UNITS OF CONCENTRATION

There are a number of different ways of expressing solute concentration that are commonly used. Some of these are listed below.

Molarity, M = moles solute/liter of solution

Normality, N = equivalents of solute/liter of solution

Weight %, Wt % = (mass of solute/mass of solution) x 100%

Parts per million, ppm = (mass of solute/mass of solution) x 10^6

Mass per volume, mg/L = mass of solute/liter of solution

molality, *m* = moles of solute/mass of solvent

mole fraction, χ = moles of solute/total moles

Concentrations expressed as ppm and N are less familiar to most students at this stage.

Parts per million:
The number of milligrams of solute per kg of solution = one ppm,
since 1 mg = 10^{-3} g and 1 kg = 10^3 g.

Assuming the density of water is 1.00 g/mL, 1 liter of solution = 1 kg and hence,
1 mg/L = 1 ppm. This is generally true for freshwater and other dilute aqueous solutions.

Parts per million concentrations are essentially mass ratios (solute to solution) x a million (10^6). In this sense, they are similar to wt %, which could be thought of as parts per hundred (although nobody uses this term).

Other variations on this theme include:
ppt – parts per thousand (used for common ions in sea water)
ppb – parts per billion (used for heavy metals and organics)
ppt – parts per trillion (used for trace metals and trace organics)

The following table summarizes common mass ratios for solutions and solids.

Unit	Solutions		Solids	
ppm	mg/L	μg/mL	mg/kg	μg/g
ppb	μg/L	ng/mL	μg/kg	ng/g
ppt	ng/L	pg/mL	ng/kg	pg/g

To convert concentrations in mg/L (or ppm in dilute solution) to molarity, divide by the molar mass of the analyte to convert mass in mg into a corresponding number of moles.

What is the molarity of a 6.2 mg/L solution of O_2(aq)?

To convert from molarity to mg/L (or ppm in dilute solution), multiply by the molar mass of the analyte to convert moles into corresponding number of moles.

The Maximum Acceptable Concentration (MAC) of Pb in drinking water is 10 ppb. If a sample has concentration of 55 nM, does it exceed the MAC?

Note 1: In seawater, 1 mg/L ≠ 1 ppm since the density of seawater is 1.03 g/mL.

Hence, 1.00 mg/$L_{sewater}$ = 1.00 mg/L x 1 mL/1.03 g x 1 L/1000 mL x 1000 mg/g
= 0.971 mg/kg or 0.971 ppm

Note 2: Some concentrations are expressed in terms the species actually measured
e.g., mg/L of NO_3^- (mass of nitrate ions per liter)

Or in terms of a particular element in a species that was measured.
e.g., mg/L of NO_3^- - N (mass of nitrogen in the form nitrate ions per liter)

To convert from one to the other of these, use the molar mass ratio of the element to that of the chemical species measured. In the example above use; 14 mg N/62 mg NO_3^-.

It is important to clearly report unit values to avoid serious error in interpretation of results. Similar situations arise in reporting the concentrations of ammonia-nitrogen, phosphates-phosphorous and others.

Note 3: Some aggregate parameters are reported in terms of a single surrogate species.

e.g., total hardness is usually reported as the mass of $CaCO_3$ that would be required to provide the same number of moles of calcium ions.

Normality is a somewhat dated concentration unit that is still encountered in may texts and lab manuals. It has advantages when carrying out titration calculations, however it can be confusing for the uninitiated. Normality is defined as the number of equivalents of solute per liter, and as such, is similar to Molarity.

equivalents of solute = mass of solute/equiv. weight

and the Equiv. Weight = M.W./K

where K = #equivalents per mole, K is an integer constant ≥ 1

Hence, N = K x M

K for a particular species is defined by the reaction type and the balanced chemical reaction.

For acid/base rxn's: K is the number of moles of H+ ions produced or neutralized per mole of acid or base supplied. Thus,

Acid/base	**K**	**M.W.**	**E.W.**
HCl	1	36.5	36.5
H_2SO_4	2	98.1	49.0
$CaCO_3$	2	100	50.0
$Al(OH)_3$	3	78.0	26.0

For oxidation/reduction reactions, K is the number of moles of e- transferred per mole of oxidant or reductant in the balanced half-reaction.

Balanced half reaction	**K**
$Fe^{3+} \rightarrow Fe$	3
$I_2 \rightarrow 2\ I^-$	2
$2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-}$	1

Using Normality in titration calculations.
Method 1: use the appropriate value of K to convert Molarity,
(i.e., 0.250 N H_2SO_4 = 0.125 M H_2SO_4) and use the coefficients in the balanced chemical equations to solve for the number of moles of analyte in given sample volume.

Method 2: use the normal concentrations directly ignoring the coefficients in the balnaced chemical equation.

# of equiv. of analyte = # equiv. titrant

The number of equiv. of analyte in a given volume of sample can now be converted to moles/L or mg/L using K or E.W., respectively.

EXAMPLES

1. When 25.0 mL of NaOH solution was titrated, 23.4 mL of 0.286 M H_2SO_4 were required to reach the end point. Find the molarity of the NaOH.

2. When 25.0 mL of NaOH solution was titrated, 23.4 mL of 0.572 N H_2SO_4 were required to reach the end point. Find the normality of the NaOH.

3. The Winkler titration for the determination of dissolved oxygen involves the treatment of the sample with iodide ion (I^-) in the presence of manganese ion catalyst (Mn^{2+}) as follows.

$$O_2 \quad + \quad 2\ Mn^{2+} \quad + \quad 4\ OH^- \quad \rightarrow \quad 2\ MnO_2(s) \quad + \quad 2\ H_2O$$

$$MnO_2(s) \quad + \quad 4\ H^+ \quad + \quad 2\ I^- \quad \rightarrow \quad Mn^{2+} \quad + \quad I_2 \quad + \quad 2\ H_2O$$

The liberated iodine is then titrated with a standard thiosulfate solution.

$$I_2 \quad + \quad 2\ S_2O_3^{2-} \quad \rightarrow \quad 2\ I- \quad + \quad S_4O_6^{2-}$$

A 50.00 mL sample of water was treated as above and the I_2 titrated with 0.01136 N $Na_2S_2O_3$, of which 8.11 mL were required to reach the end point. Determine the concentration of the dissolved O_2 in equiv./L, moles/L and mg/L.