Molar concentration is a measure of the concentration of a solution. It is defined as moles of solute per litre of solution, and has units of mol dm^{-3}.

Here, *N* is the number of molecules present in the volume. This latter is measured in litres. *A* is the Avogadro's number, 6.022×10^{23}

Molar concentration is also often represented by a capital m (M)

Another formula that may be used is M=n/v, where M is the molar concentration, n is the number of moles and v is the volume in litres.

Examples

Most proteins are present in the bacteria E.Coli at 60 copies or fewer. The volume of the bacteria is 10-15 litre, which gives us *C*=10-7 M =100nM. (nM is nanomolar *i.e.* 10-9).

2 grams of NaCl dissolved in 5mL of Water. As 58 grams of NaCl is 1 mol of molecules, and 1mL is 0.001 litre, this gives *C*= (2/58)/0.005=6.9 M.

Molarity

See Main article molar solution

Molarity (M or molar) denotes the number of moles of a given substance per litre of solution. For instance:

$$\frac{2.0 \quad moles \quad of \quad dissolved \quad particles}{4.0 \quad litres \quad of \quad liquid} = solution \quad of \quad 0.5M$$

Such a solution may be described as "0.5 molar." This can be misinterpreted as "0.5 moles of solute per 1.0 litre of solvent", which can be true only if no volume change occurs on adding the solute to the solvent. This is why preparation of a solution of known molarity involves adding an accurately weighed amount of solute to a volumetric flask, adding some solvent to dissolve it, then adding more solvent to fill to the volume mark. (Working with moles can be highly advantageous, as they enable measurement of the absolute number of particles in a solution, irrespective of their weight and volume. This is often more useful when performing stoichiometric calculations.).

When discussing the molarity of minute concentrations, such as in much pharmalogical research, molarity is expressed in millimolars (mM, 1 thousandth of a molar), micromolars (µM, 1 millionth of a molar) or nanomolars (nM, 1 billionth of a molar).

Although molarity is by far the most commonly used measure for concentration, particularly for dilute aqueous solutions, it does suffer from a number of disadvantages. Masses can be determined with great precision as balances are often very precise.

Determining volume is often not as precise. In addition, a volume of a liquid changes with temperature so that the molarity also changes without adding or removing any mass. For non-dilute solutions another problem is that the molar volume of a substance is itself a function of concentration so that volume is not strictly additive.

Following the SI system of units, the National Institute of Standards and Technology, the United States authority on measurement, considers the term molarity and the unit symbol M to be obsolete, and suggests instead the 'amount-of-substance concentration' (*c*) with units mol/m^3 or other units used alongside the SI such as mol/L.[1] This recommendation has not been generally implemented in academia yet.

Molality

Molality (m or molal) denotes the number of moles of a given substance per kilogram of *solvent* (not *solution*). Note that molality is represented by the symbol (*m*), while molarity is represented by the symbol (M). The two symbols are not meant to be confused and should not be used as symbols for units. The SI unit for molality is mol/kg. For instance: adding 1.0 moles of dissolved particles to 2.0 kilograms of solvent constitutes a solution with a molality of 0.5 mol/kg. Such a solution may be described as "0.5 molal". The term **molal solution** is used as shorthand for a "one molal solution", i.e. a solution which contains one mole of the solute per 1000 grams of the solvent.

The determination of molality *only* requires a good balance, because the masses of both solvent and solute can be obtained by *weighing*. Using a balance is often more precise than working with volumetric flasks burettes and pipettes. Another advantage of molality is that it does not change with the temperature as it deals with the mass of solvent, rather than the volume of solution. Volume typically increases with increase in temperature resulting in decrease in molarity. Molality of a solution is always constant irrespective of the physical conditions like temperature and pressure.

In a dilute aqueous solution near room temperature and standard atmospheric pressure, the molarity and molality will be very similar in value. This is because 1 kg of water roughly corresponds to a volume of 1 L at these conditions, and because the solution is dilute, the addition of the solute makes a negligible impact on the volume of the solution.

However, in all other conditions, this is usually not the case.

Mole fraction

The *mole fraction* X, (also called *molar fraction*) denotes the number of moles of solute as a proportion of the total number of moles in a solution. For instance: 1 mole of solute dissolved in 9 moles of solvent has a mole fraction of 1/10 or 0.1. Mole fractions are *dimensionless* quantities. (The *mole percentage* or *molar percentage*, denoted "mol %" and equal to 100% times the mole fraction, is sometimes quoted instead of the mole fraction.)

This measure is used very frequently in the construction of phase diagrams. It has a number of advantages:

- the measure is not temperature dependent (such as molarity) and does not require knowledge of the densities of the phase(s) involved
- a mixture of known mole fraction can be prepared by weighing off the appropriate masses of the constituents
- the measure is *symmetrical*: in the mole fractions X=0.1 and X=0.9, the roles of 'solvent' and 'solute' are reversed.

As both mole fractions and molality are only based on the masses of the components it is easy to convert between these measures. This is not true for molarity, which requires knowledge of the density.

Mass percentage

Mass percentage denotes the mass of a substance in a mixture as a percentage of the mass of the entire mixture. For instance: if a bottle contains 40 grams of ethanol and 60 grams of water, then it contains 40% ethanol by mass. Commercial concentrated aqueous reagents such as acids and bases are often labeled in concentrations of *weight percentage* with the specific gravity also listed. In older texts and references this is sometimes referred to as *weight-weight percentage* (abbreviated as *w/w*). In water pollution chemistry, a common term of measuring total mass percentage of dissolved solids in an aqueous medium is total dissolved solids.

Mass-volume percentage

Mass-volume percentage, (sometimes referred to as weight-volume percentage or percent weight per volume and often abbreviated as % m/v or % w/v) describes the mass of the solute in g per 100 mL of the resulting solution. Mass-volume percentage is often used for solutions made from a solid solute dissolved in a liquid. For example, a 40% w/v sugar solution contains 40 g of sugar per 100 mL of resulting solution.

Mass-volume ratio

Often used in medicine and pharmacology, a ratio of the weight of a drug dissolved in a volume of water, is presented as, grams of solute: mL of water. Practitioners use the term "dilution" when referring to this arcane unit. The most ubiquitous example is epinephrine solutions where a 1:100,000 solution has 1 g epinephrine in 100,000 mL water. This is equivalent to 0.01 g/L epinephrine solution.

When the volume considered is a gas a specific approach is needed: the gas' pressure and temperature conditions must be considered. A typical use is in air pollution emission quantification. It is very common to find values such as 50 g/Nm^3 or 50 g/m^3N. The "N" before or after the cubic meter indicates that the gas is under the Standard conditions for temperature and pressure.

How do you make these? The first thing you need to know is the mass of a mole of NaOH. This is just its molecular weight: 40.0 g/mol. (That's what a molecular weight means: the mass in grams of one mole of the substance.)
So, to make a 1.0 M (= 1.0 N) solution of NaOH in water, you will want to weigh out 40.0 grams of NaOH, dissolve it in about 0.8 liters of water, and then add water to the solution to take the total volume up to exactly 1.0 liters. You would do the same thing to make a 0.1 M (= 0.1 N) solution: weigh out 0.1 mole of NaOH (= 4.0 g), dissolve it in water, and add enough water to make the total volume equal to exactly 1 liter.

1. Example:
   What is the molality of a solution of 10 g NaOH in 500 g water?

   Solution:
   10 g NaOH / (4 g NaOH / 1 mol NaOH) = 0.25 mol NaOH
   500 g water x 1 kg / 1000 g = 0.50 kg water
   molality = 0.25 mol / 0.50 kg
   molality = 0.05 M / kg
   molality = 0.50 m