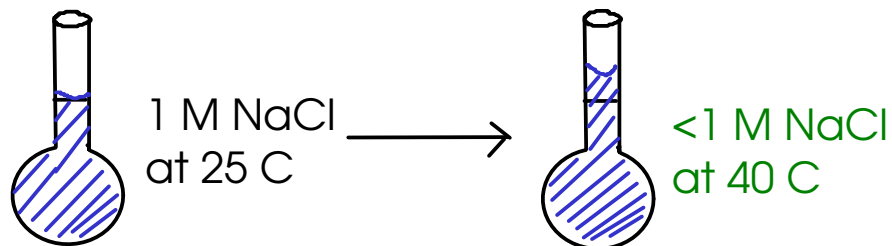


## <sup>60</sup> MOLARITY

- In the previous example, we converted between three of the four units that we discussed: mass percent, molality, and mole fraction.
- We didn't do MOLARITY, because the information given in the previous problem was not sufficient to determine molarity!

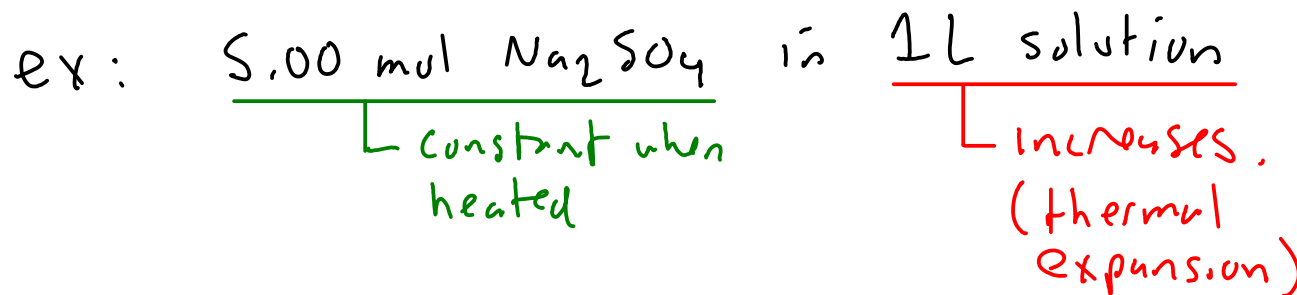
$$\underline{M} = \frac{\text{moles solute}}{\underline{\text{L solution}}}$$



Molarity is based on VOLUME, while the other three units are based on MASS. (moles and mass can be directly converted)

Volume depends on TEMPERATURE!

- If you HEAT a solution, what happens to CONCENTRATION?



... the MOLAR CONCENTRATION decreases. (But the concentration in the other three units we discussed stays the same.)

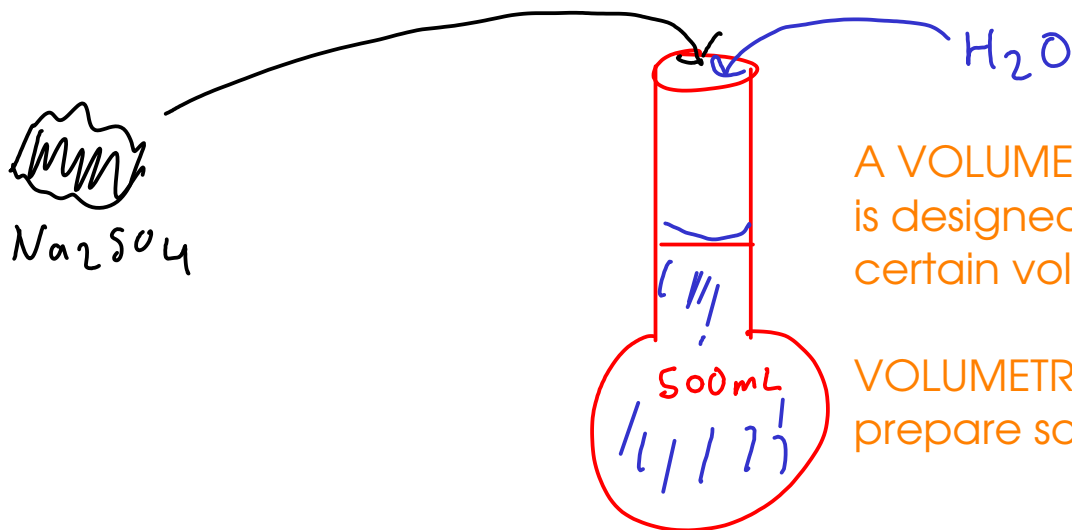
- If you COOL a solution, the MOLAR CONCENTRATION increases. (The other three units stay the same!)

<sup>61</sup> ... we use MOLARITY so much because it's easy to work with. It is easier to measure the VOLUME of a liquid solution than it is to measure mass.



Example: How would we prepare 500. mL of 0.500 M sodium sulfate in water?

Dissolve the appropriate amount of sodium sulfate into enough water to make 500. mL of solution.



A VOLUMETRIC FLASK is a flask that is designed to precisely contain a certain volume of liquid.

VOLUMETRIC FLASKS are used to prepare solutions.

volumetric flask

$$M = \frac{\text{mol Na}_2\text{SO}_4}{\text{L solution}} ; 0.500 \text{ M} = \frac{\text{mol Na}_2\text{SO}_4}{0.500 \text{ L}}$$

$$\text{mol Na}_2\text{SO}_4 = (0.500 \text{ M})(0.500 \text{ L}) = 0.250 \text{ mol Na}_2\text{SO}_4$$

$$0.250 \text{ mol Na}_2\text{SO}_4 \times \frac{142.05 \text{ g Na}_2\text{SO}_4}{\text{mol Na}_2\text{SO}_4} = \boxed{35.5 \text{ g Na}_2\text{SO}_4}$$

Dissolve 35.5 g sodium sulfate in enough water to make 500. mL of solution!

To prepare a solution of a given molarity, you generally have two options:

① Weigh out the appropriate amount of solute, then dilute to the desired volume with solvent (usually water)"

② Take a previously prepared solution of known concentration and DILUTE it with solvent to form a new solution

"stock solution"

- Use DILUTION EQUATION

The dilution equation is easy to derive with simple algebra.

$$M \times V$$

$$\frac{\text{mol}}{\text{L}} \times \text{L} = \text{moles solute}$$

... but when you dilute a solution, the number of moles of solute REMAINS CONSTANT. (After all, you're adding only SOLVENT)

$$M_1 V_1 = M_2 V_2$$

before  
dilution

after  
dilution

Since the number of moles of solute stays the same, this equality must be true!

$$M_1 V_1 = M_2 V_2 \quad \dots \text{the "DILUTION EQUATION"}$$

$M_1$  = molarity of concentrated solution

$V_1$  = volume of concentrated solution

$M_2$  = molarity of dilute solution

$V_2$  = volume of dilute solution  $\leftarrow$  (TOTAL VOLUME, NOT the volume water added!)

The volumes don't HAVE to be in liters, as long as you use the same volume UNIT for both  $V_1$  and  $V_2$

Example: Take the 0.500 M sodium sulfate we discussed in the previous example and dilute it to make 150. mL of 0.333 M solution. How many mL of the original solution will we need to dilute?

$$M_1 V_1 = M_2 V_2$$

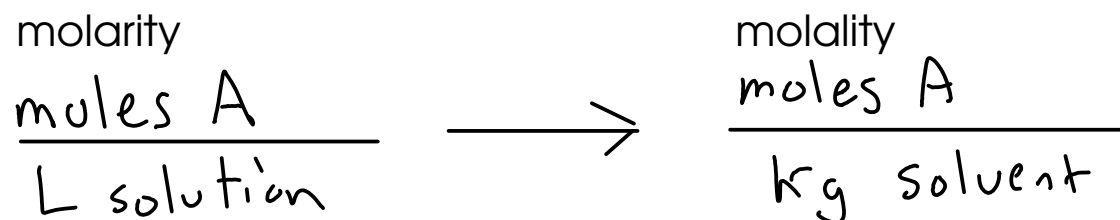
$$(0.500 \text{ M})(V_1) = (0.333 \text{ M})(150. \text{ mL})$$

$$V_1 = 99.9 \text{ mL of } 0.500 \text{ M Na}_2\text{SO}_4$$

Take 99.9 mL of 0.500 M sodium sulfate, and add enough water to make 150. mL of solution. (The TOTAL VOLUME of solution will be 150 mL, so we end up adding about 50.1 mL of water!)

## <sup>64</sup> MOLARITY and the other concentration units

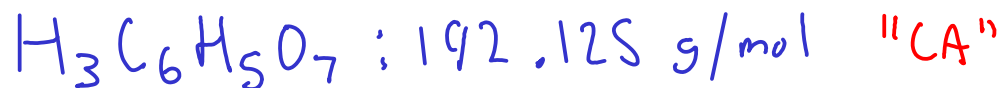
- To convert between molarity and the other three concentration units we've studied, you have to know more about the solution. For example:



- \* To perform this conversion, you can assume a liter of solution, which will give you the number of moles present. But you've then got to have a way to convert the volume of SOLUTION to the mass of the SOLVENT. How?
- \* You need DENSITY (which depends on temperature). The density of the solution will allow you to find the total mass of the solution.
- \* If you subtract out the mass of the SOLUTE, then what you have left is the mass of the SOLVENT. Express that in kilograms, and you have all the information you need to find molality!
- \* You'll run into the same situation when you use any of the other mass or mole based units. DENSITY is required to go back and forth between MOLARITY and these units.

<sup>65</sup> Example: If a solution is 0.688 m citric acid, what is the molar concentration (M) of the solution?

The density of the solution is 1.049 g/mL



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$$\frac{0.688 \text{ mol CA}}{\text{kg solvent}} \longrightarrow \frac{? \text{ mol CA}}{? \text{ L solution}}$$

molality (definition) molarity (definition)

- 1 - ASSUME A BASIS of 1 kg of solvent. Each kilogram of solvent contains 0.688 mol CA.
  - 2 - Find VOLUME of SOLUTION. We know DENSITY of the solution, but before we can use it, we must find out how much the SOLUTION weighs. So, we need to figure out how much the CA in the solution weighs, and add it to the solvent mass to get mass of solution.
- 

$$0.688 \text{ mol CA} \times \frac{192.125 \text{ g CA}}{\text{mol CA}} = 132.182 \text{ g CA}$$

$$\text{mass solution} = 1000 \text{ g solvent} + 132.182 \text{ g CA} = 1132.182 \text{ g solution}$$

Find volume solution:

$$1132.182 \text{ g solution} \times \frac{\text{mL}}{1.049 \text{ g}} \times \frac{10^{-3} \text{ L}}{\text{mL}} = 1.079296473 \text{ L solution}$$

$$M = \frac{\text{mol CA}}{\text{L solution}} = \frac{0.688 \text{ mol CA}}{1.079296473 \text{ L}} = \boxed{0.637 \text{ M CA}}$$

<sup>66</sup> An aqueous solution is 8.50% ammonium chloride by mass. The density of the solution is 1.024 g/mL  
Find: molality, mole fraction, molarity.

---



Find molality:

$$\frac{8.50 \text{ g NH}_4\text{Cl}}{100 \text{ g solution}} \longrightarrow \frac{\text{mol NH}_4\text{Cl}}{\text{kg H}_2\text{O}}$$

mass percent molality

Find mol ammonium chloride:

$$8.50 \text{ g NH}_4\text{Cl} \times \frac{\text{mol NH}_4\text{Cl}}{53.491 \text{ g NH}_4\text{Cl}} = 0.1589052364 \text{ mol NH}_4\text{Cl}$$

Find mass water:

$$100 \text{ g solution} - 8.50 \text{ g NH}_4\text{Cl} = 91.50 \text{ g H}_2\text{O} = 0.09150 \text{ kg H}_2\text{O}$$

Molality:

$$\frac{0.1589052364 \text{ mol NH}_4\text{Cl}}{0.09150 \text{ kg H}_2\text{O}} = \boxed{1.74 \text{ m NH}_4\text{Cl}}$$

Mole fraction:

$$\frac{8.50 \text{ g NH}_4\text{Cl}}{100 \text{ g solution}} \longrightarrow \frac{\text{mol NH}_4\text{Cl}}{\text{mol NH}_4\text{Cl} + \text{mol H}_2\text{O}}$$

mass percent mole fraction

We already know the moles ammonium chloride (see the molality calculation), so we just need to get the moles water. Since we already found MASS water, start there!

An aqueous solution is 8.50% ammonium chloride by mass. The density of the solution is 1.024 g/mL

Find: molality, mole fraction, molarity.



Find moles water:

$$91.50 \text{ g H}_2\text{O} \times \frac{\text{mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 5.078818828 \text{ mol H}_2\text{O}$$

$$X_{\text{NH}_4\text{Cl}} = \frac{0.1589052364 \text{ mol NH}_4\text{Cl}}{0.1589052364 \text{ mol NH}_4\text{Cl} + 5.078818828 \text{ mol H}_2\text{O}}$$

$$= \boxed{0.0303} \quad \text{If we need } X_{\text{water}}, \text{ just subtract } 1 - X_{\text{ammonium chloride}}$$

Molarity:

$$\frac{8.50 \text{ g NH}_4\text{Cl}}{100 \text{ g solution}} \longrightarrow \frac{\text{mol NH}_4\text{Cl}}{\text{L solution}}$$

mass percent molarity

We already know moles ammonium chloride from previous work, so all we need to find is the volume of SOLUTION

$$100 \text{ g solution} \times \frac{\text{mL}}{1.024 \text{ g}} \times \frac{10^{-3} \text{ L}}{\text{mL}} = 0.09765625 \text{ L solution}$$

$$\frac{0.1589052364 \text{ mol NH}_4\text{Cl}}{0.09765625 \text{ L solution}} = \boxed{1.63 \text{ M NH}_4\text{Cl}}$$