Normality, abbreviated as “N,” is a useful way of measuring the concentration of some solutions in the laboratory. While it is used in many areas of a laboratory technician’s world, we in the water and wastewater fields use it almost exclusively to measure the concentrations of acids and bases for such solutions as titrants in acidity and alkalinity analyses and for the pH adjustments of BOD, ammonia, and phosphorus samples.

Normality is similar in concept to molarity (refer to the previous article “Molarity”). Where molarity (M) represents the concentration of an ion or compound in solution, normality (N) goes one step further and represents the molar concentration only of the acid component (usually the H+ ion in an acid solution) or only the base component (usually the OH- ion in a base solution).

Here is a simple example to show the relationships of Normal acid and base solutions: a 1N solution of the acid H2SO4 will completely neutralize an equal volume of a 1N solution of the base NaOH. Even though the H2SO4 provides two (acid) H+ ions per molecule verses only one (base) OH- ion per NaOH molecule, the calculations of N take into account these differences and puts it all into an equivalent scale. In a sense, with normality calculations, you really can compare apples with oranges – acid and base-wise anyway.

If you know the Molarity of an acid or base solution, you can easily convert it to Normality by multiplying Molarity by the number of hydrogen (or hydroxide) ions in the acid (or base).

N = (M)(number of hydrogen or hydroxide ions)

For example, a 2 M H2SO4 solution will have a Normality of 4N (2 M x 2 hydrogen ions). A 2 M H3PO4, solution will have a Normality of 6N.

However, to make a solution of a predetermined normality requires a bit more calculating. First, you must determine the compound’s equivalent mass. This is done by taking the compound’s gram-molecular mass and dividing by the number of hydrogen ions or hydroxide ions. Here are a few examples:

H2SO4, sulfuric acid.
The gram-molecular mass is 98 (From the periodic chart the individual atomic masses are: H=1, S=32, O=16: {1x2}+32+{16x4}=98).
The number of acid hydrogen ions (H+ ) is 2.
Equivalent mass for H2SO4 is 98/2 = 49.

H3PO4, phosphoric acid. The gram-molecular mass is also 98. The number of hydrogen ions (H+ ) is 3. Equivalent mass for H3PO4 is 98/3 = 32.6.

NaOH, potassium hydroxide. The gram-molecular mass is 40. The number of hydroxide ions (OH-) is 1. Equivalent mass for NaOH is 40/1 = 40.

Once the equivalent mass of an acid or base is determined, you can then calculate the amount of grams needed per volume of water for N.

The formula to calculate this is:

Grams of compound needed = (N desired)(equivalent mass)(volume in liters desired).

For example, how many grams of sodium hydroxide would you need to dilute to a liter to make a 1N NaOH solution?

The equivalent mass is 40 as determined above.

Grams of NaOH needed = (1N)(40 eq. mass)(1 liter) = 40 grams of NaOH.

Similarly, to make 0.25 liters of a 0.05N potassium hydrogen phthalate (KHC8H4O4) solution (an acid), you would first determine the equivalent mass.

From the periodic chart, K=39, H=1, C=12, O=16. Its gram-molecular mass is 39+1+(12x8)+(1x4)+(16x4) = 204.

The number of hydrogen ions it can produce is 1 (The acid hydrogens are usually on the left side of a chemical formula. Hydrogens listed anywhere else usually don’t contribute to the “acid” part of the compound. In the case of KHC8H4O4, only the left most hydrogen is an “acid” hydrogen.) Its equivalent mass is 204/1 = 204.

To find the amount of potassium hydrogen phthalate (KHC8H4O4) needed to make 0.25 liters of a 0.05N solution:

Grams of KHC8H4O4 needed = (0.05N)(204 eq. mass)(0.25 liters) = 2.6 grams of KHC8H4O4.

Both of the chemicals in the examples above, sodium hydroxide and potassium hydrogen phthalate, are considered dry chemicals, which makes it relatively straightforward to calculate their Normalities. For liquid chemicals where the main compound is only a fraction of the total volume, such as the concentrated forms of hydrochloric (HCl), sulfuric (H2SO4), and phosphoric (H3PO4) acids, a few additional calculations must be performed to make a solution of a particular Normality.

The next article will describe and give examples of these additional calculations. These are not only useful for making acid and base solutions, but are useful in calculating concentrations of any type of concentrated dissolved compounds such as alum (aluminum sulfate), bleach (sodium hypochlorite), ferric chloride, and the many other solutions used in wastewater treatment.

Please note that this article specifically covers what is typically found in a wastewater treatment laboratory. There are exceptions to how the concentrations of acids and bases are measured, and this depends on the scope and application of a particular test method.

Making Normal Solutions from Concentrated Acids

The last article covered the concept of Normal solutions in the laboratory and how to calculate the equivalent mass of a compound. Then I described how to use the equivalent mass to make a solution of a predetermined Normality. However, the article did not address making Normal solutions from concentrated mineral acids like sulfuric acid, nitric acid, and hydrochloric acid. Unlike using powdered chemicals where the chemical is simply weighed out then diluted to volume, the use of liquid chemicals to make Normal solutions requires the addition of a few more calculations. This article will address these extra calculations.

First, it is important to describe a few aspects of concentrated mineral acids (as well as that of many other solutions). Most of us buy concentrated acids to use as stock solutions in the laboratory. None of these acids are one hundred percent pure. Sulfuric acid is only about 97% pure, nitric is about 69.5%, and hydrochloric acid is about 37.5% pure. Manufacturers of these acids simply cannot economically make these acids more concentrated than these respective percentages.

Another important aspect of these solutions is their specific gravities. The specific gravity of a liquid is, in most cases, synonymous with the more familiar term of density. Water has a specific gravity of 1. If the specific gravity of a liquid is greater than 1, then the liquid is heavier than water. Less than 1, and the liquid is lighter than water. The specific gravity for concentrated sulfuric acid is about 1.84, or 1.84 times heavier than an equal volume of water. The specific gravity of concentrated nitric acid is about 1.42 and that of concentrated hydrochloric acid is about 1.19.

Both the percent concentration and specific gravity values of the acid are required to determine the amount of concentrated acid needed when making a Normal solution. This information is usually printed on a label attached to the bottle of acid. Specific values vary depending on the manufacturer and lot of acid.

To make a solution of a predetermined Normality, you must first determine the equivalent mass of the chemical and then determine the grams needed of that chemical. These calculations were described in the last article, “Normality.” Then you must convert the number of grams into its volume equivalent. Once this volume is determined, it is a simple dilution after that.

Here is an example:
You want to make only 250 mL of a 1 N H2SO4 solution that will be used to adjust the pH of BOD samples prior to analysis. How many milliliters of concentrated sulfuric acid do you need to make 250 mL of a 1 N solution?

To determine how many grams of sulfuric acid you will need, you will first need to calculate the equivalent mass of H2SO4. This is the gram-formula weight divided by the number of acid hydrogens in the compound. It is 98/2 = 49.

Then you can calculate the amount of grams of H2SO4 that are needed.

The formula to calculate this is:

Grams of compound needed = (N desired)(equivalent mass)(volume in liters desired).

Substituting the above numbers into the equation, we get:
grams of compound needed = (1 N)(49)(0.250 liters) = 12.25 grams.

A 1 N solution requires 12.25 g of a pure sulfuric acid powder (if one existed) diluted to 250 mL. But the acid is a liquid and it is not one hundred percent pure active sulfuric acid. You will need to calculate what volume of the concentrated acid that contains 12.25 grams of sulfuric acid. The formula for this is:

Volume of concentrated acid needed = (grams of acid needed)/(percent concentration x specific gravity)

Continuing with the sulfuric acid example, plug into the formula the percent concentration and specific gravity from the label on the acid container. For this example, I am using those values previously mentioned in this article: volume of concentrated acid needed = (12.25 grams)/(0.97 x 1.84) = 6.9 mL

If you took 6.9 mL of concentrated sulfuric acid and diluted it to 250 mL, you would have a 1 N H2SO4 solution.

(Important note: Always add the acid (or base) to water, in that order. Pour slowly with constant mixing. This will help prevent rapid heat generation and spattering of the mixture. Fill a container about half way or more with distilled water, add the acid, and then bring up to volume with more water. In the example above, fill a flask with about 150 mL or more with distilled water, add 6.9 mL of concentrated sulfuric acid, then continue to dilute with water to the 250 mL mark.)

As with any acid or base made from a concentrated stock solution, the resulting Normality will be an approximate value, which won’t be accurate enough for analytical work. However, it will, in conjunction with a pH meter, be good for adjusting the pH of samples. For analytical procedures where the Normality needs to be accurately known, as in alkalinity titrations, acidity titrations, and volatile acid titrations, you will need to standardize the acid or base. An overview of standardization and the shelf life of acids and bases will be covered in a future article.