

# HOW TO MAKE STANDARD SOLUTIONS FOR CHEMISTRY

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Chemists make two common types of "standard solutions":

- Molar solutions
- Normal solutions

Both of these solutions are concentrations (or "strengths") of a particular component (solute) that is dissolved in a solvent.

Making a Molar solution usually involves fewer mental steps than does making a Normal solution. Below, I will describe both methods.

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## PART 1: MAKING MOLAR (M) SOLUTIONS

A 1 Molar solution (1M) contains 1 mole of solute dissolved in a solution totaling 1 liter. If you use water as the solvent, it must be distilled and deionized. Do not use tap water. A mole is the molecular weight (MW) expressed in grams (sometimes referred to as the 'gram molecular weight' (gMW) of a chemical). Thus, 1 M = 1 gMW of solute per liter of solution.

**Problem: How much sodium chloride is needed to make 1 liter of an aqueous 1 M solution?**

Answer:

First, we calculate the molecular weight (MW) of sodium chloride. Checking the Periodic Table of Elements, we find that the atomic weight of sodium (Na) is 23 and the atomic weight of chlorine (Cl) is 35.5.

Therefore, the molecular weight of sodium chloride (NaCl) is:

$\text{Na (23) + Cl (35.5) = 58.5 grams/mole.}$

To make a 1M aqueous solution of NaCl, dissolve 58.5 grams of NaCl in some distilled deionized water (the exact amount of water is unimportant; just add enough water to the flask so that the NaCl dissolves). Then add more water to the flask until it totals 1 liter. You're done.

Similarly, a 2M solution of sodium chloride contains 117 grams of the salt ( $2 \times 58.5$  grams), topped-off with water to the one liter mark.

Likewise, a 0.1M solution of sodium chloride contains 5.85 grams (58.5 grams/10) of the salt, topped-off with water to the one liter mark.

### MAKING MOLAR SOLUTIONS FROM CONCENTRATED AQUEOUS ACIDS AND BASES

Making a standard molar solution from aqueous acids or bases is a bit more involved than making a standard molar solution from a solid chemical. This is because nearly all liquid acids, no matter how concentrated they are, are already diluted to some extent with water (for instance, pure HCl is a gas, not a liquid, and it is rarely sold in its pure form). Let's use sulfuric acid as our first example.

**Problem: Make one liter of a 1 Molar (1M) aqueous solution of  $\text{H}_2\text{SO}_4$ .**

Answer:

The first step is to read the label on the bottle of the  $\text{H}_2\text{SO}_4$  reagent.

The label will tell you it's molarity. Although there are a variety of concentrations of acids, concentrated  $\text{H}_2\text{SO}_4$  often comes from the factory at a 18.0 Molar concentration (Table 1). This means that there are 18 moles of  $\text{H}_2\text{SO}_4$  in each liter of solution (note: do *not* rely on Table 1; always check the label on the bottle). You need to make a much more diluted solution, so you will add one mole of the concentrated reagent to a fresh batch of water. Your task is to calculate how many milliliters of reagent contain one mole of the acid.

We know from reading the label on the bottle ("18.0 Molar") that one liter of reagent contains 18.0 moles of  $\text{H}_2\text{SO}_4$ . This means that 1 ml of reagent contains 0.018 moles of  $\text{H}_2\text{SO}_4$ . Therefore,

$$\frac{1 \text{ ml}}{0.018 \text{ moles}} = \frac{x \text{ ml}}{1 \text{ mole}}$$

Solving for x, we find that we need 55.6 ml of  $\text{H}_2\text{SO}_4$  reagent.

Therefore, we *slowly* add 55.6 ml of the  $\text{H}_2\text{SO}_4$  reagent to about 500 ml of distilled deionized water, and then we top it off with more water to exactly the "1-liter" mark on the flask.

You have successfully made a 1 Molar  $\text{H}_2\text{SO}_4$  solution. This procedure works similarly with aqueous bases.

**Caution:** *Never* add water into a large volume of concentrated acid! You risk creating an explosion! The rule is: “Acid into water = you’re doing what ya oughta.”  
“Water into acid = you might get blasted!”

Therefore, *always* add a smaller volume of acid into a larger volume of water.

Table 1

TYPICAL CONCENTRATIONS OF CONCENTRATED ACIDS AND BASES  
(as written on the labels of their containers)

ACID/BASE NAME	WT%	DENSITY (sp. gr) (g/ml)	MOLARITY
Acetic acid	99.7%	1.05 g/ml	17.4
Ammonium hydroxide (aqueous ammonia)	28%	0.89 g/ml	14.6
Hydrochloric acid	37%	1.18 g/ml	12.0
Nitric acid (HNO <sub>3</sub> )	70%	1.40 g/ml	15.6
Phosphoric acid	85%	1.69 g/ml	14.7
Sulfuric acid	96%	1.84 g/ml	18.0

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**PART 2: MAKING NORMAL SOLUTIONS (N)**

Compared to making Molar solutions, making Normal solutions can be a bit confusing. Aqueous solutions of acids and bases are often described in terms of their normality rather than their molarity. In order to properly make a Normal solution, the worker must understand the difference between a *pure* reagent and a *diluted* reagent.

A "1 Normal" solution (1 N) contains 1 “gram equivalent weight” (gEW) of solute, topped-off to one liter of solution. The gram equivalent weight is equal to the solute’s molecular weight, expressed as grams, divided by the valence (n) of the solute:

$$\text{Equivalent weight (EW)} = \frac{\text{molecular weight}}{n}$$

After the equivalent weight (or milliequivalent weight) has been calculated, then the following equation is used:

$$N = \frac{\text{Weight of solute}}{\text{milliequivalent weight of solute} \times \text{Volume (in ml) of dilution}}$$

The equivalent weight (or milliequivalent weight) of a substance depends upon the type of reaction in which the substance is taking part. Some different types of chemical reactions, along with how to determine a solute's equivalent weight for each reaction, are given below.

### MAKING A NORMAL SOLUTION WITH SALTS

**Problem: Calculate the normality of a sodium chloride solution prepared by dissolving 2.9216 grams of NaCl in water and then topping it off with more water to a total volume of 500.0 ml.**

Answer:

- Checking the Periodic Table of Elements, we find that the molecular weight of NaCl is 58.44
- $n = 1$  (because there is room in the molecule for only one replaceable  $H^+$  ion). In other words, one hydrogen atom can replace the sodium atom in NaCl.
- Therefore, the equivalent weight of NaCl is:  $\frac{58.44}{1}$  or 58.44

- Therefore, the *milliequivalent* weight of NaCl is:  $\frac{58.44}{1000}$  or 0.05844

- The normality (N) is:

$$N = \frac{2.9216 \text{ grams}}{[0.05844 \times 500 \text{ ml}]}$$

$$N = 0.099$$

### MAKING NORMAL SOLUTIONS WITH PURE (NON-AQUEOUS) ACIDS

The equivalent weight of an acid is its molecular weight, divided by the number of replaceable hydrogen atoms in the reaction.

To clarify this concept, we will consider the following acids:

Hydrochloric acid (HCl) has one replaceable hydrogen ion ( $H^+$ ). Sulfuric acid ( $H_2SO_4$ ) has two replaceable hydrogen ions ( $2H^+$ ). The valences of these acids are determined by their respective replaceable hydrogen ions (Table 2):

Table 2

HCl	n = 1
HNO <sub>3</sub>	n = 1
H <sub>2</sub> SO <sub>4</sub>	n = 2
HF	n = 1

So, for pure HCl, its MW is 36.46, its EW is 36.46 and therefore a 1N solution would be 36.46 grams of the pure chemical per liter. Note that, in the case of HCl, a 1N solution has the same concentration as a 1M solution..

To make a 1N  $H_2SO_4$  solution from pure  $H_2SO_4$ , its MW is 98.08, and its EW is:

$$\frac{98.08}{2}$$

Therefore,

$$\text{EW} = 49.04 \text{ grams per liter (98.08/2)}$$

(or 49.04 grams per 1000 milliliters)

So, a 1 N solution would be 49.04 grams of the pure chemical per liter.

To make Normal acid solutions from *aqueous* reagents, see p. 7.

### **MAKING NORMAL SOLUTIONS FROM PURE ALKALIS (BASES)**

The equivalent weight of a base is defined as "Its molecular weight divided by the number of hydrogen ions that are required to neutralize the base".

To understand the valences of alkalis, consider the following examples::

The  $(OH)^-$  ion in Sodium Hydroxide (NaOH) can be neutralized by one hydrogen ion. The  $(OH)_2^{2-}$  ions in Calcium hydroxide ( $Ca(OH)_2$ ) can be neutralized by two hydrogen ions. As was the case with acids, the valences (n) of these bases are determined by their respective replaceable hydrogen ions (Table 3):

Table 3

NaOH	n = 1
Ca(OH) <sub>2</sub>	n = 2

So, for NaOH, its MW is 40, its EW is 40, and therefore a 1N solution would be 40 grams of the pure chemical per liter of water. You will also note that, in the case of NaOH, a 1N solution is the same concentration as a 1M solution.

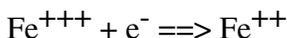
For Ca(OH)<sub>2</sub>, its MW is 74, its EW is  $74/2 = 37$  (because  $n = 2$ ). Therefore, a 1 N aqueous solution of Ca(OH)<sub>2</sub> is 37 grams of the pure chemical per liter of water. Of course, many acidic reagents and basic reagents come from the factory in a diluted aqueous form. To make Normal basic solutions from *aqueous* reagents, see p. 7.

### HOW TO DETERMINE THE EQUIVALENT WEIGHT OF A SOLUTE IN REDOX REACTIONS

The equivalent weight of a substance undergoing oxidation or reduction is equal to its molecular weight divided by the total number of electrons gained or lost per molecule in that particular reaction. All you have to do is examine the redox equation and count the number of electrons involved in the reaction.

$$\text{equivalent weight} = \frac{\text{molecular weight of solute}}{\text{number of } e^- \text{ gained or lost}}$$

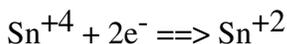
Example #1:



$$\text{Equivalent weight of Fe} = \frac{58.85}{1}$$

Therefore, EW = 58.85

Example #2:

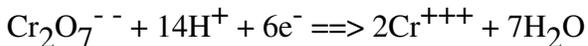


$$\text{Equivalent weight Sn} = \frac{118.69}{2}$$

Therefore, EW = 59.34

Example #3:

Potassium Dichromate ( $K_2Cr_2O_7$ )



$$\text{Equivalent weight } K_2Cr_2O_7 = \frac{294.19}{6}$$

Therefore, EW = 49.03

### **PREPARATION OF NORMAL SOLUTIONS FROM CONCENTRATED AQUEOUS REAGENTS**

To prepare a standard Normal solution from a concentrated aqueous laboratory reagent, knowledge of the specific gravity (in grams/ml) and the percentage composition (by weight) of the reagent are required. The following example will produce a solution of a known normality. Subsequent titration would be necessary to achieve a highly precise normality, however a discussion of titration techniques is beyond the scope of this document. Consult a chemistry text book on quantitative analysis for more details on how to do a titration.

**Problem: Calculate the volume of concentrated aqueous sulfuric acid, having a specific gravity of 1.842 and containing 96.0%  $H_2SO_4$  (by weight), required to prepare 2.0 liters of 0.20 N  $H_2SO_4$ .**

Answer:

Consulting the Periodic Table of Elements, we find that the molecular weight of  $H_2SO_4$  is 98.08.

$$\text{Equivalent weight} = \frac{\text{molecular weight}}{n}$$

Since  $n = 2$ , the equivalent weight is:

$$\frac{98.08}{2} = 49.04 \text{ g/eq}$$

The weight of the H<sub>2</sub>SO<sub>4</sub> reagent that is required is calculated as follows:

Equation #1:

$$\text{Weight H}_2\text{SO}_4 \text{ needed} = \frac{0.20 \text{ eq}}{\text{liter}} \times 2.0 \text{ liters} \times \frac{49.04 \text{ g}}{\text{eq}}$$

Multiplying and canceling-out some units, we get:

$$\text{Weight H}_2\text{SO}_4 \text{ needed} = 19.616 \text{ grams}$$

However, we want to find the *volume* of concentrated H<sub>2</sub>SO<sub>4</sub> that is needed, not its weight. The weight of H<sub>2</sub>SO<sub>4</sub> is related to its current diluted volume by the following equation:

Equation #2:

$$\text{Weight} = \frac{V(\text{ml}) \times 1.842 \text{ g} \times 0.96}{\text{ml}}$$

The “0.96” refers to the weight percent of H<sub>2</sub>SO<sub>4</sub> in the reagent, which is written on the bottle’s label. In this case, it is 96%. Simplifying the equation, we arrive at:

$$\text{Weight} = \frac{V(\text{ml}) \times 1.76832 \text{ grams}}{\text{ml}}$$

Since “weight” is common to both equations #1 and #2, we can combine them into one equation:

Equation #3:

$$19.616 \text{ grams} = \frac{V(\text{ml}) \times 1.76832 \text{ grams}}{\text{ml}}$$

Solving for V(ml) we get:

$$V(\text{ml}) = \frac{19.616 \text{ grams}}{1.76832 \text{ grams}} \times \text{ml}$$

Canceling-out some units and rounding off, we get:

$$V(\text{ml}) = 11.1 \text{ ml}$$

We *slowly* pour 11.1 ml of  $\text{H}_2\text{SO}_4$  into 1 liter of distilled deionized water, and then top it off with more water until the total volume of the solution is 2 liters. We have successfully created a 0.2 N solution.

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**Problem: Calculate the volume of concentrated aqueous HCl, having a density of 1.188 g/ml and containing 38% HCl by weight, needed to prepare 2 liters of a 0.20 N hydrochloric acid solution.**

Answer:

The molecular weight of HCl = 36.461

The valence (n) of hydrogen = 1

Equation #1

$$\text{Weight of HCl needed} = 0.20 \text{ eq/Liter} \times 2.0 \text{ Liters} \times 36.461 \text{ g/eq}$$

Re-writing the equation:

$$\text{Weight of HCl needed} = \frac{0.20 \text{ eq}}{1000 \text{ ml}} \times 2000 \text{ ml} \times \frac{36.461 \text{ grams}}{\text{eq}}$$

$$\text{Weight} = 14.5844 \text{ grams}$$

The weight of HCl is related to its current diluted volume by the following equation:

Equation #2

$$\text{Weight} = \frac{V(\text{ml}) \times 1.188 \text{ grams} \times 0.38}{\text{ml}}$$

Combining equations #1 and #2 gives us equation #3:

Equation #3

$$14.5844 \text{ grams} = \frac{V(\text{ml}) \times 1.188 \text{ grams} \times 0.38}{\text{ml}}$$

Solving for V(ml):

$$V(\text{ml}) = \frac{14.5844 \text{ grams} \times \text{ml}}{1.188 \text{ grams} \times 0.38}$$

Canceling-out some units and rounding off, we arrive at our answer:

$$V(\text{ml}) = 32.3 \text{ ml}$$

Therefore, we *slowly* add 32.3 milliliters of concentrated HCl to, say, 500 milliliters of distilled deionized water (remember: always add a *smaller* volume of acid to a larger volume of water), and then we top it off with more water until we have a total volume of 2 liters.

### SOME HELPFUL TIPS

- Always use a fume hood when handling highly concentrated acids and bases. Wear a plastic apron, plastic gloves, and eye goggles. This is particularly important when working with hydrofluoric acid.
- Highly basic (alkali) reagents can, over time, slowly leach silica from the glass container into the reagent. For some applications in chemistry, a trace amount of silica contamination isn't a problem, but if you need to make highly precise chemical measurements, the dissolved silica can affect the reagent's chemical properties. If you need a reagent that can produce highly precise results in quantitative analysis, then you should store all newly prepared strong basic solutions in labeled plastic containers.
- Highly acidic reagents can be stored in either glass containers or in plastic containers. The exception is hydrofluoric acid, which etches glass and weakens it. Store hydrofluoric acid in a labeled plastic container.
- It is advisable to store concentrated reagents that come straight from the supplier in their original containers, with their labels intact.
- When preparing Molar solutions and Normal solutions, students sometimes make the mistake of adding the solute to a set volume of solvent. Doing this will produce a solution of the wrong concentration. Example: When making one liter of a 1 M solution of NaCl, do *not* add 58.5 grams of NaCl to 1 liter of water. Instead, the correct way to make the solution is to add 58.5 grams of NaCl into a container and then top it off with water to a total volume of 1 liter.