BIOL 442 Plant Physiology  
Lab #1: Making Solutions

Report due: (no report for this lab)

The preparation and measuring of solutions is an important part of many lab activities. This short exercise is designed to re-acquaint you with the simple calculations and lab procedures involved in preparing solutions and in use of the balance.

Objectives:
- Produce a 0.5 M solution
- Dilute the 0.5 M solution to 0.3 M
- Become familiar with the osmometer
- Pipette aliquots with accuracy
- Test your pipette accuracy by weighing the solutions and calculating the volume

Materials (for each group of two people):
- KCl
- 100 ml graduated reaction flask
- 1 ml graduated pipette
- Pipette bulb (do not pipette by mouth, even if the solution seems harmless)
- Weighing boat
- 3 screw cap vials

Preparing a 0.5 M Solution

Remember that M stands for Molar (see SI units handout) and means 1 mole of the solute per one liter of solution. Strictly speaking, the "Molar" unit has become obsolete and recent usage would call for concentration units of mol/m$^3$, where 1 mole of solute per liter would be equivalent to 1000 mol/m$^3$. We do not need one liter of solution, 100 ml (or even less) will do.

Look up the mass of one mole of KCl (its molar mass). Chemical containers often list this value as F.W. (formula weight). What is the formula weight of KCl? ______________

What are the units of this value? ______________

Remember to always include units when reporting a value. This will help you remember what that number represents.

To calculate the mass of KCl needed, use the following formulae:
\[ \text{mol}_{\text{KCl}} = C_{\text{KCl}} \times V_{\text{soln}} \]

\[ \text{mass of KCl} = \text{mol}_{\text{KCl}} \times \text{molar mass} \]

In this case, you are to make a solution with a KCl concentration \(C_{\text{KCl}}\) of 0.5 M. Make a volume \(V_{\text{soln}}\) of 100 ml (i.e., 0.1 L). When mixing the KCl and water, be certain that all of the powder is dissolved very well. If some of the KCl crystals are not dissolved, you will not produce the desired results.

**Dilution of the 0.5 M solution to a 0.3 M solution**

To create a 0.3 M solution from the 0.5 M solution, we need to dilute the original 0.5 M solution. Use the following formula:

\[ C_{\text{final}} \times V_{\text{final}} = C_{\text{initial}} \times V_{\text{initial}} \]

We want 5 ml of 0.3 M KCl. These values are the final volume and concentration, respectively. What is the initial concentration of the solution? _________________

What volume of that initial solution must you mix with water to make the final solution? _________________

Next, you will use the osmometer to check your result. The value you will receive is an osmolality (Osm) rather than a molarity (M). The instrument result (osmolality) is related to the total number of molecules in the solution: dissolving 1 mole of KCl in water would actually give you nearly 2 moles of ions because almost all of the KCl separates into individual K and Cl ions. What is the difference?

If you need help with this concept, ask me. I will provide further instruction on the use of the osmometer during the lab section.

What is the Osm of your solution?

On the last page, you will find a figure relating Osm to M for KCl. (This relationship is different for each compound. You must look up each chemical for which you wish to use this test.)

What is the Molarity of your solution? How close is this value to 0.3 M?
Pipetting Solutions and Verifying Pipetting Accuracy

Transfer 1 ml of 0.3 M KCl into pre-weighed vials (record the weight of the vial on the table below) and compare the weight increase with the volume you pipetted. To do this, you first need to look up the density of the KCl solution on the CRC Handbook table provided. Use the concentration (molarity) you determined with osmometer. The concentration of your solution will not appear on the table. You must interpolate the actual density by determining the line defined by the two concentrations and densities just surrounding your value. For those of you who don’t remember how to do this, read the next paragraph. Otherwise, skip the next paragraph.

This is called linear interpolation. Recall the relationship for a straight line: 

\[ y = mx + b \]

Let’s say you have a 2 M solution. The (molar, density) values from the table surrounding 2 M are (1.733, 1.0768) and (2.048, 1.0905). The slope is defined as the rise divided by the run, or \( \Delta y/\Delta x \). In this case, the resulting value for the slope is 0.04349. The y-intercept can be determined by solving the algebraic expression for \( b \):

\[ b = y - mx \]

where the slope was just calculated as \( m \), \( y \) is the molarity of one of the points used to calculate the slope and \( x \) is the density that corresponds to the molarity you used as \( y \). In this case, the \( y \)-intercept, \( b \), is 1.0014. Now, you put the concentration (2 M in this example) into the point-slope form of the line equation \( (y = mx + b) \) to calculate the density, in g/ml. A 2 M solution of KCl would have a density of 1.0884 g/ml.

To determine the mass of your solution, weigh the vial with the solution within it and subtract the mass of the empty vial from the mass of the vial + solution. Lastly, you will use the following relationship to calculate the actual volume of solution you pipetted into the vial:

\[ \text{Volume (ml)} = \frac{\text{Mass (g)}}{\text{Density (g/ml)}} \]

<table>
<thead>
<tr>
<th></th>
<th>Empty vial (g)</th>
<th>Vial + Aliquot (g)</th>
<th>Aliquot (g)</th>
<th>Volume</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vial 1</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Vial 2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Vial 3</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Osmotic properties of KCl Solutions

Curve fit to data:

\[ M = 0.5543 Osm - 0.0019 \]

or

\[ Osm = \frac{M + 0.0019}{0.5543} \]
Find the density of a 0.3 M KCl solution. Consulting a density table gives the following values; notice that the table does not include exactly 0.3 M:

<table>
<thead>
<tr>
<th>M</th>
<th>Density (g/cm$^3$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.271</td>
<td>1.0110</td>
</tr>
<tr>
<td>0.409</td>
<td>1.0174</td>
</tr>
</tbody>
</table>

So let’s use a method called Linear Interpolation to approximate the value that we need. Here’s how it works, graphically, based on similar triangles:

$$\frac{1.0174 - ???}{1.0174 - 1.0110} = \frac{0.409 - 0.3}{0.409 - 0.271}$$

$$??? = 1.0174 - \frac{0.409 - 0.3}{0.409 - 0.271}(1.0174 - 1.0110)$$

Unknown density = 1.0123 g/cm$^3$