Osmolarity, Tonicity, and Equivalents Study Guide & Practice

1. Osmolarity

Recall that the osmotic pressure of a solution is created by solutes dissolved in a solvent, and the dissolved solutes create a ‘pull’ for water to move INTO that solution by osmosis.

Osmolarity is simply another measure of concentration, like molarity. (See Molarity Study Guide Handout for more on molarity.) However, unlike molarity which tells us the concentration of one type of molecule, osmolarity allows us to look at the concentration of multiple solutes dissolved in a solution. By now, you can certainly appreciate that body fluids have many more than just one type of particle dissolved in them. So when talking about body fluids it is much better to use osmolarity.

The more solute dissolved in a solution, the more osmotic pressure that solution has, and the higher the osmolarity of the solution.

The following is important to understand: The osmolarity of a solution depends upon the total NUMBER, NOT THE TYPE, of particles dissolved in a solvent. What do we mean by particles? In terms of osmolarity, particles can be atoms, ions, molecules, or even cells in solution without regard to molecular weight or electrical charge of the individual particles.

So, 1 particle (or mole or millimole) of glucose would be equivalent to 1 particle (or mole, or millimole) of Na, or 1 particle of Cl, or 1 particle of Mg, or anything else in solution. The type of particle does NOT matter; only the NUMBER of particles in solution contributes to osmolarity.

Covalently bonded molecules, e.g., glucose, proteins, lipids, do not break apart when dissolved in aqueous solvent. (They are called nonelectrolytes because they do not generate an electrical charge when dissolved in solution.) However, ionically bonded compounds like NaCl, CaCl₂, and NaHCO₃ do break apart into ions when dissolved in aqueous solutions, i.e., they are electrolytes. (Recall that elements on the left side of the periodic table generally will form ionic bonds with elements on the right side of the periodic table.) This is what happens to them in solution:

\[
\begin{align*}
\text{NaCl} & \rightarrow \text{Na}^+ + \text{Cl}^- \quad (1 \text{ particle of NaCl dissociates into TWO particles in solution}) \\
\text{CaCl}_2 & \rightarrow \text{Ca}^{2+} + 2 \text{Cl}^- \quad (1 \text{ particle of CaCl}_2 \text{ dissociates into THREE particles in solution}) \\
\text{NaHCO}_3 & \rightarrow \text{Na}^+ + \text{HCO}_3^- \quad (1 \text{ particle of NaHCO}_3 \text{ dissociates into TWO particles in solution})
\end{align*}
\]
Osmolarity is measured in osmoles (Osm) / L of solution. (1/1000 of an Osm is a milliosmole, or mOsm) A 5 Osm/L solution would have a higher osmotic pressure than a 2 Osm/L solution.

In each of the reactions below, one mole of a substance goes into solution. How many moles of particles does each reaction produce? (That is, for each particle that goes into solution, how many particles result from dissolving that substance?)

NaCl → Na⁺ + Cl⁻
CaCl₂ → Ca²⁺ + 2 Cl⁻
NaHCO₃ → Na⁺ + HCO₃⁻

**IMPORTANT:** The # of osmoles in a solution is ALWAYS greater than, or equal to, the number of moles/L of the original solute, i.e., Osm/L ≥ moles/L. (Use this rule on the equations above to see for yourself that it’s true.)

In order to calculate the osmolarity of a solution, simply take the number of moles/L (or molarity, M) of the compound (solute) to be dissolved and MULTIPLY by the number of particles into which the dissolved compound dissociates. So,

\[
Osm/L = \text{mol/L} \times \# \text{ particles yielded upon dissociation}
\]

**Example 1:** The osmolarity of a 2 M NaCl solution would be:

\[
Osm/L = 2 \text{ M NaCl} \times 2 \text{ particles (see dissociation equations above)} = 4 \text{ Osm/L NaCl}
\]

**Example 2:** The osmolarity of a 2 mM CaCl₂ solution would be:

\[
Osm/L = 2 \text{ mM CaCl₂} \times 3 \text{ particles (see dissociation equations above)} = 6 \text{ mOsm/L CaCl₂}
\]

The osmolarity of human body fluids is about 290-300 mOsm/L. Most of this (about 270 mOsm) is due to Na⁺, Cl⁻, and HCO₃⁻. Other solutes that make contributions to this osmolarity include: glucose, urea, plasma proteins, and other cations/anions.
2. Tonicity of body fluids

Tonicity is a measure of the difference in osmotic pressure between two solutions, e.g., the solution inside a cell vs. the fluid outside the cell. **Tonicity deals with the relative osmotic pressures of different solutions.**

**IMPORTANT:** Water will always move from a solution with lower osmotic pressure into a solution with a higher osmotic pressure. [Remember that there is more water in a solution with lower osmotic pressure, so it is just diffusing (by osmosis) down its concentration gradient according to the law of diffusion.]

An **isotonic** solution is one in which two solutions have the same osmolarity (osmotic pressure) so, there is NO NET MOVEMENT OF WATER between the solutions. This is an equilibrium situation, so there will be movement of water between the two solutions, but the number of water molecules that move one way will equal the number that move the other way. Thus we say there is no NET movement of water.

Two important isotonic solutions you should already know are: 0.9% NaCl and 5.0% glucose (dextrose).

A **hypertonic** solution is a solution that has a higher osmolarity (osmotic pressure) than a second solution (which would be hypotonic to it – see below): **there is a net movement of water INTO the hypertonic solution** (see the rule above about water always moving from a solution with lower osmotic to a solution with higher osmotic pressure).

A **hypotonic** solution is a solution that has a lower osmolarity (osmotic pressure) than a second solution (which would be hypertonic to it): **there is a net movement of water OUT OF the hypotonic solution into the hypertonic solution** (see the rule above about water always moving from a solution with lower osmotic to a solution with higher osmotic pressure).

**Example 1:** What would happen to an erythrocyte (RBC) if it is placed in a hypertonic solution? There would be a net movement of water from the hypotonic solution (inside of the RBC) OUT OF the RBC and the RBC would shrink (crenate).

**Example 2:** What would happen if an erythrocyte (RBC) is placed in a hypotonic solution? There would be a net movement of water from the hypotonic solution (outside of the RBC) INTO the RBC and the RBC would swell, and possibly lyse.
3. Equivalents (Equivalent weights)

Recall that osmolarity is a measure of the osmotic pressure of a solution created by solutes dissolved in a solvent. Solutes create a ‘pull’ for water to move INTO that solution by osmosis. Also recall that osmolarity, rather than molarity, is used to indicate the total number of particles produced by a multiple solutes when they are dissolved in a solvent (typically H₂O), regardless of the types of particles. In other words, we don’t care if the particle is charged or not. If it’s in solution, it contributes to osmolarity (and tonicity).

However, when assessing the body’s electrolytes (charged particles that dissociate in water), it becomes important to know how many positive and how many negative charges there are in solution, not merely the total number of particles. Many processes in the body depend upon the right balance of charges in order to function correctly, e.g., nerve and muscle cells so this balance is crucial for homeostasis. In the body, the total number of positive charges (contributed by cations) will equal the total number of negative charges (contributed by anions).

As an illustration of this, a Na⁺ ion (+1 charge) in solution has only half the charge of a Ca²⁺ ion (+2 charge), but both particles are identical in terms of their effect on osmolarity. So you can see that there is an important difference when we are assessing osmolarity vs. when we are assessing charge.

We measure the number of charges of a particular ion in a solution by using something called its equivalent weight per volume of solute or, more simply, equivalents/L. The term Equivalent is abbreviated as ‘Eq’, so we normally state an ion’s charge in solution as Eq/L.

Since the number of charges contributed by any particular ion in the body is very small, we generally use the term milliequivalents (1/1000 of an Equivalent) instead of Equivalents. Milliequivalent can be abbreviated as mEq, and concentrations of an ion’s charges in the body are usually stated as mEq/L.

Formally, the definition of an equivalent is: something that has the same amount of charge as 1 mol of H⁺.

For example, since 1 mole of Na⁺ and one mole of H⁺ have the same number of particles (recall the definition of a mole) and each particle has a +1 charge, we can say that 1 mole of Na⁺ = 1 Eq of Na⁺. (We can say this because it has the same amount of charge as 1 mol of H⁺.)

Let’s look at an ion that doesn’t have just a +1 charge: the Ca²⁺ ion. Now, 1 mole of Ca²⁺ has the same number of particles as 1 mole of H⁺, but is there an equal amount of charge? Definitely not! There are twice as many charges in a mole of Ca²⁺ as there are in a mole of H⁺.
So, how do we make the two ‘equivalent’? Easy. Just cut the amount of Ca\(^{2+}\) in half, i.e., use half a mole of Ca\(^{2+}\). We can say, therefore, that 0.5 mol of Ca\(^{2+}\) = 1 Eq. of Ca\(^{2+}\) (has the same amount of charge as 1 mole of H\(^{+}\)).

Here’s another important point: When we’re trying to determine the number of equivalents for an ion in solution, it doesn’t matter whether the charge is positive or negative; just that it IS a charge. In other words, the sign of the charge doesn’t matter, only the magnitude of the charge is important.

For example, since one mole of Cl\(^{-}\) and one mole of H\(^{+}\) have the same number of particles and each particle has one charge (any charge!), we can say that 1 mole of Cl\(^{-}\) = 1 Eq of Cl\(^{-}\). Of course, if we’re looking to balance charges, well that’s something else again. Then we would have to look at the number of positive vs. the number of negative charges. But remember that all we’re talking about here is determining the number of equivalents for an ion in solution.

**IMPORTANT:** The # of equivalents in a solution is ALWAYS greater than, or equal to, the number of moles/L of the original ion, i.e., Eq/L ≥ moles/L. Remember that we used a similar rule when talking about the Osm/L above.

In order to calculate the Eq of an ion in solution, simply take the number of moles/L (or molarity, M) of the solute to be dissolved and MULTIPLY by the number of charges on that ion when dissolved in solution, i.e., it’s valence. So,

\[
\text{Eq/L} = \text{mol/L (this is molarity, M)} \times \# \text{ ion charges when dissolved in solution}
\]

**Example 1:** How many equivalents are there in a 1 M solution of Na\(^{+}\)?

\[
\text{Eq/L} = 1 \text{ M Na}^{+} \times 1 \text{ (since the valence of Na is 1)} = 1 \text{ Eq/L Na}^{+}
\]

**Example 2:** How many equivalents are there in a 1 mM solution of Ca\(^{2+}\)?

\[
\text{Eq/L} = 1 \text{ mM Ca}^{2+} \times 2 \text{ (since the valence of Ca is 2)} = 2 \text{ mEq/L Ca}^{2+}
\]

**Example 3:** How many equivalents are there in a 3 mM solution of HCO\(_3\)^\(-\)?

\[
\text{Eq/L} = 3 \text{ mM HCO}_3^- \times 1 \text{ (since the charge of HCO}_3^- \text{ is 1)} = 3 \text{ mEq/L HCO}_3^-
\]

*Remember that when determining the number of equivalents for an ion in solution, only the magnitude of the sign matters, not the sign of the charge.)*

Try the practice problems on the following page.
Practice Problems:

1. If you put 1 mole of KCl in 2 L of H₂O, what would be the osmolarity of the solution?

2. How many moles of glucose would you need to make a 2L solution that is isotonic with the one in question 1?

3. What would be the osmolarity of the solution in question 2?

4. The plasma Ca²⁺ concentration is about 2.5 mmol/L (2.5 mM). What would be the concentration of Ca²⁺ in mEq/L?