Calculating Concentrations

What is a concentration: The amount of solute in a solvent

There are many ways to calculate concentrations. The most common are:

Name	Symbol	Units	Areas of Application
Mass Percent	**%m/m**	**%**	**Solid – liquid phase solutions (commercial products)**
Parts per million	**ppm**	$\frac{g\ solute}{1000\,000\ g\ solution}$	**Express small concentrations**
Molality	***m***	$\frac{mole\ solute}{kg\ solvent}$	**With calculations of properties such as boiling point elevation and freezing point depression**
Molarity	**M**	$\frac{mole\ solute}{L\ solution}$	**Solution Stoichiometry calculations**

Mass Percent:

$$Mass\ \% = \%\left(\frac{m}{m}\right) = \frac{mass\ of\ solute}{mass\ of\ solution}\ x\ 100$$

NOTE: solution = solute + solvent. You must add these in the denominator.

Example: Calculate the mass % for 15.0 grams of potassium nitrate being dissolved in 135 grams of water.

$$Mass\ \% = \frac{15.0\ g}{(15.0\ g + 135\ g)} x100$$

$$= 10.0\%\ potassium\ nitrate\ solution$$

Parts per Million (PPM):

$$ppm = \frac{(mass\ of\ solute)}{(mass\ of\ solution)}\ x\ 10^6$$

Example1: A sample of 300.0 g of drinking water is found to contain 38 mg of lead. What is the concentration in parts per million?

$$ppm = \frac{(38\ mg\ x\ \frac{1\ g}{1000\ mg})}{300.0\ g}\ x\ 10^6 = 130\ ppm$$

Example 2: A 900.0 g sample of sea water is found to contain 6.7 x 10^{-3} g zinc. Express this concentration in ppm.

$$ppm = \frac{6.7x10^{-3}g}{900.0\ g}\ x\ 10^6 = 7.4\ ppm$$

Note*- the drinking water and sea water IS the solution*

Molality:

$$m = \frac{moles\ of\ solute}{kg\ of\ solvent}$$

Example 1: Suppose we had 1.00 mole of sucrose (it's about 342.3 grams) and proceeded to mix it into exactly 1.00 liter water. It would dissolve and make sugar water. What would be the molality of this solution?

$$m = \frac{1.00\ mole}{(1Lx\frac{1000\ mL}{1\ L}\ x\ \frac{1.00\ g}{1mL}\ x\ \frac{1\ kg)}{1000\ g}} = 1.00\ \frac{mol}{kg} = 1.00m$$

Example 2: What is the molality when 43.83 g NaCl is dissolved in 2.50 L of water?

$$m = \frac{(43.83\ g\ x\ \frac{1\ mol}{58.44\ g})}{(2.50Lx\frac{1000\ mL}{1\ L}\ x\ \frac{1.00\ g}{1mL}\ x\ \frac{1\ kg)}{1000\ g}} = 0.300\ \frac{mol}{kg}$$

Molarity: Most common and most important!

$$M = \frac{moles\ of\ solute}{Liters\ of\ solution} = \frac{moles}{L}$$

Example 1: Calculate the molarity if 9.99 grams of potassium bromide is dissolved in water to make 2.50 liters of solution.

$$9.99\ g\ KBr\ x\ \frac{1\ mole\ KBr}{119.00\ g\ KBr} = 0.08394958\ mol\ KBr$$

$$M = \frac{mol}{L} = \frac{0.08394958\ mol\ KBr}{2.50\ L} = 0.0336\ mol/L$$

Example 2: Calculate M if 5.0 g of sodium phosphate is dissolved in water to make 255 mL of solution.

$$5.0\ g\ Na_3PO_4 x\ \frac{1\ mol}{163.94\ g} = 0.03049896\ mol$$

$$255\ mL\ x\ \frac{1\ L}{1000\ mL} = 0.255\ L \text{ (molarity MUST be in liters)}$$

$$M = \frac{mol}{L} = \frac{0.03049896\ mol}{0.255\ L} = 0.12\ mol/L$$

Example 3: How many grams of NaCl is needed to make 2.0L of a 2.5 M solution? (Use M as a conversion factor as $\frac{2.5\ mol}{1\ L}$. Remember always start with what is NOT a fraction)

$$2.0\ L\ x\ \frac{2.5\ mol}{1\ L}\ x\ \frac{58.44\ g}{1\ mole} = 292.2\ g = 290\ g\ NaCl$$

Example 4: What volume is 13.6 g of a 2.00 M KOH solution?

$$13.6\ g\ x\ \frac{1\ mol\ KOH}{56.105\ g}\ x\ \frac{1\ L\ KOH}{2.00\ mol}\ x\ \frac{1000\ mL}{1\ L} = 121\ mL\ KOH\ solution$$