potential of the solvent in a binary solution is given by Equation 3.24. The subscripts outside the bracket on the left-hand side indicate that temperature, pressure and amount of component 1 (the solute in this case) remain constant:

$$\left(\frac{\partial G}{\partial n_2}\right)_{T,P,n_1} = \bar{G}_2 = \mu_2 \quad (3.24)$$

Since (by definition) only solvent molecules can pass through a semi-permeable membrane, the driving force for osmosis arises from the inequality of the chemical potentials of the solvent on opposing sides of the membrane. Thus the direction of osmotic flow is from the dilute solution (or pure solvent), where the chemical potential of the solvent is highest because of the higher concentration of solvent molecules, into the concentrated solution, where the concentration and consequently the chemical potential of the solvent are reduced by the presence of more solute. The chemical potential of the solvent in the more concentrated solution can be increased by forcing its molecules closer together under the influence of an externally applied pressure. Osmosis can be prevented by such means, hence the term *osmotic pressure*.

The relationship between osmotic pressure (π) and concentration of a non-electrolyte is defined for dilute solutions, which may be assumed to exhibit ideal behaviour, by the van't Hoff equation (Eqn 3.25):

$$\pi V = n_2 RT \quad (3.25)$$

where V is the volume of solution, n_2 is the number of moles of solute, T is the absolute temperature and R is the gas constant. This equation, which is similar to the ideal gas equation, was derived empirically but it does correspond to a theoretically derived equation if approximations based on low solute concentrations are taken into account.

If the solute is an electrolyte, Equation 3.25 must be modified to allow for the effect of ionic dissociation, because this will increase the number of particles in the solution. This modification is achieved by insertion of the van't Hoff correction factor (i) to give:

$$\pi V = i n_2 RT \quad (3.26)$$

$$\text{where } i = \frac{\text{observed colligative property}}{\text{colligative property expected if dissociation did not occur}}$$

Osmolality and osmolarity

The amount of osmotically active particles in a solution is sometimes expressed in terms of osmoles or milliosmoles. These osmotically active particles may be either molecules or ions. Osmole values depend on the number of particles dissolved in a solution, regardless of charge. For substances that maintain their molecular structure when they dissolve (e.g. glucose), osmolarity and the molarity are essentially the same. For substances that dissociate when they dissolve, the osmolarity is the number of free particles times the molarity. Thus a 1 molar solution of pure NaCl solution would be 2 osmolar (1 for Na^+ and 1 for Cl^-).

The concentration of a solution may therefore be expressed in terms of its *osmolarity* or its *osmolality*. Osmolarity is the number of osmoles per litre of solution and osmolality is the number of osmoles per kilogram of solvent.

Isoosmotic solutions

If two solutions are separated by a perfect semi-permeable membrane, i.e. a membrane which is permeable only to solvent molecules, and no net movement of solvent occurs across the membrane, then the solutions are said to be *isoosmotic* and have equal osmotic pressures.

Isotonic solutions

Biological membranes do not always function as perfect semi-permeable membranes and some solute molecules in addition to water are able to pass through them. If two isoosmotic solutions remain in osmotic equilibrium when separated by a biological membrane, they may be described as being *isotonic* with respect to that particular membrane.

Adjustment of isotonicity is particularly important for formulations intended for parenteral routes of administration (this is discussed in Chapter 36). Excessively hypotonic or hypertonic solutions can cause biological damage.

Diffusion in solution

The components of a solution, by definition, form a homogeneous single phase. This homogeneity arises