



## Review

# Basic concepts and practical equations on osmolality: Biochemical approach



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## ARTICLE INFO

## Article history:

Received 12 January 2016

Received in revised form 2 June 2016

Accepted 3 June 2016

Available online 22 June 2016

## Keywords:

Osmolarity

Osmosis

Osmotic pressure

Osmolar gap

Tonicity

## ABSTRACT

The terms osmotic pressure, osmotic coefficient, osmole, osmolarity, osmolality, effective osmolality and delta osmolality are formally defined. Osmole is unit of the amount of substance, one mole of nonionized impermeant solute is one osmole. Assuming an ideal solution, osmotic pressure ( $\pi$ ) in mmHg is 19.3 times the osmolarity. Osmolarity is defined as the number of milliosmoles of the solutes per liter of solution. Suitable equations are presented for the rapid calculation of the osmolarity of different solutions. The concentrations of electrolytes are expressed by mEq/L that is, equal to their osmolarity as mOsm/L. If the solute concentration (C) is expressed as mg/L, mg/dL and g%, osmolarity is calculated as:  $C.n'/MW$ ,  $C.n'(10)/MW$  and  $C.n'(10^4)/MW$  respectively. Osmolality is milliosmoles of solutes per one kilogram (or liter) of water of solution (plasma) and is calculated by osmolarity divided to plasma water. The osmolal concentration is corrected to osmolal activity by using the osmotic coefficient,  $\phi$ . The salts of sodium (chloride and bicarbonate) and nonelectrolyte glucose and urea are the major five osmoles of plasma. The equation:  $Posm = 2 [Na^+] + glucose (mg/dL)/18 + BUN (mg/dL)/2.8$  is also the simplest and best formula to calculate plasma osmolality. The concentration of only effective osmoles evaluates effective osmolality or tonicity as:  $Eosm = 2 [Na^+] + glucose/18$ . The normal range of plasma tonicity is 275–295 mOsm/kg of water. The difference between the measured and calculated osmolality is called osmolal gap. It is recommended to withdraw the formula of Dorwart–Chalmers from the textbooks and autoanalyzers and to use the simplest equation of Worthley *et al.* as the best equation for calculating serum osmolality. Furthermore the normal ranges of osmolal gap also must be corrected to  $0 \pm 2$  mOsm/L.

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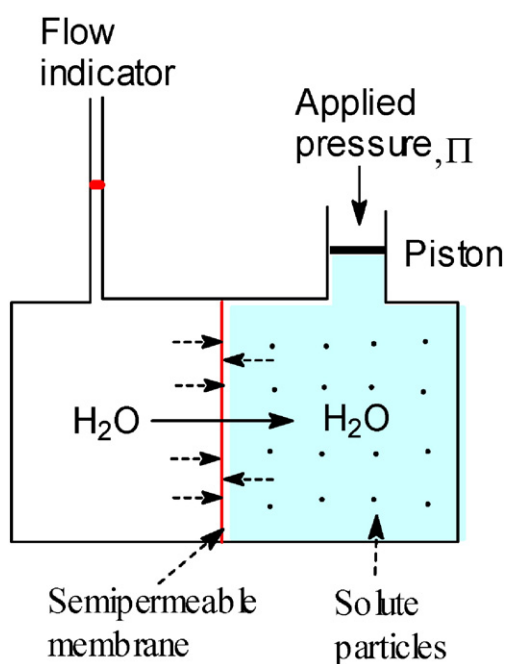
## 1. Introduction

The balance and disorders of water and electrolytes have a central position in clinical sciences [1]. Most of the equations in this field have been deduced on the basis of the osmosis phenomenon [2]. The terms involved in the issue of osmosis are often used by professional people including authors, physicians, pharmacists, nurses, dietitians and clinical chemists. The wide using of various therapeutic and nutritional fluids needs to use correct terminology. Despite its importance, the subject has not been reviewed in recent decades [3]. In the present study, the basic concepts in osmosis were surveyed and the practical equations and their applications are listed.

### 1.1. Osmosis

When a solution is divided into two compartments by a membrane which is permeable to water molecules and impermeable to solutes, the membrane is called selective- or semi-permeable. If the solute concentration is not equal on the two sides of a semi-permeable membrane, the water molecules will migrate to the more concentrated side [4]. The net diffusion of water molecules due to the difference in solutes concentrations is called 'water osmosis'. The osmosis of water molecules causes a hydrostatic pressure in the solvent side called 'osmotic pressure',  $\pi$  (Fig. 1).

The phenomenon of osmosis can be studied by both kinetic and thermodynamic approaches [2]. The molecular kinetic theory of gases has also been applied to solutions. In the kinetic model, osmotic flow is caused by the difference in the number of water molecules striking the membrane from the solution and the solvent sides [5]. A higher number of water molecules strike the membrane in solvent side determines the net direction of water flow. Kinetic derivation of the equation



**Fig. 1.** Schematic diagram for an apparatus to measure osmotic pressure. There is the solvent (pure water) in the left side and a solution with a defined osmolarity in the right side. The water osmosis has been shown by arrows and causes osmotic pressure in the right side.

for osmotic pressure gives:

$$\pi = \frac{RT}{\bar{v}_i} X_2$$

In which  $X_2$  is molar fraction of the solute and  $\bar{v}_i$  is partial molar volume of the solvent [2].

In the thermodynamic model, the equations are derived on the basis of the equilibrium of the solvent, that is, water to which the membrane is permeable [6–9]. The difference in solute concentration causes a chemical potential ( $\mu$ ) for solute across the membrane from the high toward the low level. Since the solute can not pass across the membrane, the water molecules will travel toward the more concentrated side. According to Gibbs-Duhem equation [6–9], there is an inverse relationship between the chemical potentials (or the activities,  $a$ ) of the solute and solvent. This means that the activity and chemical potential of water molecules is higher in the side in which the activity of solute is lower. So, the movement of water molecules is favorable thermodynamically from the solvent to the solution side. In general, the chemical potential ( $\mu$ ) is related to the pressure ( $P$ ) by the equation:

$$\left(\frac{\partial \mu}{\partial P}\right)_{T, X_i} = \bar{v}_i \quad \text{or} \quad \pi = -\frac{RT}{\bar{v}_i} \ln X_1$$

In which  $X_1$  is molar fraction of the solvent [8,9].

### 1.2. Osmotic pressure ( $\pi$ ) and osmotic coefficient ( $\varphi$ )

Both the kinetic [2,5] and thermodynamic models [6–9] finally lead to the same equation::

$$\pi V = nRT$$

In which,  $\pi$  is the osmotic pressure in mmHg,  $V$  is the volume of solution in liter,  $n$  is the number of moles of the solute and  $T$  is absolute temperature in Kelvin ( $^{\circ}\text{K}$ ).  $R$  is the gas constant and is 2 cal/mol. $^{\circ}\text{K}$  or as a more convenient unit is equal to 62.36 mmHg.L/mol. $^{\circ}\text{K}$  [8]. The equation is analogous to the state function of an ideal gas ( $PV = nRT$ ). For an ideal solution:

$$\pi V/nRT = 1$$

The 'osmotic coefficient' ( $\varphi$ ) is applied to quantify the deviation from ideality:

$$\pi V/nRT = \varphi$$

The osmotic coefficient is dependent on the nature and molality (ionic strength) of the solute and temperature [9]. Several methods have been described to calculate osmotic coefficients for different solutes [9–11]. The osmotic coefficient of NaCl in ionic strength comparable to plasma is about 0.93 and for non-electrolytes glucose and urea is about 1.0 [8–10]. By definition, for a particle the number of moles per volume i.e. molarity is the same as osmolarity:

$$\pi = \varphi \cdot (n/V) \cdot RT = \varphi \cdot \text{Osmolarity} \cdot RT$$

$$\pi = \varphi \cdot \text{Osmolarity} \cdot (62.36 \text{ mmHg.L/mol.}^{\circ}\text{K}) (37 + 270.15 \text{ }^{\circ}\text{K}) / 1000$$

$$\pi(\text{mmHg}) = 19.3 \varphi \text{ Osmolarity (mOsm/L)}$$

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