SOLUTIONS OF NONELECTROLYTES

- **a system** is generally considered to be a bounded space or an exact quantity of a **material substance**
- Material substances can be mixed together to form a variety of **pharmaceutical mixtures (or dispersions):**
  a) true solutions
  b) colloidal dispersions
  c) coarse dispersions
- **A dispersion** consists of at least two phases with one or more dispersed (internal) **phases** contained in a single continuous (external) **phase.**
- phase is defined as a distinct homogeneous part of a system separated by definite boundaries from other parts of the system.

A. A **true solution** is defined as a mixture of two or more components that form a homogeneous molecular dispersion (one-phase system).

B. The **coarse dispersions** diameter of particles is greater than ~500 nm (0.5 µm).

    A. Two common pharmaceutical coarse dispersions are:

        A. emulsions (liquid–liquid dispersions)

        B. suspensions (solid–liquid dispersions)

C. A **colloidal dispersion** represents a system having a particle size intermediate between that of a true solution and a coarse dispersion, roughly 1 to 500 nm. It may be considered as a two-phase (heterogeneous) system under some circumstances and as a one-phase system (homogeneous) under other circumstances.

e,.g. blood, liposomes, and zinc oxide paste

- A solution composed of only two substances is known as a binary solution, and the components or constituents are referred to as the **solvent** (the greater amount) and the **solute** (the lesser amount).

- When a solid is dissolved in a liquid, however, the liquid is usually taken as the solvent and the solid as the solute, irrespective of the relative amounts of the constituents.

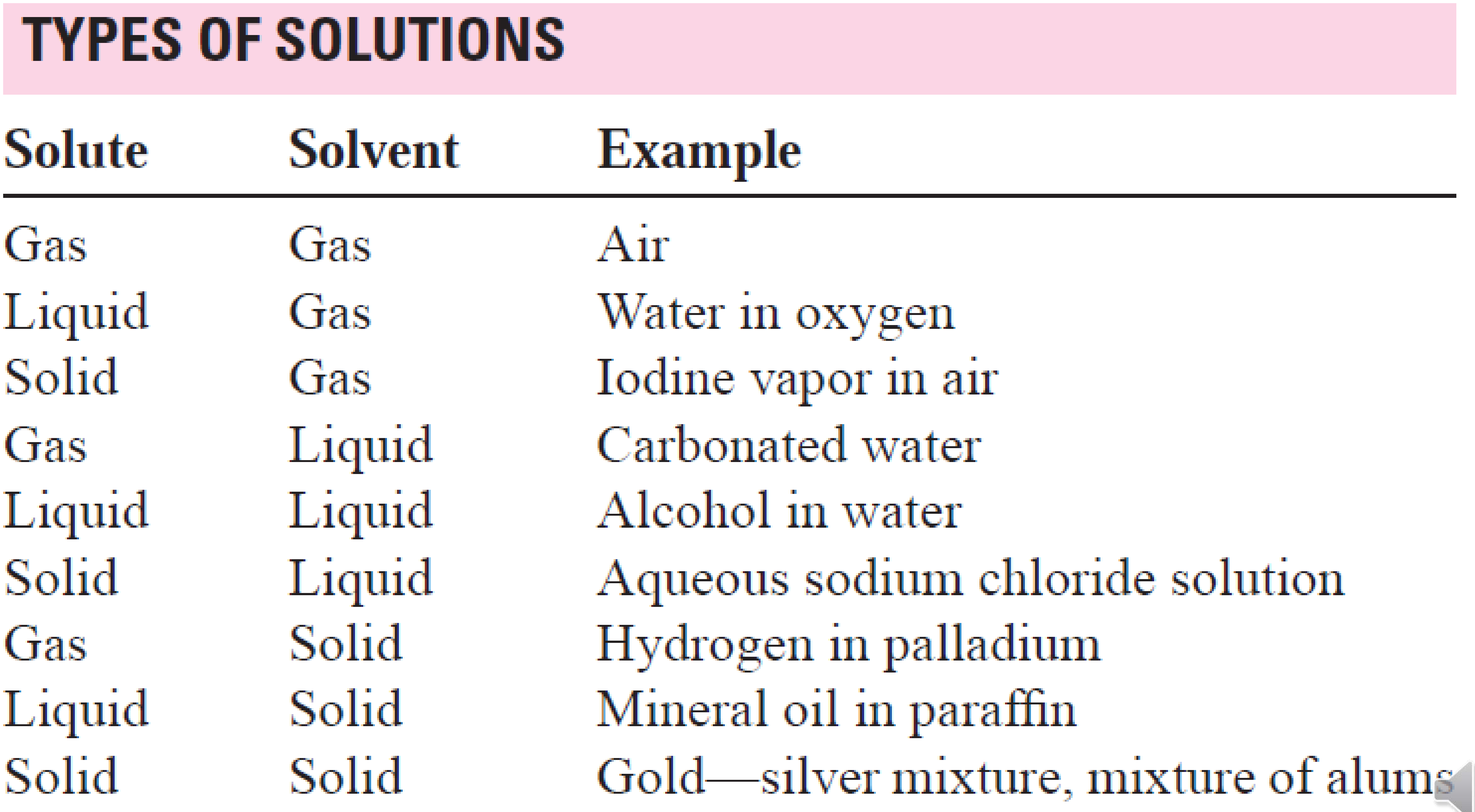
PHYSICAL PROPERTIES OF SUBSTANCES

- The physical properties of substances can be classified as colligative, additive, and constitutive.

- **Colligative properties** depend mainly on the **number of particles** in a solution.
- The colligative properties of solutions are
  a) osmotic pressure,
  b) vapor pressure lowering,
  c) freezing point depression
  d) boiling point elevation.

- The values of the colligative properties are approximately the same for equal concentrations of different nonelectrolytes in solution regardless of the species or chemical nature of the constituents

- *Types of Solutions*
- A solution can be classified according to the states in which the solute and solvent occur, and because three states of matter (gas, liquid, and crystalline solid) exist, nine types of homogeneous mixtures of solute and solvent are possible



TYPES OF SOLUTIONS

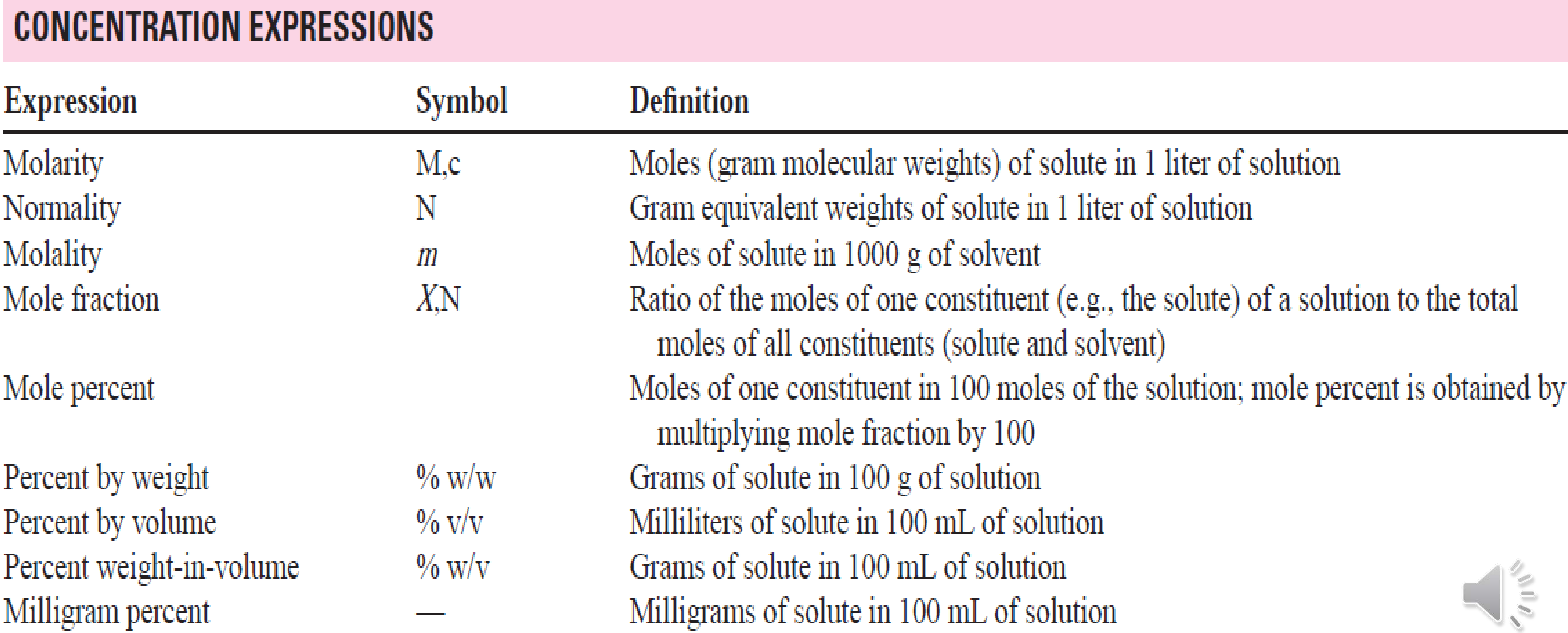
Solute	Solvent	Example
Gas	Gas	Air
Liquid	Gas	Water in oxygen
Solid	Gas	Iodine vapor in air
Gas	Liquid	Carbonated water
Liquid	Liquid	Alcohol in water
Solid	Liquid	Aqueous sodium chloride solution
Gas	Solid	Hydrogen in palladium
Liquid	Solid	Mineral oil in paraffin
Solid	Solid	Gold—silver mixture, mixture of alums

- The solutes (whether gases, liquids, or solids) are divided into two main classes:
  a) nonelectrolytes
  b) Electrolytes

- **Nonelectrolytes** are substances that do not ionize when dissolved in water and therefore do not conduct an electric current through the solution. E.g. sucrose, glycerin, naphthalene, and urea.
- The colligative properties of solutions of nonelectrolytes are fairly regular.

- **Electrolytes** are substances that form ions in solution, conduct electric current, and show apparent “anomalous” colligative properties; that is, they produce a considerably greater freezing point depression and boiling point elevation than do nonelectrolytes of the same concentration.
- Electrolytes may be subdivided further into strong electrolytes and weak electrolytes depending on whether the substance is completely or only partly ionized in water.
- Hydrochloric acid and sodium sulfate are strong electrolytes, whereas ephedrine and phenobarbital are weak electrolytes.

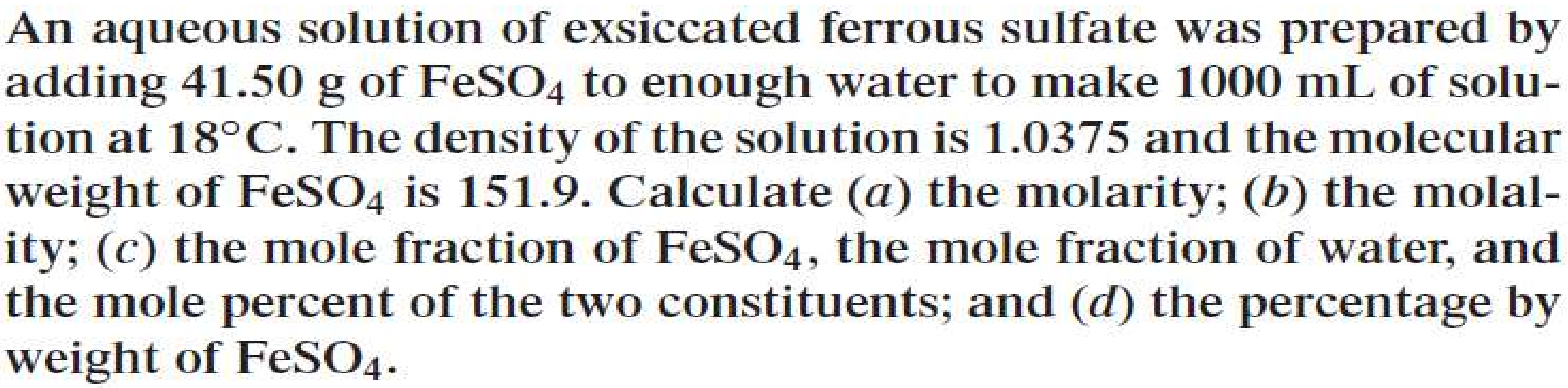
CONCENTRATION EXPRESSIONS



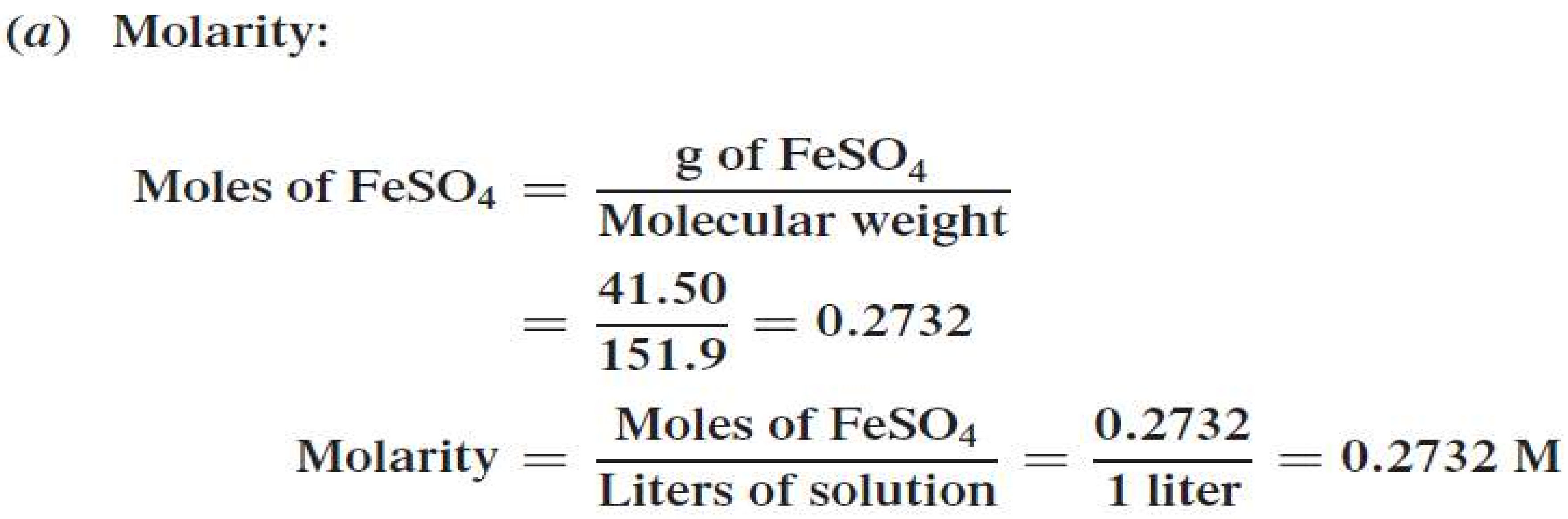
CONCENTRATION EXPRESSIONS

Expression	Symbol	Definition
Molarity	M,c	Moles (gram molecular weights) of solute in 1 liter of solution
Normality	N	Gram equivalent weights of solute in 1 liter of solution
Molality	*m*	Moles of solute in 1000 g of solvent
Mole fraction	*X*,N	Ratio of the moles of one constituent (e.g., the solute) of a solution to the total moles of all constituents (solute and solvent)
Mole percent		Moles of one constituent in 100 moles of the solution; mole percent is obtained by multiplying mole fraction by 100
Percent by weight	% w/w	Grams of solute in 100 g of solution
Percent by volume	% v/v	Milliliters of solute in 100 mL of solution
Percent weight-in-volume	% w/v	Grams of solute in 100 mL of solution
Milligram percent	—	Milligrams of solute in 100 mL of solution

Example



An aqueous solution of exsiccated ferrous sulfate was prepared by adding 41.50 g of $FeSO_4$ to enough water to make 1000 mL of solution at 18°C. The density of the solution is 1.0375 and the molecular weight of $FeSO_4$ is 151.9. Calculate (*a*) the molarity; (*b*) the molality; (*c*) the mole fraction of $FeSO_4$, the mole fraction of water, and the mole percent of the two constituents; and (*d*) the percentage by weight of $FeSO_4$.



(*a*) Molarity:

$$\text{Moles of FeSO}_4 = \frac{\text{g of FeSO}_4}{\text{Molecular weight}}$$

$$= \frac{41.50}{151.9} = 0.2732$$

$$\text{Molarity} = \frac{\text{Moles of FeSO}_4}{\text{Liters of solution}} = \frac{0.2732}{1\ \text{liter}} = 0.2732\ \text{M}$$

(*b*) Molality:

$$\text{Grams of solution} = \text{Volume} \times \text{Density}$$

$$1000 \times 1.0375 = 1037.5\ \text{g}$$

$$\text{Grams of solvent} = \text{Grams of solution} - \text{Grams of } FeSO_4 = 1037.5 - 41.5 = 996.0\ \text{g}$$

$$\text{Molality} = \frac{\text{Moles of } FeSO_4}{\text{kg of solvent}} = \frac{0.2732}{0.996} = 0.2743\, m$$

(*c*) Mole fraction and mole percent:

$$\text{Moles of water} = \frac{996}{18.02} = 55.27\ \text{moles}$$

Mole fraction of $FeSO_4$:

$$X_2 = \frac{\text{Moles of } FeSO_4}{\text{Moles of water} + \text{Moles of } FeSO_4}$$

$$= \frac{0.2732}{55.27 + 0.2732} = 0.0049$$

Mole fraction of water:

$$X_1 = \frac{55.27}{55.27 + 0.2732} = 0.9951$$

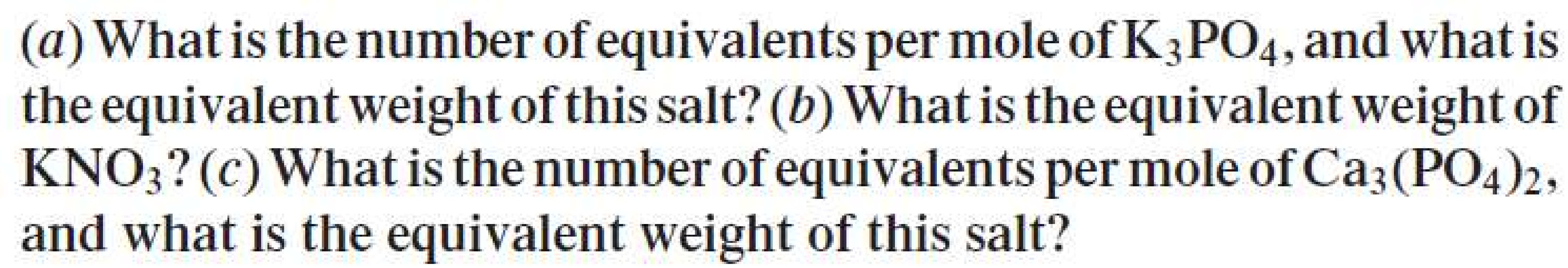
Notice that
$X_1 + X_2 = 0.9951 + 0.0049 = 1.0000$
Mole percent of $FeSO_4 = 0.0049 \times 100 = 0.49\%$
Mole percent of water $= 0.9951 \times 100 = 99.51\%$

(*d*) Percentage by weight of $FeSO_4$:

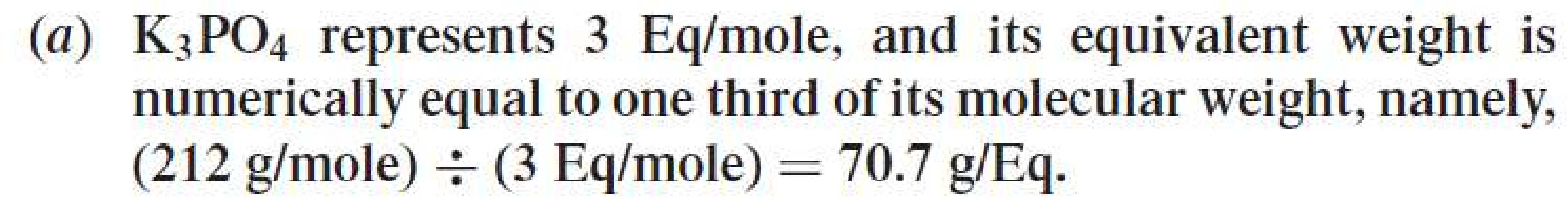
$$= \frac{\text{g of FeSO}_4}{\text{g of solution}} \times 100$$

$$= \frac{41.50}{1037.5} \times 100 = 4.00\%$$

EQUIVALENT WEIGHTS



(*a*) What is the number of equivalents per mole of K_3PO_4, and what is the equivalent weight of this salt? (*b*) What is the equivalent weight of KNO_3? (*c*) What is the number of equivalents per mole of $Ca_3(PO_4)_2$, and what is the equivalent weight of this salt?



(*a*) K_3PO_4 represents 3 Eq/mole, and its equivalent weight is numerically equal to one third of its molecular weight, namely, (212 g/mole) $\div$ (3 Eq/mole) $=$ 70.7 g/Eq.

(*b*) The equivalent weight of KNO_3 is also equal to its molecular weight, or 101 g/Eq.

(*c*) The number of equivalents per mole for $Ca_3(PO_4)_2$ is 6 (i.e., three calcium ions each with a valence of 2 or two phosphate ions each with a valence of 3). The equivalent weight of $Ca_3(PO_4)_2$ is therefore one sixth of its molecular weight, or $310/6 = 51.7$ g/Eq.

IDEAL AND REAL SOLUTIONS

Ideality

- ✓Ideality in a gas implies the complete absence of attractive forces
- ✓ideality in a solution means complete uniformity of attractive forces.
- Because a liquid is a highly condensed state, it cannot be expected to be devoid of attractive forces
- if, in a mixture of A and B molecules, the forces between A and A, B and B, and A and B are **all of the same order**, the solution is considered to be ideal according to the definition just given.

- an ideal solution as one in which there is no change in the properties of the components, other than dilution, when they are mixed to form the solution.

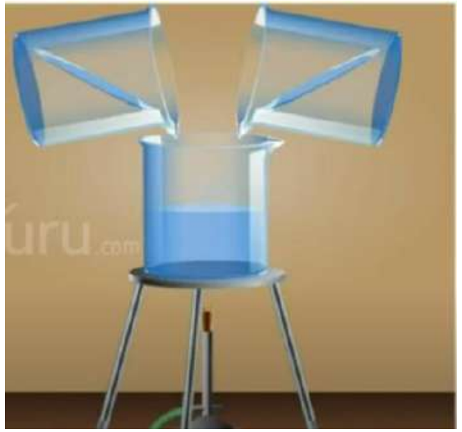
- Mixing substances with similar properties forms ideal solutions.

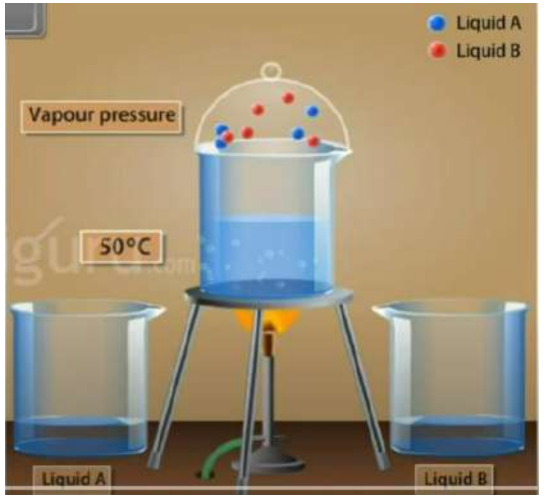
- For example, when 100 mL of methanol is mixed with 100 mL of ethanol, the final volume of the solution is 200 mL, and no heat is evolved or absorbed. The solution is nearly ideal.

- When 100 mL of sulfuric acid is combined with 100 mL of water, however, the volume of the solution is about 180mL at room temperature, and the mixing is attended by a considerable evolution of heat; the solution is said to be nonideal, or real.

Ideal Solutions and Raoult's Law

- in an ideal solution, the partial vapor pressure of each volatile constituent is equal to the vapor pressure of the pure constituent multiplied by its mole fraction in the solution.



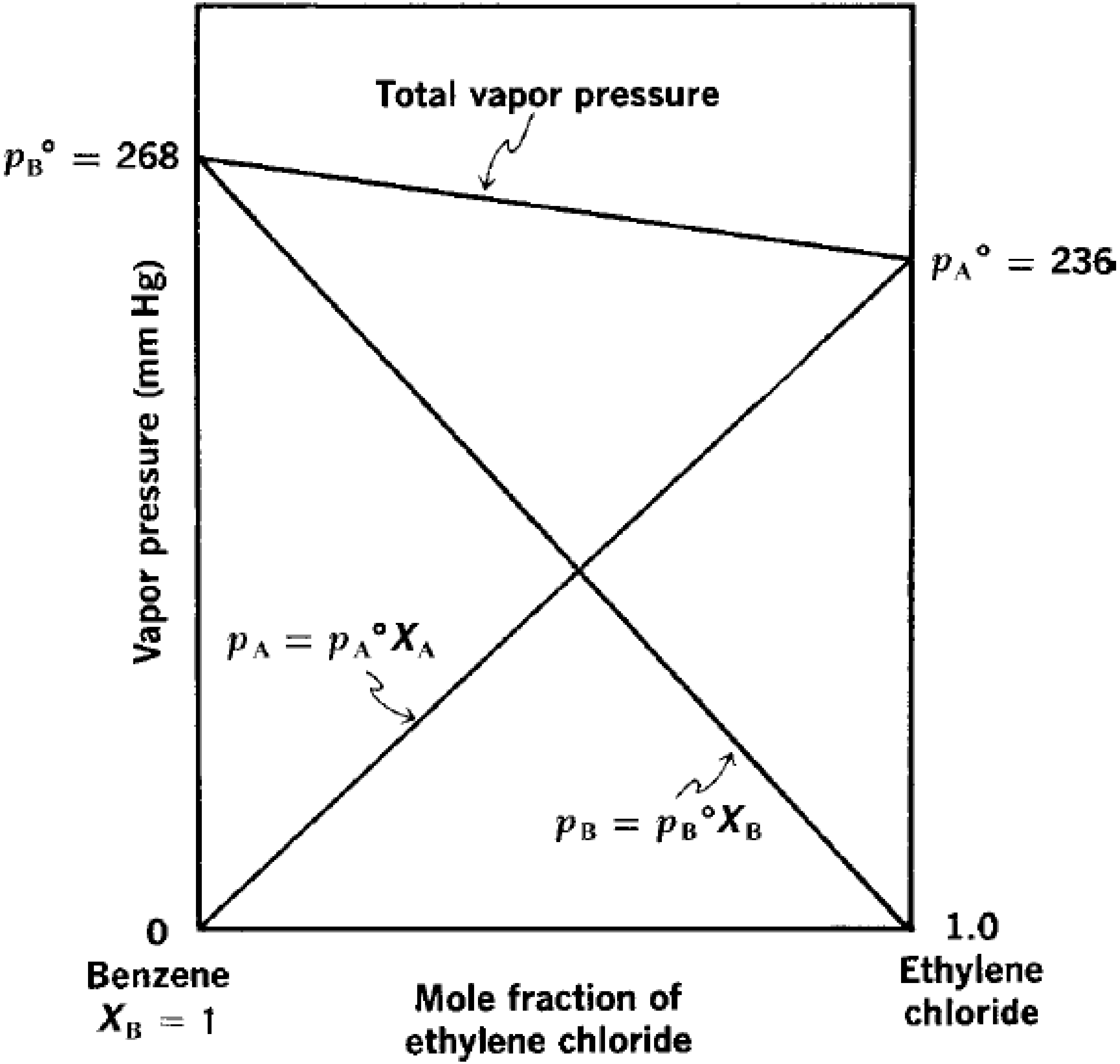


$$p_A = p_A{}^{\circ} X_A$$

$$p_B = p_B{}^{\circ} X_B$$

- where pA and pB are the partial vapor pressures of the constituents over the solution when the mole fraction concentrations are XA and XB, respectively.
- The vapor pressures of the pure components are pA◦ and pB◦, respectively.
- $X_A + X_B = 1$

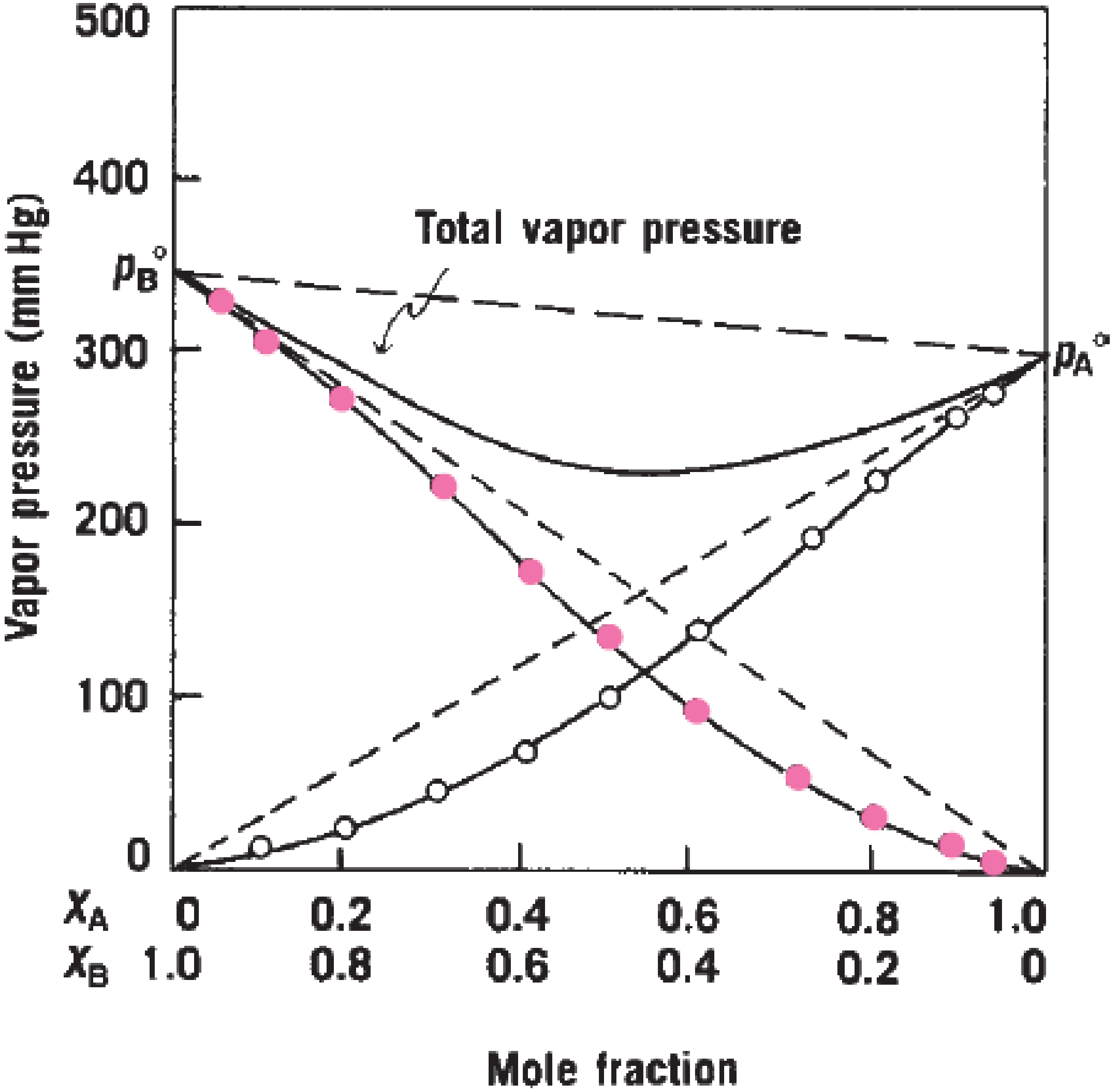
Vapor pressure–composition curve for an ideal binary system



Real Solutions

- Ideality in solutions presupposes complete uniformity of attractive forces.
- Many examples of solution pairs are known, however, in which the “cohesive” attraction of A for A exceeds the “adhesive” attraction existing between A and B.
- Similarly, the attractive forces between A and B may be greater than those between A and A or B and B. This may occur even though the liquids are miscible in all proportions
- Two types of deviation from Raoult’s law are recognized:
  - negative deviation
  - positive deviation

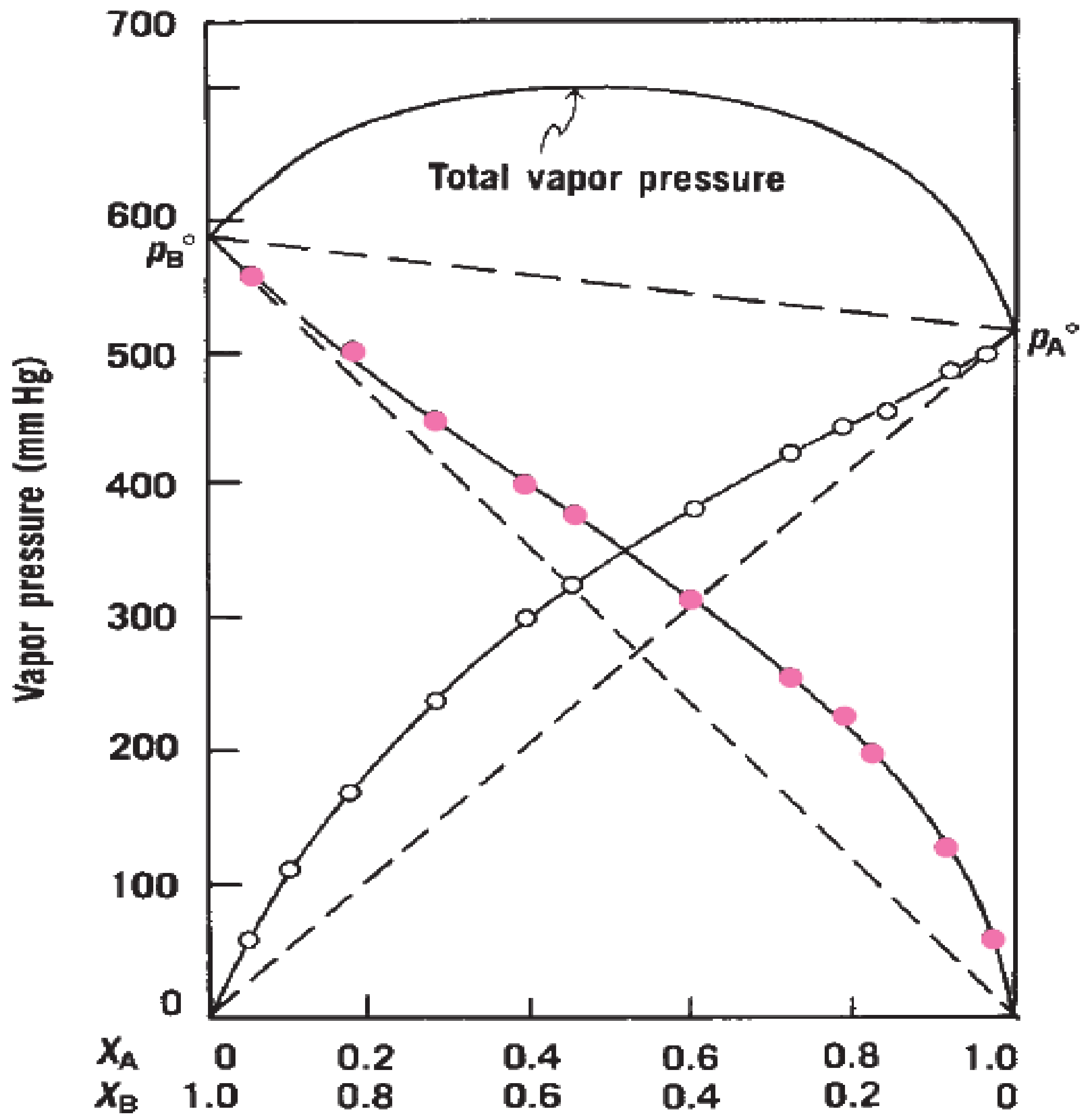
Vapor pressure of a system showing negative deviation from Raoult's law



Chloroform and acetone manifest such an attraction for one another through the formation of a hydrogen bond, thus further reducing the escaping tendency of each constituent.

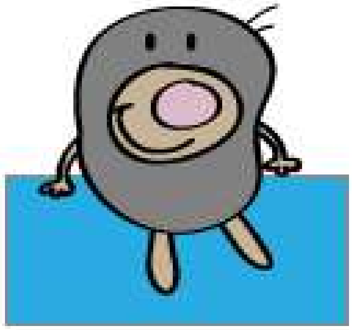
$$\mathrm{Cl_3C-H\cdots O=C(CH_3)_2}$$

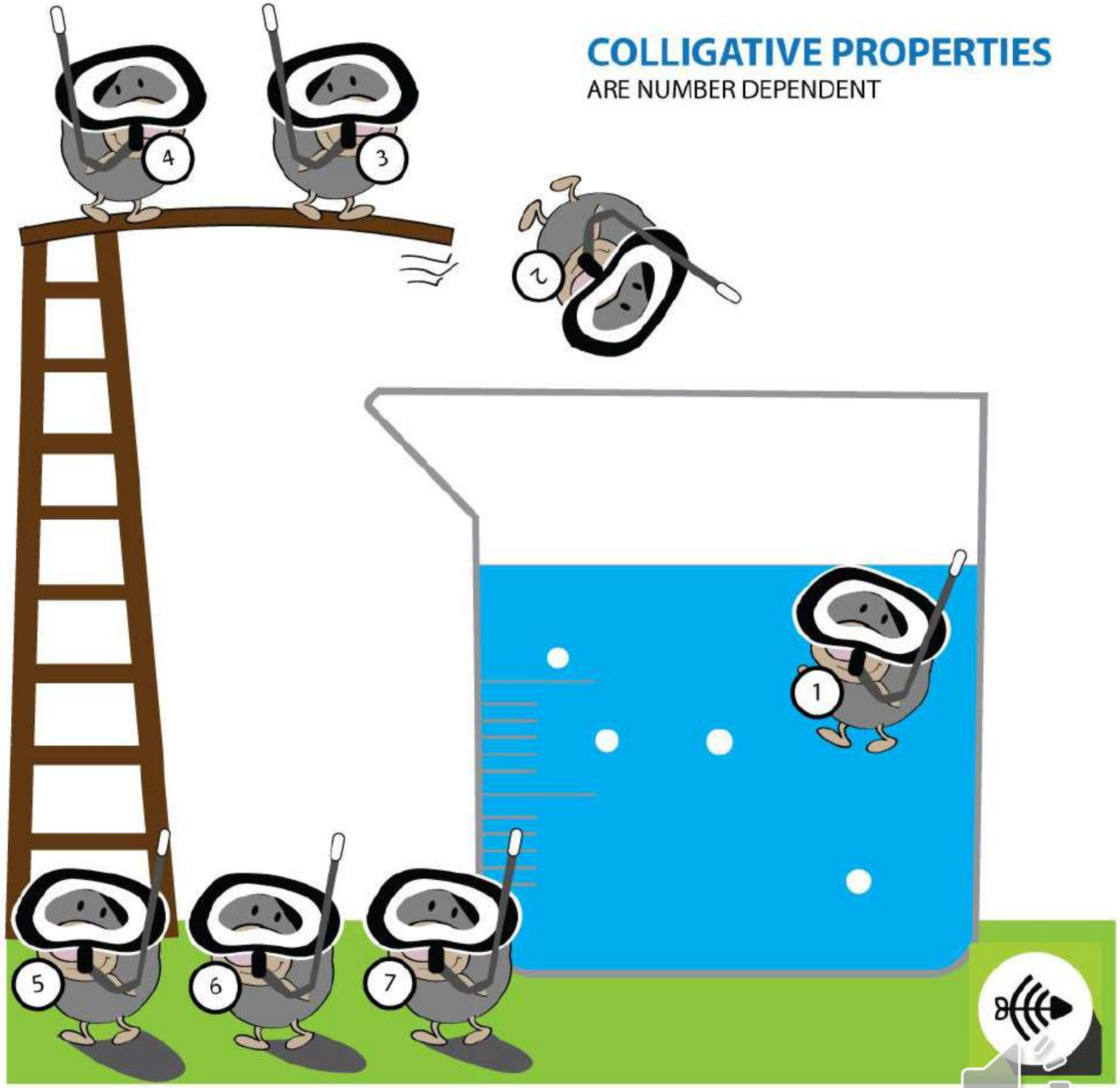
Vapor pressure of a system showing positive deviation from Raoult's law.



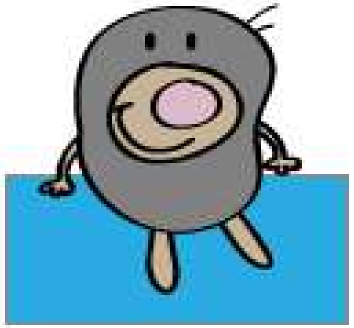
COLLIGATIVE PROPERTIES

- **Colligative properties** are those properties of solutions that depend on the ratio of the number of solute particles to the number of solvent molecules.





- Moles are not only cute but highly essential for colligative property calculations.

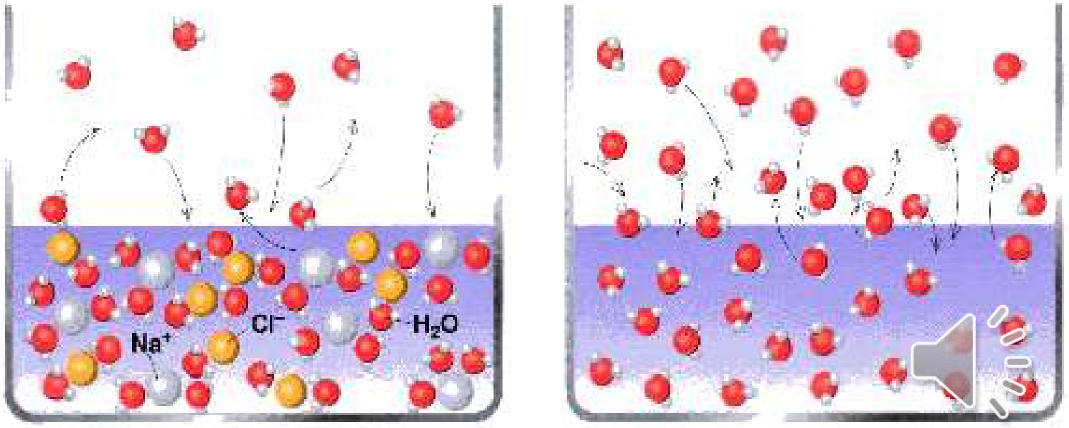


- That’s because the mole can determine the change in boiling point, freezing point or vapor pressure of the solution using what is called Molality.

- Molality (m): is measured by the number of moles of particles in a kg of solvent.

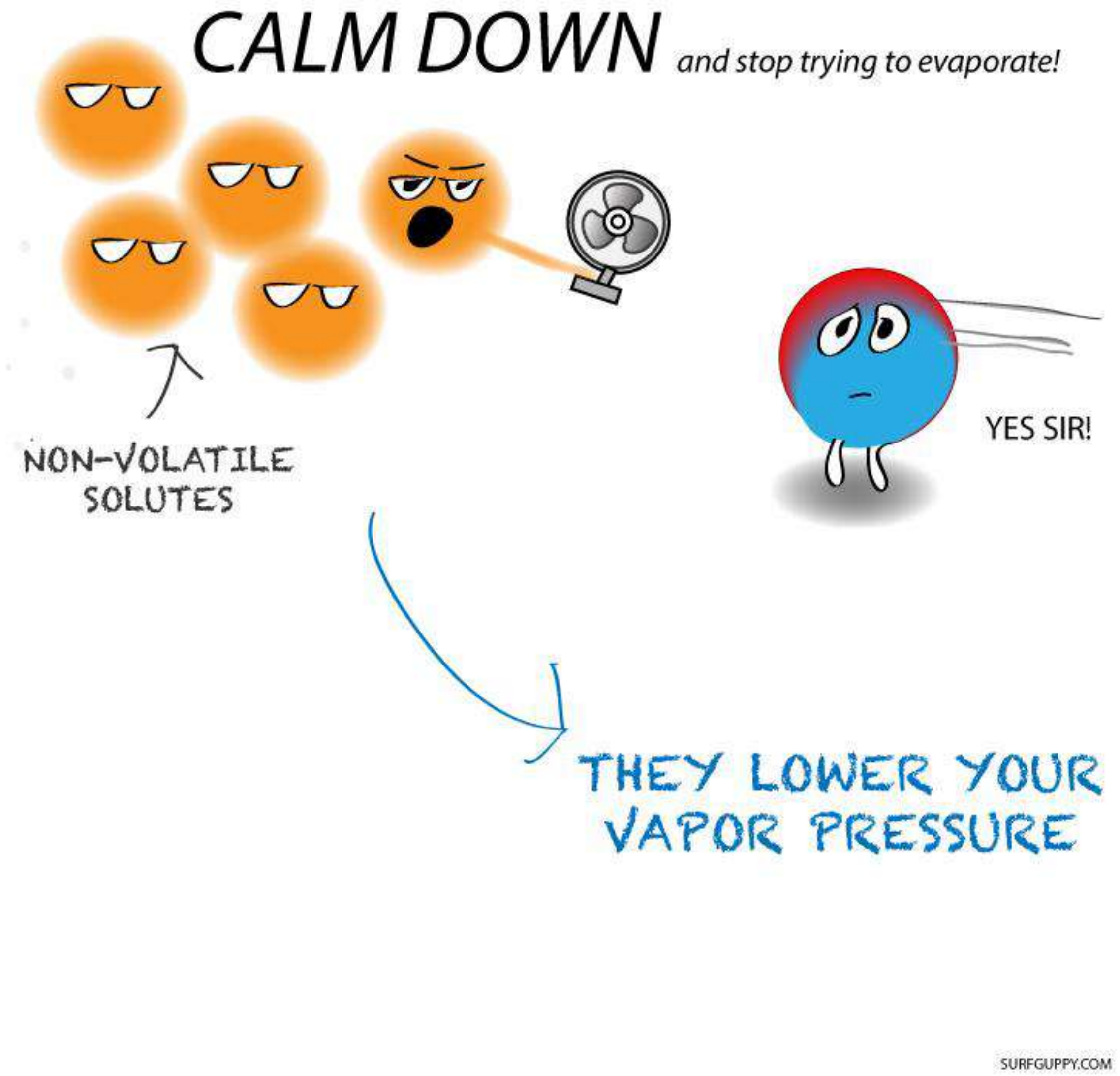
Raoult's Law

- the partial vapor pressure of each volatile constituent is equal to the vapor pressure of the pure constituent multiplied by its mole fraction in the solution
- When a nonvolatile solute is combined with a volatile solvent, the vapor above the solution is provided solely by the solvent.
- The solute reduces the escaping tendency of the solvent, and, on the basis of Raoult's law, the vapor pressure of a non-volatile solute is lowered proportional to the relative number.
- The Vapor Pressure of the solvent is proportional to its mole fraction



(a) Sea water (b) Pure water

Lowering of the Vapor Pressure



according to Raoult’s law, the vapor pressure, $p1$, of a solvent over a dilute solution is equal to the vapor pressure of the pure solvent, $p1\ \circ$, times the mole fraction of solvent in the solution, $X1$.

- Because the solute under discussion here is considered to be non-volatile, the vapor pressure of the solvent, $p1$, is identical to the total pressure of the solution, p.

- It is more convenient to express the vapor pressure of the solution in terms of the concentration of the solute rather than the mole fraction of the solvent, and this may be accomplished in the following way.

- The sum of the mole fractions of the constituents in a solution is unity:

$$X_1 + X_2 = 1$$

$$X_1 = 1 - X_2$$

- where *X*1 is the mole fraction of the solvent and *X*2 is the mole fraction of the solute.

$$p = p_1{}^{\circ}(1 - X_2)$$

$$p_1{}^{\circ} - p = p_1{}^{\circ} X_2$$

$$\frac{p_1{}^{\circ} - p}{p_1{}^{\circ}} = \frac{\Delta p}{p_1{}^{\circ}} = X_2 = \frac{n_2}{n_1 + n_2}$$

the lowering of the vapor pressure

the relative vapor pressure lowering

- The relative vapor pressure lowering depends only on the mole fraction of the solute, X2, that is, on the number of solute particles in a definite volume of solution.
- Therefore, the relative vapor pressure lowering is a colligative property.
- The freezing point, boiling point, and osmotic pressure of a solution also depend on the relative proportion of the molecules of the solute and the solvent.
- These are called colligative properties (Greek: “collected together”) because they depend chiefly on the number rather than on the nature of the constituents.

EXAMPLE

- Calculate the relative vapor pressure lowering at 20°C for a solution containing 171.2 g of sucrose (*w*2) in 1000 g (*w*1) of water. The molecular weight of sucrose (M2) is 342.3 g/mole and the molecular weight of water (M1) is 18.02 g/mole.

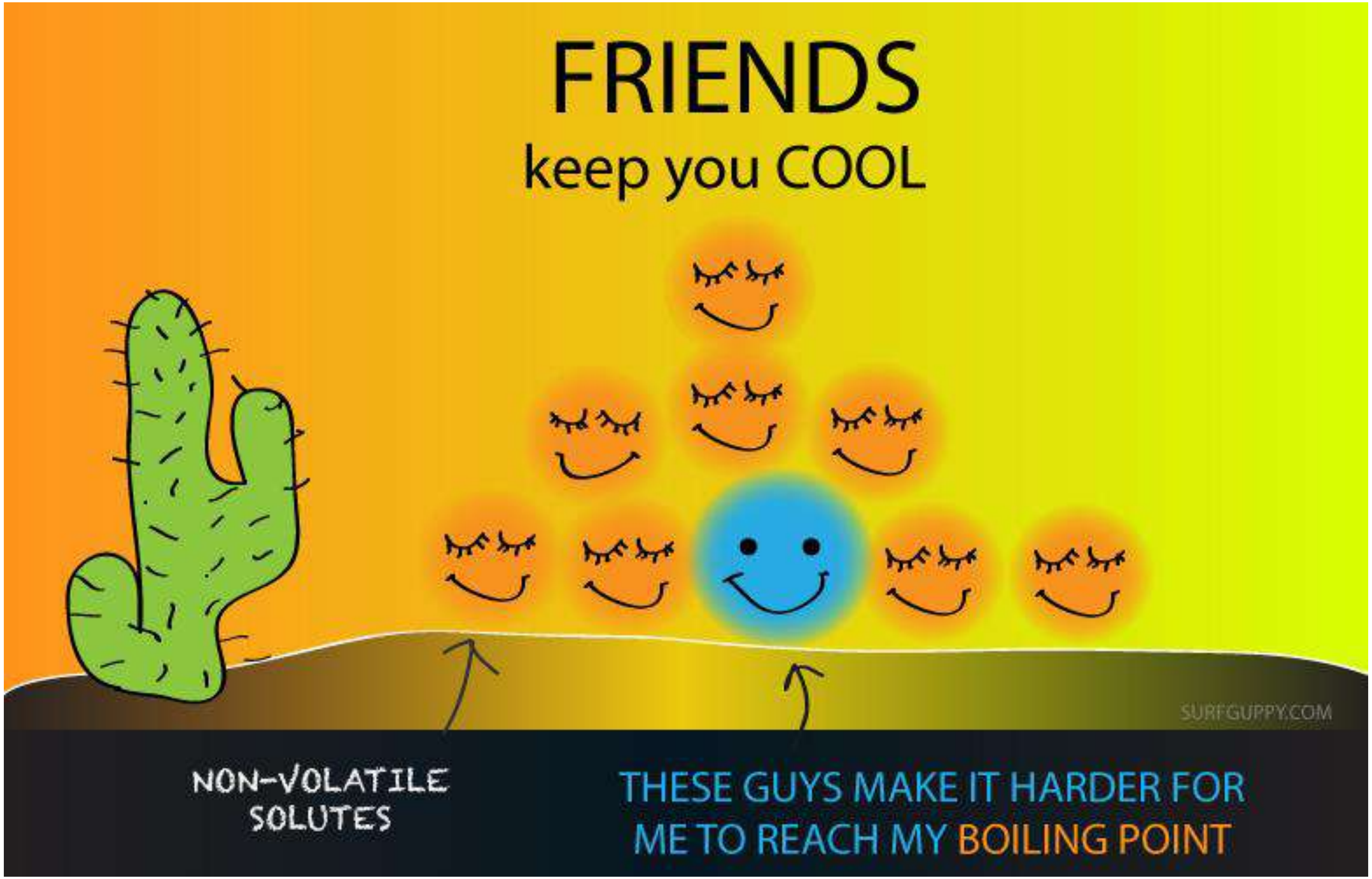
$$\text{Moles of sucrose} = n_2 = \frac{w_2}{M_2} = \frac{171.2}{342.3} = 0.500$$

$$\text{Moles of water} = n_1 = \frac{w_1}{M_1} = 1000/18.02 = 55.5$$

$$\frac{\Delta p}{p_1{}^{\circ}} = X_2 = \frac{n_2}{n_1 + n_2}$$

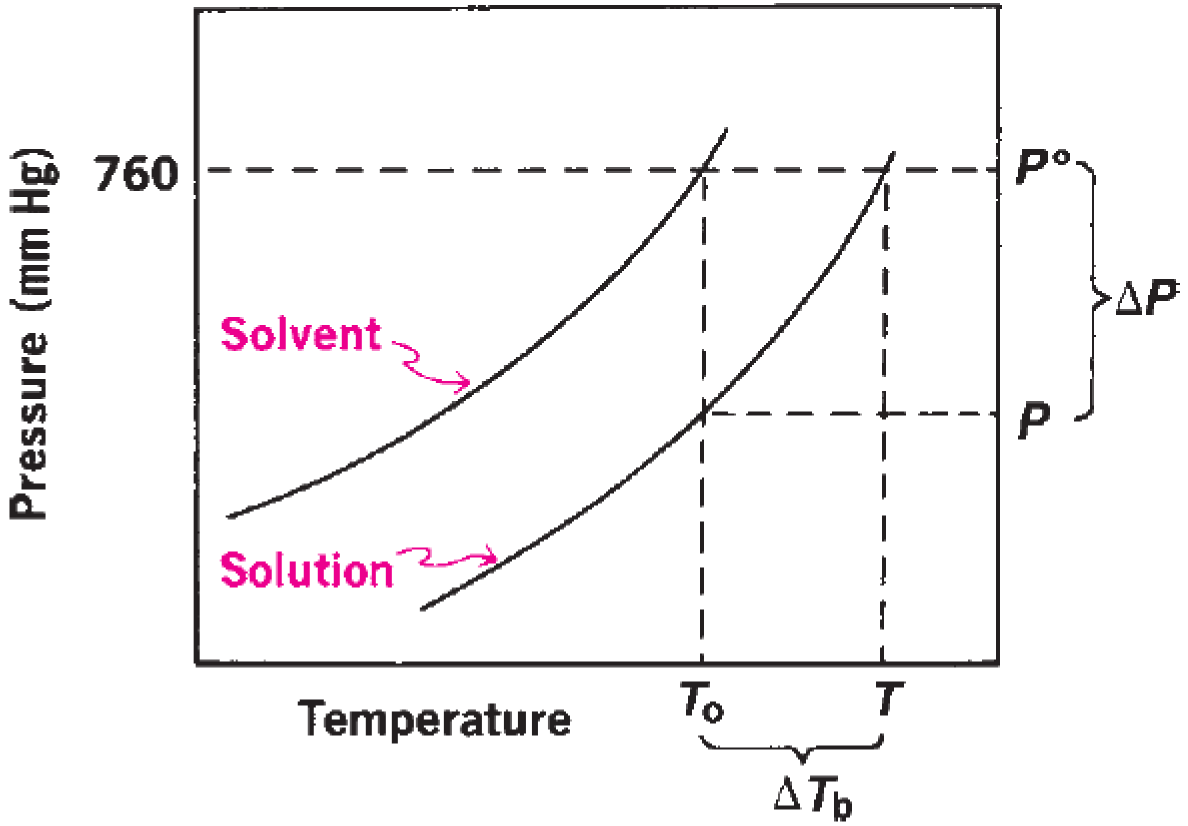
$$\frac{\Delta p}{p_1{}^{\circ}} = \frac{0.50}{55.5 + 0.50} = 0.0089$$

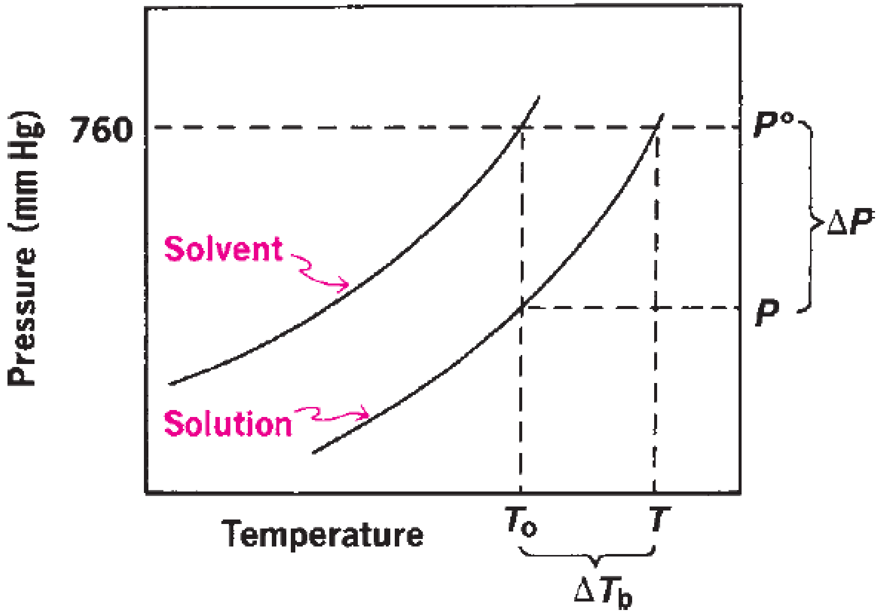
Elevation of the Boiling Point



- The normal boiling point is the temperature at which the vapor
- pressure of the liquid becomes equal to an external pressure of 760 mm Hg.
- A solution will boil at a higher temperature than will the pure solvent.

- The vapor pressure curve for the solution lies below that of the pure solvent, and the temperature of the solution must be elevated to a value above that of the solvent in order to reach the normal boiling point.





- The elevation of the boiling point is shown in the figure as $T - T_o = \Delta T_b$.
- The ratio of the elevation of the boiling point to the vapor pressure lowering, $\Delta p = p° - p$, at 100◦C is approximately a constant at this temperature; it is written as

$$\frac{\Delta T_b}{\Delta p} = k' \qquad \Delta T_b = k' \Delta p$$

Moreover, because $p_1°$ is a constant, the boiling point elevation may be considered proportional to $\Delta p/p_1°$, lowering of vapor pressure $\Delta T_b = kX_2$

- Because the boiling point elevation depends only on the mole fraction of the solute, it is a colligative property.
- In dilute solutions, X2 is equal approximately to m/(1000/M1)
- X2 $= \frac{n2}{n1+n2} \approx \frac{n2}{n1} = \frac{W2/M2}{1000/M1}$

$$\Delta T_b = \frac{kM_1}{1000} m$$

$$\Delta T_b = K_b m$$

- where Tb is known as the boiling point elevation and Kb is called the molal elevation constant or the ebullioscopic constant.
- Kb has a characteristic value for each solvent

EBULLIOSCOPIC (K_b) AND CRYOSCOPIC (K_f) CONSTANTS FOR VARIOUS SOLVENTS

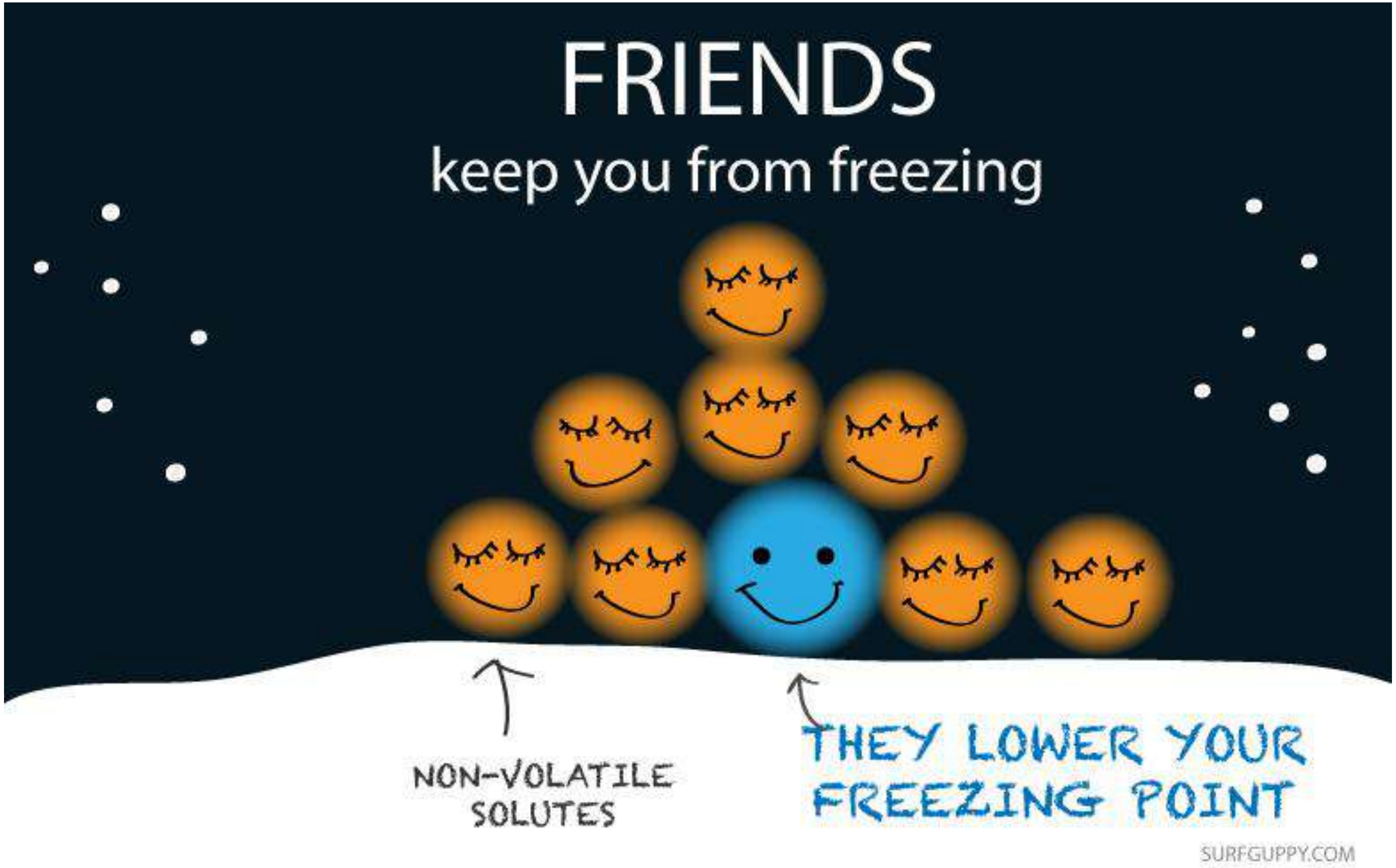
Substance	Boiling Point (°C)	K_b	Freezing Point (°C)	K_f
Acetic acid	118.0	2.93	16.7	3.9
Acetone	56.0	1.71	−94.82*	2.40*
Benzene	80.1	2.53	5.5	5.12
Camphor	208.3	5.95	178.4	37.7
Chloroform	61.2	3.54	−63.5	4.96
Ethyl alcohol	78.4	1.22	−114.49*	3*
Ethyl ether	34.6	2.02	−116.3	1.79*
Phenol	181.4	3.56	42.0	7.27
Water	100.0	0.51	0.00	1.86

EXAMPLE

- A 0.200 m aqueous solution of a drug gave a boiling point elevation of 0.103◦C. Calculate the approximate molal elevation constant for the solvent, water.

$$K_b = \frac{\Delta T_b}{m} = \frac{0.103}{0.200} = 0.515 \text{ deg kg/mole}$$

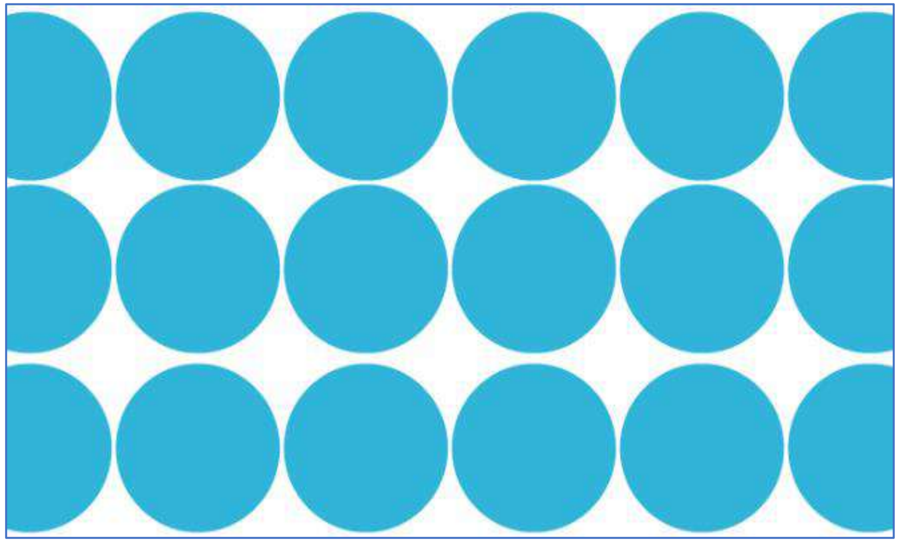
Freezing point depression

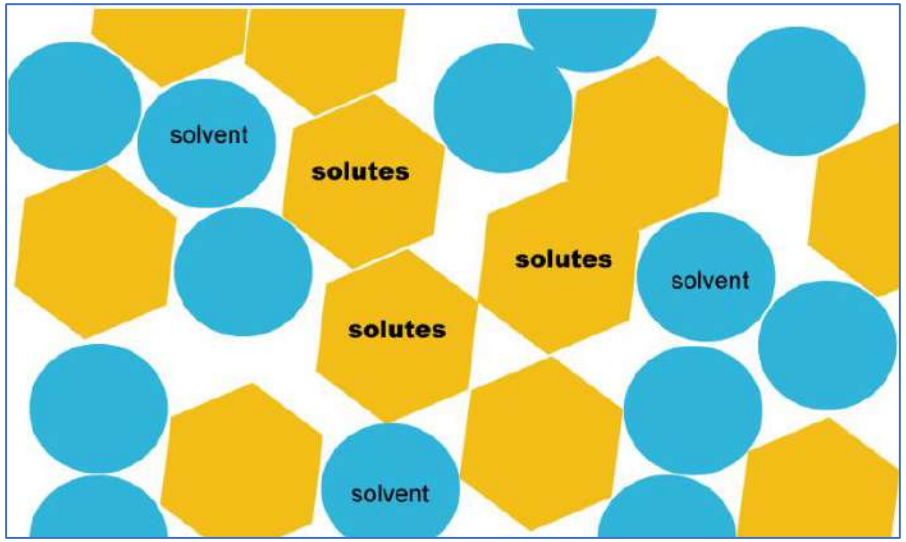


- The normal freezing point or melting point of a pure compound is the temperature at which the solid and the liquid phases are in equilibrium under a pressure of 1 atm.

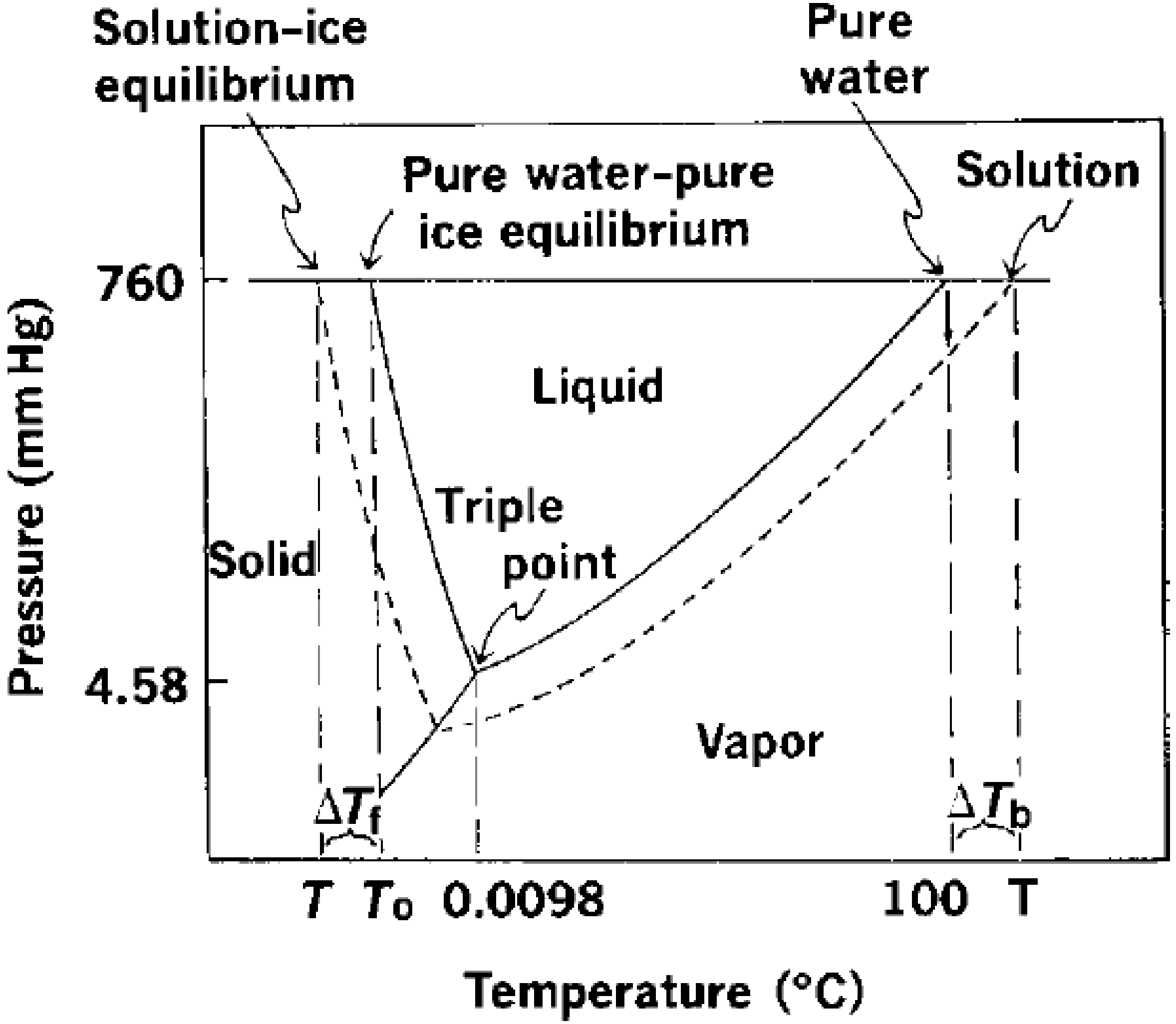
- the freezing point of a solution is always lower than that of the pure solvent.

- That is because the solute prevents the molecules from getting together into an organized structure and the temperature has to fall a lot lower to get rid of all the kinetic energy that causes the molecules to bounce around.





- The more concentrated the solution, the farther apart are the solvent and the solution curves in the diagram and the greater is the freezing point depression.
- Accordingly, a situation exists analogous to that described for the boiling point elevation, and the freezing point depression is proportional to the molal concentration of the solute.



$$\Delta T_f = K_f m$$

$$\Delta T_f = K_f \frac{1000 w_2}{w_1 M_2}$$

- ΔTf is the freezing point depression, and Kf is the molal depression constant or the cryoscopic constant, which depends on the physical and chemical properties of the solvent.
- The freezing point depression of a solvent is a function only of the number of particles in the solution, and for this reason it is referred to as a colligative property.

EXAMPLE

What is the freezing point of a solution containing 3.42 g of sucrose and 500 g of water? The molecular weight of sucrose is 342. In this relatively dilute solution, K_f is approximately equal to 1.86.

$$\Delta T_f = K_f m = K_f \frac{1000\, w_2}{w_1 M_2}$$

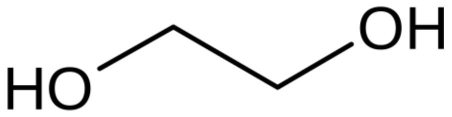
$$\Delta T_f = 1.86 \times \frac{1000 \times 3.42}{500 \times 342}$$

$$\Delta T_f = 0.037^\circ\mathrm{C}$$

Therefore, the freezing point of the aqueous solution is -0.037°C.

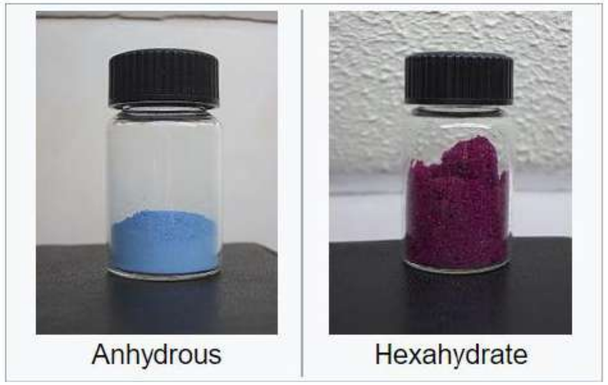
Ethylene glycol is a chemical commonly used in many commercial and industrial applications including **antifreeze** and **coolant**.

Ethylene glycol helps keep your car's engine from freezing in the winter and acts as a coolant to reduce overheating in the summer



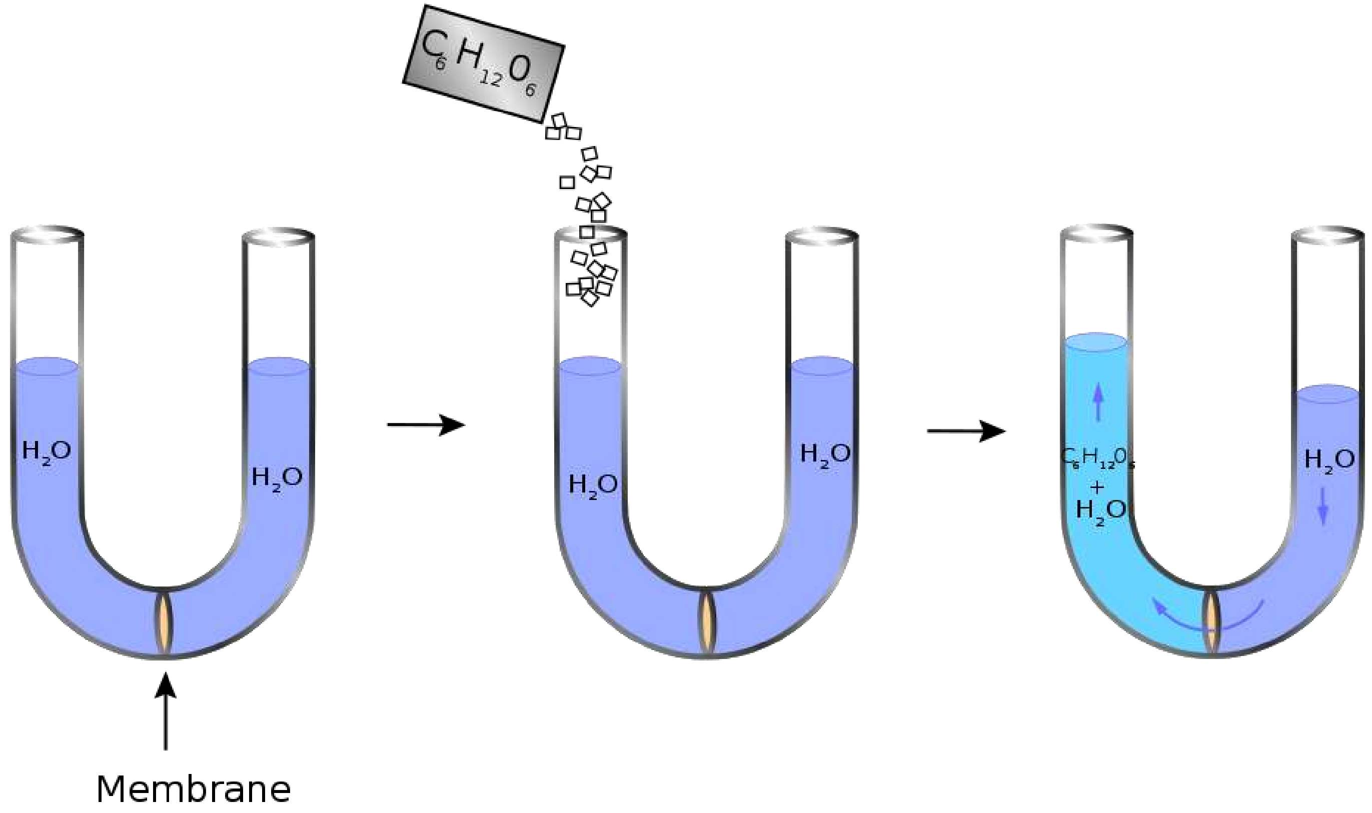
Osmotic Pressure

Cobalt(II) chloride



- If cobalt chloride is placed in a parchment sac and suspended in a beaker of water, the water gradually becomes red as the solute diffuses throughout the vessel.
- In this process of diffusion, both the solvent and the solute molecules migrate freely.
- On the other hand, if the solution is confined in a membrane permeable only to the solvent molecules, the phenomenon known as **osmosis** (Greek: “a push or impulse”) occurs, and the barrier that permits only the molecules of one of the components (usually water) to pass through is known as a semipermeable membrane.

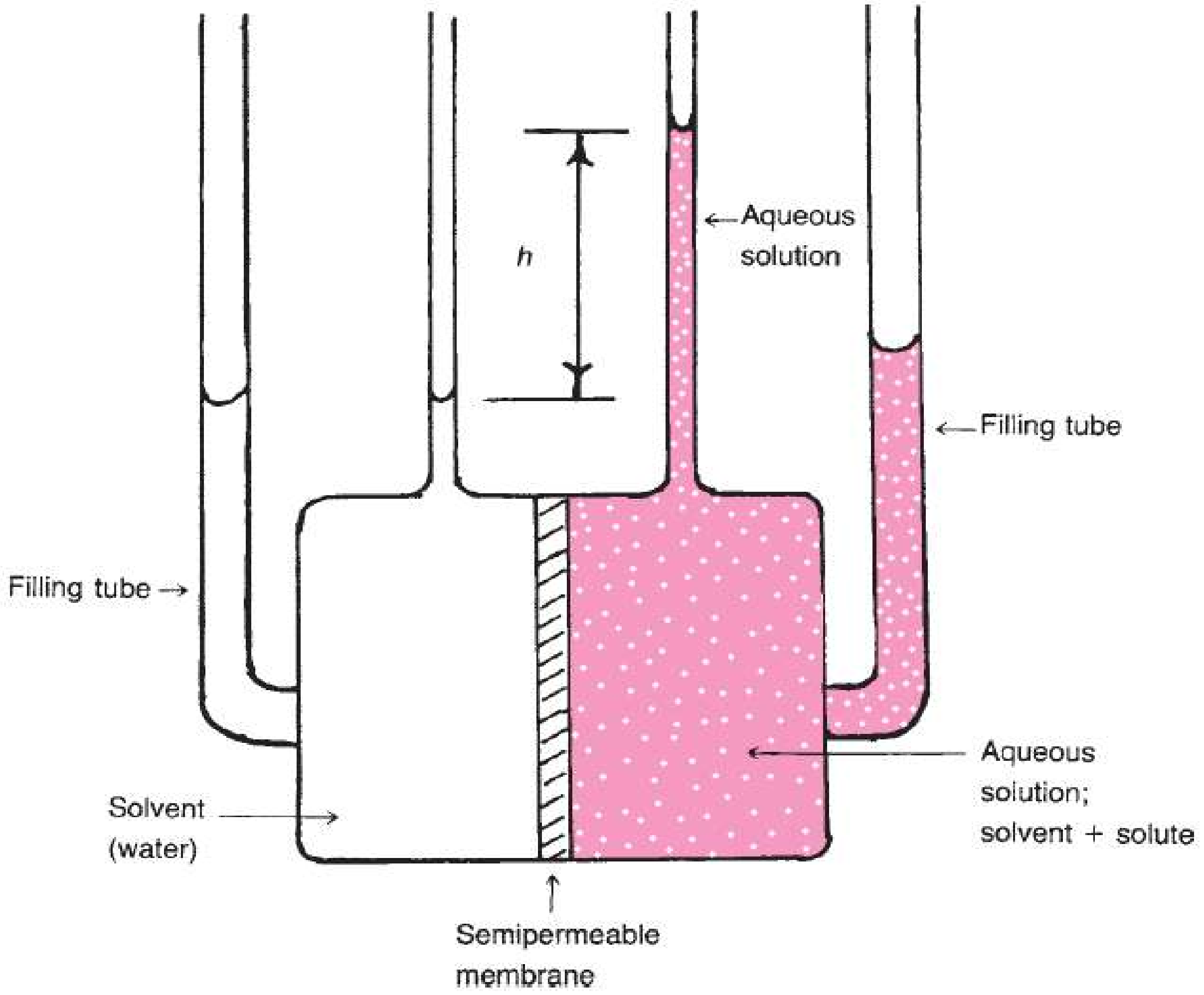
- Osmosis is the passage of the solvent into a solution through a semipermeable membrane.
- Osmotic pressure is the minimum pressure which needs to be applied to a solution to prevent the inward flow of its pure solvent across a semipermeable membrane.



- Solvent therefore passes spontaneously into the solution until the chemical potentials of solvent and solution are equal.
- Consider osmosis in terms of the following sequence of events.
- (*a*) The addition of a nonvolatile solute to the solvent forms a solution in which the vapor pressure of the solvent is reduced (see Raoult’s law).
- (*b*) If pure solvent is now placed adjacent to the solution but separated from it by a semipermeable membrane, solvent molecules will pass through the membrane into the solution in an attempt to dilute out the solute and
- raise the vapor pressure back to its original value (namely, that of the original solvent).
- (*c*) The osmotic pressure that is set up as a result of this passage of solvent molecules can be determined either by measuring the hydrostatic head appearing in the solution or by applying a known pressure that just balances the osmotic pressure and prevents any net movement of solvent molecules into the solution.

Osmotic pressure osmometer

osmotic pressure π (atm)=Height h ×Solution density ρ × Gravity acceleration



- The osmotic pressure thus obtained is proportional to the reduction in vapor pressure brought about by the concentration of solute present.
- Because this is a function of the molecular weight of the solute, osmotic pressure is a colligative property

van't Hoff and Morse Equations for Osmotic Pressure

- van't Hoff recognized the proportionality between osmotic pressure, concentration, and temperature, suggested a relationship that corresponded to the equation for an ideal gas.

$$\pi V = nRT$$

- where π is the osmotic pressure in atm, V is the volume of the solution in liters, n is the number of moles of solute, R is the gas constant, equal to 0.082 liter atm/mole deg, and T is the absolute temperature.

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

- where c is the concentration of the solute in moles/liter (molarity).

- Morse and others have shown that when the concentration is expressed in molality rather than in molarity, the results compare more nearly with the experimental findings.
- The Morse equation is

$$\pi = RTm$$

EXAMPLE

One gram of sucrose, molecular weight 342, is dissolved in 100 mL of solution at 25°C. What is the osmotic pressure of the solution?

$$\begin{aligned} \text{Moles of sucrose} &= \frac{1.0}{342} = 0.0029 \\ \pi \times 0.10 &= 0.0029 \times 0.082 \times 298 \\ \pi &= 0.71\,\text{atm} \end{aligned}$$

MOLECULAR WEIGHT DETERMINATION

- The four colligative properties that have been discussed in this chapter—vapor pressure lowering, freezing point lowering, boiling point elevation, and osmotic pressure—can be used to calculate the molecular weights of nonelectrolytes present as solutes.

1- using vapor pressure lowering

$$\frac{p_1^\circ - p_1}{p_1^\circ} = \frac{n_2}{n_1 + n_2} = \frac{w_2/M_2}{(w_1/M_1) + (w_2/M_2)}$$

In dilute solutions

$$\frac{\Delta p}{p_1^\circ} = \frac{w_2/M_2}{w_1/M_1} \qquad M_2 = \frac{w_2 M_1 p_1^\circ}{w_1 \Delta p}$$

2- using boiling point elevation

$$\Delta T_b = K_b m$$

$$\Delta T_b = K_b \frac{1000 w_2}{w_1 M_2}$$

$$M_2 = K_b \frac{1000 w_2}{w_1 \Delta T_b}$$

3- using freezing point depression

$$M_2 = K_f \frac{1000 w_2}{\Delta T_f w_1}$$

EXAMPLE

A solution containing 10.0 g of sucrose dissolved in 100 g of water has a boiling point of 100.149°C. What is the molecular weight of sucrose?

$$M_2 = 0.51 \times \frac{1000 \times 10.0}{100 \times 0.149}$$

$$= 342 \text{ g/mole}$$

EXAMPLE

The freezing point depression of a solution of 2.000 g of 1,3-dinitrobenzene in 100.0 g of benzene was determined by the equilibrium method and was found to be 0.6095°C. Calculate the molecular weight of 1,3-dinitrobenzene.

$$M_2 = 5.12 \times \frac{1000 \times 2.000}{0.6095 \times 100.0} = 168.0 \text{ g/mole}$$

EXAMPLE

Fifteen grams of a new drug dissolved in water to yield 1000 mL of solution at 25°C was found to produce an osmotic pressure of 0.6 atm. What is the molecular weight of the solute?

$$\pi = cRT = \frac{c_g RT}{M_2}$$

where c_g is in g/liter of solution. Thus,

$$\pi = \frac{15 \times 0.0821 \times 298}{M_2}$$

or

$$M_2 = \frac{15 \times 24.45}{0.6} = 612 \text{ g/mole}$$