of two component solutions as a function of mole fraction is shown below.



- The intermolecular forces of attraction between solute-solute molecules and solvent-solvent molecules are weaker than those between solute-solvent molecules. This results in the decreasing of vapour pressure.

- Example − Chloroform and acetone mixture

  The intermolecular attractive forces between solute-solvent molecules increase due to the formation of H-bond.



Azeotropes

- Binary mixtures which have the same composition in liquid and vapour phase, and have constant boiling points

- Not possible to separate the components by fractional distillation

- Two types − Minimum boiling azeotrope

  Maximum boiling azeotrope

- Minimum boiling azeotrope − Solution showing a large positive deviation from Raoult’s law of specific composition

  Example: Ethanol-water mixture containing ethanol approximately 95% by volume

- Maximum boiling azeotrope − Solution showing a large negative deviation from Raoult’s law at specific composition

  Example: Nitric acid-water mixture containing 68% nitric acid and 32% water by mass

  The boiling point of this azeotrope is 393.5 K.

Colligative Properties and Determination of Molar mass

- Properties which depend upon the number of solute particles present in the solution irrespective of their nature, relative to the total number of particles present in the solution
- Examples: Relative lowering of vapour pressure of the solvent, depression of freezing point of the solvent, elevation of boiling point of the solvent, osmotic pressure of the solution

Relative Lowering of Vapour Pressure

- Relative lowering of vapour pressure is equal to the mole fraction of the solute.

i.e., $\frac{p_1^0 - p_1}{p_1^0} = x_2$

Where,

p_1^0 = Vapour pressure of the solvent

p_1 = Vapour pressure of the solution

x_2 = Mole fraction of the solute

However, $x_2 = \frac{n_2}{n_1 + n_2}$

Where, n_2 = Number of moles of solute

n_1 = Number of moles of solvent

Now, $\frac{p_1^0 - p_1}{p_1^0} = \frac{n_2}{n_1 + n_2}$

For dilute solution, $n_1 >> n_2$

Therefore, $\frac{p_1^0 - p_1}{p_1^0} = \frac{n_2}{n_1}$

Or, $\frac{p_1^0 - p_1}{p_1^0} = \frac{\frac{w_2}{M_2}}{\frac{w_1}{M_1}}$

Or, $\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$

Where,

w_2 = Mass of solute

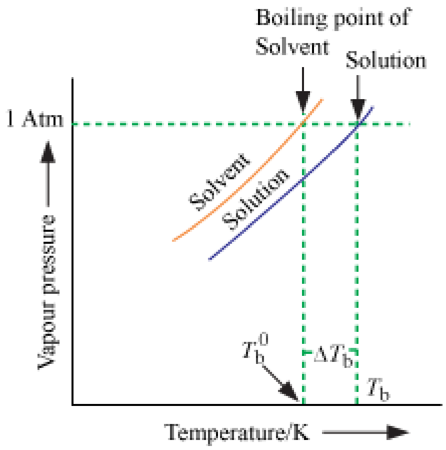
M_2 = Molar mass of solute

w_1 = Mass of solvent

M_1= Molar mass of solvent

Elevation of Boiling Point

- With addition of non-volatile solute, vapour pressure decreases and hence, boiling point increases.



- Here, $\Delta T_b = T_b - T_b^0$ = Elevation of boiling point

  Where,

  T_b = Boiling of solution

  T_b^0 = Boiling point of solvent

- $\Delta T_b \propto m$ (for dilute solution)

  $\Rightarrow \Delta T_b = K_b m$

  Where,

  m = Molality

  K_b = Proportionality constant known as Boiling Point Elevation Constant or Molal Elevation Constant or Ebullioscopic constant (Unit = K kg mol^{-1})

  However, $m = \dfrac{\dfrac{w_2}{M_2}}{\dfrac{w_1}{1000}} = \dfrac{1000 \times w_2}{M_2 \times w_1}$

  Where,

  w_2 = Mass of solute

  M_2 = Molar mass of solute

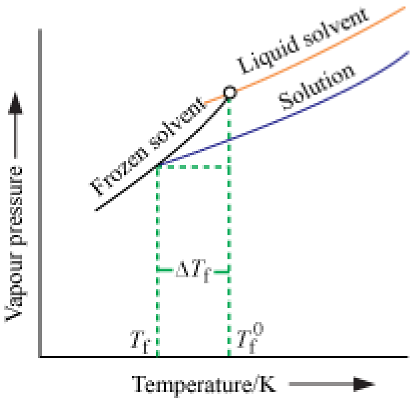
  w_1 = Mass of solvent

  Now, $\Delta T_b = \dfrac{K_b \times 1000 \times w_2}{M_2 \times w_1}$

  Or, $M_2 = \dfrac{1000 \times w_2 \times K_b}{\Delta T_b \times w_1}$

  Depression of Freezing point

- With addition of non-volatile solute, vapour pressure decreases, which leads to a decrease in freezing point.



- Here, $\Delta T_f = T_f^0 - T_f$ = Depression of freezing point

  Where,

  T_f^0 = Freezing point of solvent

  T_f= Freezing point of solution

- For dilute solution (ideal solution),
  $\Delta T_f \propto m$

  $\Rightarrow \Delta T_f = K_f m$

  Where,

  m = Molality

  K_f = Proportionality constant known as Freezing Point Constant or Molal Depression Constant or Cryoscopic constant (Unit = K kg mol^{-1})

  However, $$m = \frac{\frac{w_2}{M_2}}{\frac{w_1}{1000}} = \frac{1000 \times w_2}{M_2 \times w_1}$$

  Where,

  w_2 = Mass of solute

  M_2 = Molar mass of solute

  w_1 = Mass of solvent

Now, $\Delta T_f = \dfrac{K_f \times w_2 \times 1000}{M_2 \times w_1}$

Or, $M_2 = \dfrac{1000 \times w_2 \times K_f}{\Delta T_f \times w_1}$

- The value of K_b and K_f can be determined by the following relations:

$$K_b = \frac{\mathrm{R} \times M_1 \times T_b^2}{1000 \times \Delta_{\mathrm{vap}} H}$$

$$K_f = \frac{\mathrm{R} \times M_1 \times T_f^2}{1000 \times \Delta_{\mathrm{fus}} H}$$

Where,

M_1 = Molar mass of the solvent

R = Gas constant

$\Delta_{\mathrm{vap}}H$ = Enthalpy of vaporisation

$\Delta_{\mathrm{fus}}H$ = Enthalpy of fusion

Antifreeze

Antifreeze solutions are added to a liquid in a cooling system (such as the water in an automobile engine) to lower its freezing point and prevent ice build-up in the system at cooler temperatures.

Functions of Antifreeze

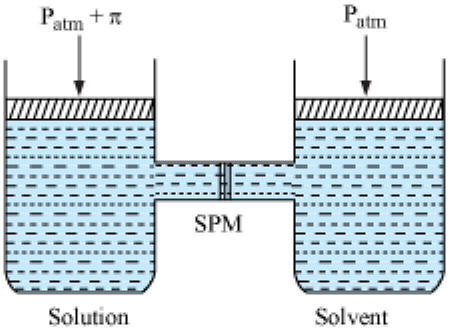
- It contains chemicals that prevent corrosion and scale formation in the engine and radiator of a vehicle.
- It provides protection against boiling during summers. At 1 atmosphere pressure, water boils at 100 °C; but a 50-50 blend of water−ethylene glycol boils at 106 °C.
- A mixture of antifreeze and distilled water (in the ratio, one part antifreeze to one part water) provides freeze protection down to −36.67 °C and boils over protection up to 129.4 °C.

Precautions

- Avoid using concentrated antifreeze in a cooling system. At least 40% of the mixture should be water.
- Do not increase the concentration of antifreeze above 60% as it damages the freezing and overheating protection of the engine.
- Do not add too much water to the cooling system as it lowers the concentration of the corrosion inhibitor and antifreeze. This results in decreased protection against corrosion and freezing.

Osmosis and Osmotic Pressure

- The process of flow of solvent molecules from pure solvent to solution or from solution of lower concentration to solution of higher concentration through a semi-permeable membrane is called osmosis.
- The pressure required to just stop the flow of solvent due to osmosis is called osmotic pressure (π) of the solution.
- The osmotic pressure π has to be applied to the solution to prevent osmosis.



- For dilute solution, osmotic pressure is directly proportional to the molarity C of the solution at a given temperature T.

That is, $\pi = C\mathrm{R}T$ (R = Gas constant)

$$\Rightarrow \pi = \left(\frac{n_2}{V}\right)\mathrm{R}T$$

Where,

n_2 = Number of moles of solute

V= Volume of the solution in litres

Again, $n_2 = \frac{w_2}{M_2}$

Where, w_2 = Mass of the solute

M_2 = Molar mass of the solute

Now, $\pi = \left(\frac{\frac{w_2}{M_2}}{V}\right) \mathrm{R}T$

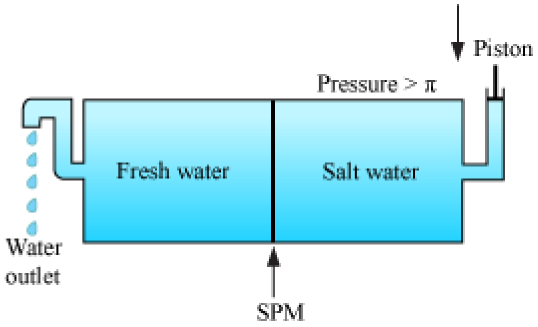
Or, $\pi V = \frac{w_2 \mathrm{R}T}{M_2}$

Or, $M_2 = \frac{w_2 \mathrm{R}T}{\pi V}$

- Now, osmosis does not occur between two isotonic solutions.
- Isotonic solutions − Solutions having same concentration of solute These solutions have same osmotic pressure.
- Hypotonic solution − Solution having a low concentration of solute relative to another.
- Hypertonic solution − Solution having a high concentration of solute relative to another.

Reverse Osmosis and Water purification

- If the pressure applied on the solution is greater than its osmotic pressure, then the direction of osmosis is reversed i.e., the solvent starts passing from solution to solvent. The phenomenon is called reverse osmosis.
- This phenomenon is used in purification (desalination) of sea water.



Abnormal Molar Masses

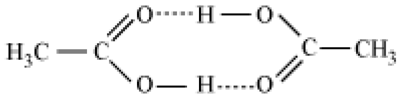
- Colligative property $\propto \dfrac{1}{\text{Molar mass of the solute}}$
- Due to association or dissociation of molecules, the molar mass of a substance calculated from its colligative property is either lower or higher than the expected or normal value. Such molar mass is called **abnormal molar mass**.

  Dissociation:

  $KCl \longleftrightarrow K^+ + Cl^-$

  Association:

  $2CH_3COOH \longleftrightarrow (CH_3COOH)_2$

- To account for the extent of dissociation or association, van't Hoff introduced a factor *i*, known as the van't Hoff factor.

  $$i = \frac{\text{Normal molar mass}}{\text{Abnormal molar mass}} = \frac{\text{Observed colligative property}}{\text{Calculated colligative property}} = \frac{\text{Total number of moles of particles after association/dissociation}}{\text{Number of moles of particles before association/dissociation}}$$

- Value of *i*:

For association, $i < 1$

For dissociation, $i > 1$

No association or dissociation, $i = 1$

- Modified equations for colligative properties after inclusion of van’t Hoff factor
- Relative lowering of vapour pressure of solvent,

$$\frac{p_1^0 - p_1}{p_1^0} = i\frac{n_2}{n_1}$$

- Elevation of boiling point, $\Delta T_b = iK_bm$
- Depression of freezing point, $\Delta T_f = iK_fm$
- Osmotic pressure of solution, $\pi = in_2RT/V$