

Concentration Units

The concentration of a dissolved salt in water refers to the amount of salt (solute) that is dissolved in water (solvent).

Chemists use the term **solute** to describe the substance of interest and the term solvent to describe the material in which the solute is dissolved. For example, in a can of soda pop (a **solution** of sugar in carbonated water), there are approximately twelve tablespoons of sugar (the solute) dissolved in the carbonated water (the solvent). In general, the component that is present in the greatest amount is termed the solvent.

Solution: a mixture consisting of a solute and a solvent

Solute: component of a solution present in the lesser amount

Solvent: component of a solution present in the greater amount

Concentration: amount of a solute present in a solution per amount of solvent

There are many ways to express concentrations. Concentration may be expressed several different ways, using percent composition by mass, volume percent, mole fraction, molarity, molality, or normality. Some of the more common concentration units are:

1. Percent Composition by Mass (%)

This is the mass of the solute divided by the mass of the solution (mass of solute plus mass of solvent), multiplied by 100. It is also called weight percent or percent by weight, this is simply the mass of the solute divided by the total mass of the solution and multiplied by 100%:

$$\text{Percent by mass} = \left(\frac{\text{Mass of component}}{\text{Mass of solution}} \right) (100\%)$$

The mass of the solution is equal to the mass of the solute plus the mass of the solvent. For example, a solution consisting of 30 grams of sodium chloride and 70 grams of water would be 30% sodium chloride by mass: $[(30 \text{ g NaCl}) / (30 \text{ g NaCl} + 70 \text{ g water})] \times 100\% = 30\%$.

To avoid confusion whether a solution is percent by weight or percent by volume, the symbol "w/w" (for weight to weight) is often used after the concentration such as "10% potassium iodide solution in water (w/w)".

Example: Determine the percent composition by mass of a 100 g salt solution which contains 20 g salt.

Solution: $20 \text{ g NaCl} / 100 \text{ g solution} \times 100 = 20\% \text{ NaCl solution}$.

2. Weight/Weight Percent (w/w%)

This unit of concentration is often used for concentrated solutions, typically acids and bases. If you were to look on a bottle of a concentrated acid or base solution the concentration expressed as a weight/weight percent. A weight/weight percent is defined as:

$$\text{w/w}\% = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100$$

Mass/mass concentrations are commonly expressed as parts per million, part per billion or parts per trillion. For example 1 mg of a solute placed in 1 kg of solvent

equals 1 ppm. Part per million by mass (referred to a ppm) is defined as the number of units of mass of a chemical per million units of total mass.

Mass/mass concentrations can also be reported with the units explicitly show (e.g mg/kg, $\mu\text{g}/\text{kg}$). In soil and sediments, 1 ppm equals 1 mg of pollutant per kg of solid (mg/kg) and 1 ppb equals 1 $\mu\text{g}/\text{kg}$. Percent by mass is analogously equals to the number of g pollutant per 100 g total.

3. Volume Percent (% v/v)

Volume percent or volume/volume percent most often is used when preparing solutions of liquids. It is also called volume percent or percent by volume, this is typically only used for mixtures of liquids. Percent by volume is simply the volume of the solute divided by the sum of the volumes of the other components multiplied by 100%:

$$\text{Percent by volume} = \left(\frac{\text{Volume of component}}{\text{Volume of component} + \text{Volume of solvent}} \right) (100\%)$$

Note that volume percent is relative to volume of solution, not volume of *solvent*.

If we mix 30 mL of ethanol and 70 mL of water, the percent ethanol by volume will be 30% BUT the total volume of the solution will NOT be 100 mL (although it will be close). That's because ethanol and water molecules interact differently with each other than they do with themselves.

To avoid confusion whether we have a percent by weight or percent by volume solution, we could label this as "50% ethanol in water (v/v)" where v/v stands for "volume to volume".

The advantage of volume/volume units is that gaseous concentrations reported in these units do not change as a gas is compressed or expanded. Atmospheric concentrations expressed as mass/volume (e.g $\mu\text{g}/\text{m}^3$) decrease as the gas expands, since the pollutant mass remains constant but the volume increases.

For example, 70% v/v rubbing alcohol may be prepared by taking 700 mL of isopropyl alcohol and adding sufficient water to obtain 1000 mL of solution (which will not be 300 mL).

4. Mole Fraction (X)

This is the number of moles of a compound divided by the total number of moles of all chemical species in the solution. Keep in mind, the sum of all mole fractions in a solution always equals 1.

A fraction is defined as a part over a whole. Multiplying this fraction by 100 would give the percent. Thus, a mole fraction involves knowing the moles of solute or component of interest over the total moles of all components in the solution mixture:

$$X_{\text{solute}} = \frac{\text{moles of solute}}{\text{total moles of all components}}$$

Example:

What are the mole fractions of the components of the solution formed when 92 g glycerol is mixed with 90 g water? (molecular weight water = 18; molecular weight of glycerol = 92)

Solution:

90 g water = $90 \text{ g} \times 1 \text{ mol} / 18 \text{ g} = 5 \text{ mol water}$

92 g glycerol = $92 \text{ g} \times 1 \text{ mol} / 92 \text{ g} = 1 \text{ mol glycerol}$

total mol = $5 + 1 = 6 \text{ mol}$

$X_{\text{water}} = 5 \text{ mol} / 6 \text{ mol} = 0.833$

$X_{\text{glycerol}} = 1 \text{ mol} / 6 \text{ mol} = 0.167$

It's a good idea to check your math by making sure the mole fractions add up to 1:

$X_{\text{water}} + X_{\text{glycerol}} = 0.833 + 0.167 = 1.000$.

5. **Molarity (M)**

Molarity is probably the most commonly used unit of concentration. It is the number of moles of solute dissolved in one liter of solution. For example, if we have 90 grams of sucrose (molar mass = 180 grams per mole) this is $(90 \text{ g}) / (180 \text{ g/mol}) = 0.50$ moles of sucrose. If we place this in a flask and add water until the total volume = 1 liter we would have a 0.5 molar solution. Molarity is usually denoted with a capital M, i.e. a 0.50 M solution.

Recognize that molarity is moles of solute per liter of solution, not per liter of solvent!! Also recognize that molarity changes slightly with temperature because the volume of a solution changes with temperature.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{L solution}}$$

Molarity is the most common concentration unit involved in calculations dealing with volumetric stoichiometry.

Example: What is the molarity of a solution made when water is added to 11 g CaCl_2 to make 100 mL of solution?

Solution:

$$11 \text{ g CaCl}_2 / (110 \text{ g CaCl}_2 / \text{mol CaCl}_2) = 0.10 \text{ mol CaCl}_2$$

$$100 \text{ mL} \times 1 \text{ L} / 1000 \text{ mL} = 0.10 \text{ L}$$

$$\text{molarity} = 0.10 \text{ mol} / 0.10 \text{ L}$$

$$\text{molarity} = 1.0 \text{ M.}$$

6. **Molality (m)**

Molality is the number of moles of solute dissolved in one kilogram of solvent. Notice the two key differences between molarity and molality. Molality uses mass rather than volume and uses solvent instead of solution.

Unlike molarity, molality is independent of temperature because mass does not change with temperature. If we were to place 90 grams of sucrose (0.50 moles) in a flask and then add one kilogram of water we would have a 0.50 molal solution. Molality is usually denoted with a small m, i.e. a 0.50 m solution.

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kg solvent}}$$

Molality is often used as the concentration unit involved in calculations dealing with colligative properties, such as freezing point depression, boiling point elevation and osmotic pressure.

Because the density of water at 25°C is about 1 kilogram per liter, molality is approximately equal to molarity for dilute aqueous solutions at this temperature. This is a useful approximation, but remember that it is only an approximation and doesn't apply when the solution is at a different temperature, isn't dilute, or uses a solvent other than water.

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

Solution:

$$10 \text{ g NaOH} / (40 \text{ g NaOH} / 1 \text{ mol NaOH}) = 0.25 \text{ mol NaOH}$$

$$500 \text{ g water} \times 1 \text{ kg} / 1000 \text{ g} = 0.50 \text{ kg water}$$

$$\text{molality} = 0.25 \text{ mol} / 0.50 \text{ kg}$$

$$\text{molality} = 0.50 \text{ m / kg}$$

$$\text{molality} = 0.50 \text{ m.}$$

7. **Normality (N)**

Normality is equal to the *gram equivalent weight* of a solute per liter of solution. A gram equivalent weight or equivalent is a measure of the reactive capacity of a given molecule. Normality is the only concentration unit that is reaction dependent.

Example:

1 M sulfuric acid (H_2SO_4) is 2 N for acid-base reactions because each mole of sulfuric acid provides 2 moles of H^+ ions. On the other hand, 1 M sulfuric acid is 1 N for sulfate precipitation, since 1 mole of sulfuric acid provides 1 mole of sulfate ions.

8. **Mass per unit volume (mg/mL)**

For example, milligrams per milliliter (mg/mL) or milligrams per cubic centimeter (mg/cm^3). Note that $1 \text{ mL} = 1 \text{ cm}^3$ and that cm^3 is sometimes denoted as a "cc".

Mass per unit volume is handy when discussing how soluble a material is in water or a particular solvent. For example, "the solubility of substance X is 3 grams per liter".

In air pollution measurements, the mass of pollutant is expressed as micrograms of pollutant per cubic meter of air (1 millionth of a grams per cubic meter of air), $\mu\text{g}/\text{m}^3$ - mass of species per unit volume of air. Also it can be given in mg/m^3 and mg/l . In water, mass/volume concentration of mg/l and $\mu\text{g}/\text{m}^3$ are common. In most aqueous systems, ppm is equivalent to mg/l .

9. **Parts per million (ppm)**

Parts per million (1 in 1,000,000) works like percent by mass, but is more convenient when there is only a small amount of solute present. ppm is defined as the mass of the component in solution divided by the total mass of the solution multiplied by 10^6 (one million):

$$\text{parts per million} = \left(\frac{\text{Mass of component}}{\text{Mass of solution}} \right) (1,000,000)$$

A solution with a concentration of 1 ppm has 1 gram of substance for every million grams of solution. Because the density of water is 1 g per mL and we are adding such a tiny amount of solute, the density of a solution at such a low concentration is approximately 1 g per mL. Therefore, in general, one ppm implies one mg of solute per liter of solution.

$$\text{ppm} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 10^6$$

Since the amount of solute relative to the amount of solvent is typically very small, the density of the solution is to a first approximation the same as the density of the solvent. For this reason, parts per million may also be expressed in the following two ways:

$$\text{ppm} = \frac{\text{mg of solute}}{\text{L solution}}$$

$$\text{ppm} = \frac{\text{mg of solute}}{\text{kg solution}}$$

At 25 °C and 1 atm (101.3 kPa) pressure, the relationship between Parts per million and micrograms per cubic meter is given by

$$\mu\text{g}/\text{m}^3 = \frac{\text{ppm} \times \text{MW}}{0.0245}$$

where 0.0245 represented the multiplication of RT ($8.205 \times 10^{-5} \times 298$). R is the gas constant in $\text{m}^3 \text{ atm/mole K}$. MW is the molecular weight of the species. For example, concentration of carbon monoxide (MW=28) is 1.5% by volume. It equals 1.5% by mole and equals 10^4 ppm. So when applying the above equation it gives $17.1 \times 10^6 \mu\text{g/m}^3$. Also the concentration of $415 \mu\text{g/m}^3$ of sulfur dioxide (MW=64) equals 0.159 ppm and the concentration of $118 \mu\text{g/m}^3$ of ozone (MW=48) equals 0.060 ppm and the concentration of $100 \mu\text{g/m}^3$ of nitrogen dioxide (MW=46) equals 0.05 ppm.

Finally, recognize that one percent = 10,000 ppm. Therefore, something that has a concentration of 300 ppm could also be said to have a concentration of $(300 \text{ ppm}) / (10,000 \text{ ppm/percent}) = 0.03\%$ percent by mass.

10. **Parts per billion (ppb)**

This works like above, but we multiply by one billion (10⁹; caution: the word billion has different meanings in different countries). A solution with 1 ppb of solute has 1 microgram (10^{-6}) of material per liter.

$$\text{ppb} = \frac{\text{g of solute}}{\text{grams of solution}} \times 10^9$$

Owing to the dilute nature of the solution, once again, the density of the solution will be about the same as the density of the solvent. Thus, we may also express parts per billion as:

$$\text{ppb} = \frac{\mu\text{g of solute}}{\text{L solution}}$$

$$\text{ppb} = \frac{\mu\text{g of solute}}{\text{kg solution}}$$

11. **Parts per trillion (ppt)**

This works like parts per million and parts per billion except that we multiply by one trillion (10^{12}). There are few, if any, solutes which are harmful at concentrations as low as 1 ppt.

Dilutions

You dilute a solution whenever you add solvent to a solution. Adding solvent results in a solution of lower concentration. You can calculate the concentration of a solution following a dilution by applying this equation: $M_i V_i = M_f V_f$

where M is molarity, V is volume, and the subscripts i and f refer to the initial and final values.

Example: How many milliliters of 5.5 M NaOH are needed to prepare 300 mL of 1.2 M NaOH?

Solution:

$$5.5 \text{ M} \times V_1 = 1.2 \text{ M} \times 0.3 \text{ L}$$

$$V_1 = 1.2 \text{ M} \times 0.3 \text{ L} / 5.5 \text{ M}$$

$$V_1 = 0.065 \text{ L}$$

$$V_1 = 65 \text{ mL}$$

So, to prepare the 1.2 M NaOH solution, you pour 65 mL of 5.5 M NaOH into your container and add water to get 300 mL final volume.