

Basic numerical skills: CONCENTRATIONS AND DILUTIONS

1. Introduction (easy)

The term 'concentration' describes the amount of a substance contained within a known quantity of another substance. The most common use of the term is in describing solutions, where one substance (the 'solute') is dissolved in a liquid (the 'solvent'). However, concentration could also apply to mixtures.

The term 'dilution' refers to the action or effect of reducing concentration by adding more of the carrier substance (eg the solvent). In the case of a solution, adding more pure solvent increases the overall volume of the solution, whilst the quantity of the solute remains unchanged. As a result, concentration of the solute decreases.

2. Units (easy)

Concentration is the amount of one substance expressed in relation to the amount of a second substance in which the first is dissolved or dispersed. It is possible to express concentration using different units, depending on how the amounts of the two substances are measured:

Concentrations of solutions are typically measured as the mass of solute per unit volume of solution. Such units might be mg l^{-1} or kg m^{-3} . Note the use of superscripts to indicate that mass is divided by volume, rather than mg/l – there is a separate **NuMBerS** helpsheet covering units if you are unfamiliar with this type of notation.

Sometimes solutions are expressed in relation to the mass of the solution rather than its volume. For instance, older methods for the determination of the salinity (salt content) of seawater expressed the result as a value for grammes of salt per kilogramme of seawater. Strictly speaking, this is dimensionless - it has no units and is just the ratio between two masses. Modern salinity determinations based on electrical conductance are expressed in PSU (practical salinity units) which are equivalent to salt concentration as described earlier.

Mixtures of liquids and gases are typically expressed in terms of volume, describing the proportion of the overall mixture accounted for by one or more of the ingredients. An example would be the oxygen concentration in air, where oxygen is 21% by volume. Note again that this is a ratio, and concentration expressed in this way has no units.

3. Molarity (intermediate)

A concentration of a solution expressed in, say, grammes per litre (g l^{-1}) tells us how much solute is dissolved, but it does not provide information about the reactivity of the solute. Expressing the mass of the solute as a proportion of its atomic or molecular mass is a way of providing this information. An example:

The salt sodium nitrate (NaNO_3) has a molecular mass of 85 (Na = 23, N = 14, O = 16 relative to a carbon atom of mass 12). A mass of sodium nitrate of 85 g is termed one 'mole' of the salt (ie 1 mol is the molecular mass in grammes). A sodium nitrate solution of 85 g l^{-1} can be expressed as 1 mol l^{-1} . Similarly, a solution containing 85 mg m^{-3} would be expressed as 1 mmol m^{-3} , and a solution of $85 \text{ } \mu\text{g l}^{-1}$ would be $1 \text{ } \mu\text{mol l}^{-1}$.

You may see a solution of one mole per litre (1 mol l^{-1}) described as a 'molar solution'. The unit may be abbreviated as 1 M, and subsidiary units may be found such as millimolar (mM) and micromolar (μM).

4. Making solutions of known concentration (easy)

Concentrations of solutions expressed on a mass per unit volume basis (eg mg l^{-1} or mmol m^{-3}) require knowledge of the mass of solute added and the volume of the solution. To make a solution of known concentration, a specified quantity of the solute is weighed out accurately. This is added to a specially calibrated volumetric flask, and solvent is added to make an exact volume of solution. An example:

To make 250 ml of a 8.5 g l^{-1} solution of sodium nitrate, first calculate the amount of solute needed for the volume of solution. The volumetric flask contains 250 ml = 0.25 l. If a litre of solution contains 8.5 g, 250 ml will contain a quarter of that amount of solute. So the mass of solute needed to make the solution will be $0.25 \times 8.5 = 2.125 \text{ g}$.

Dissolving 2.125 g of sodium nitrate in 250 ml of water gives a solution of 8.5 g l^{-1} , which can be expressed in terms of molarity (see previous section) as 0.1 mol l^{-1} .

5. Making a range of standard solutions (easy)

Having made a solution of known concentration, it is possible to make various mixtures of the standard ('stock' solution) and the pure solvent in order to prepare a range of standard solutions. This set of standards can be used to calibrate an analytical technique.

The usual way of making standard solutions is to prepare individual mixtures of the stock standard solution and the pure solvent. Adjustable pipettes are used to provide an exact amount of each. The table below indicates a standard series based on the sodium nitrate solution described in the last section, where a series of 20 ml standard samples have been prepared.

Amount of stock solution (8.5 g l^{-1})	Amount of distilled water	Final concentration (g l^{-1})	Final concentration (mol l^{-1})
0 ml	20 ml	0 g l^{-1}	0 mol l^{-1}
5 ml	15 ml	2.125 g l^{-1}	0.25 mol l^{-1}
10 ml	10 ml	4.25 g l^{-1}	0.5 mol l^{-1}
15 ml	5 ml	6.375 g l^{-1}	0.75 mol l^{-1}
20 ml	0 ml	8.5 g l^{-1}	0.1 mol l^{-1}

If you are unfamiliar with the use of adjustable piston pipettes, you may find it useful to look at a University of Michigan practical schedule at:

<http://www.umd.umich.edu/casl/natsci/slc/slconline/MICRPIP/index.html>

In this example, the same volume of solution (20 ml) is being produced in each case. Sometimes, analytical tests may involve 'spiking' a constant amount of solvent with different amounts of the stock solution. In this case, the final volume of each standard is different, so that the concentration of each of the standard series has to be calculated based on the added solute and the final volume. For instance, if you use a pipette that dispenses 250 μl you can make a standard series like this, by adding different numbers of 'shots' to 10 ml batches of solvent:

Number of 250 μl additions	Total solute added (mg)	Final volume (ml)	Final concentration (mg ml^{-1})	Final concentration (mmol l^{-1})
0	0	10	0	0
1	2.125	10.25	0.207	2.44
2	4.25	10.5	0.405	4.76
3	6.375	10.75	0.593	6.98
4	8.5	11	0.773	9.09
5	10.625	11.25	0.944	11.11

Note that as more shots of standard are added, the final volume increases so that the fractional increase in concentration declines.

6. Measuring solute concentrations in real samples using colorimetry (intermediate)

The concentration of solute is a key measurement in a very wide range of studies. Most techniques involve the use of optical instruments to measure the absorption of the solution. Beer's Law states that the absorbance of light at a particular wavelength is proportional to the amount of the absorbing substance in the solution for a given pathlength through the solution. You can find out more about Beer's Law at this webpage from Sheffield Hallam University:

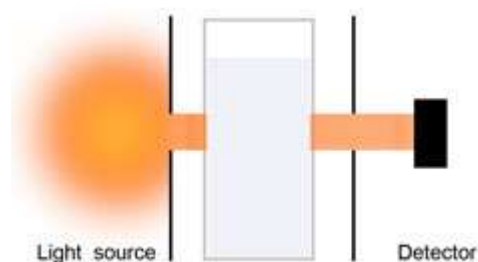
<http://teaching.shu.ac.uk/hwb/chemistry/tutorials/molspec/beers1.htm>

In some cases, the solution has a natural colour, that is stronger the more concentrated the solution. An example would be coloured salt solutions such as copper sulphate (CuSO_4). However, most analyses require that a chemical reaction is set up to produce a coloured end-product whose concentration is proportional to the concentration of the reactant (the solute being tested).

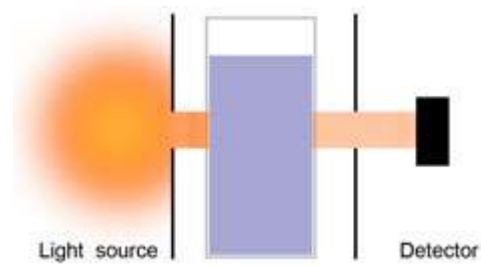
Precision instruments are used to measure the absorption of light. A colorimeter is an instrument that uses light at a single wavelength, whilst a spectrophotometer allows the user to select the wavelength used.

Samples to be measured are placed in cuvettes that are optically identical and have the same pathlength (the distance the light beam passes through the solution). The instrument is zeroed against a sample with all of the reagents but none of the reactant, so that it fulfils the requirements for a blank that is optically equal to real samples with a zero concentration.

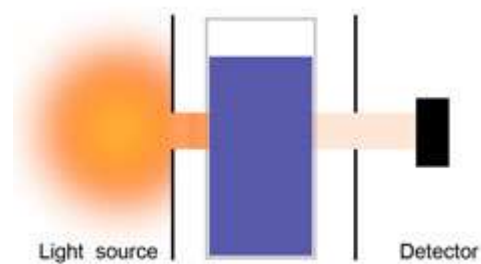
The zero sample (control) can be represented in this diagram, where a light beam passes through a sample in a cuvette, and produces an output from the detector:



A sample with low concentration absorbs more light than the control:



A sample with high concentration absorbs even more light:



The method involves making measurements of a series of standards. These allow the relationship between absorbance and reactant concentration to be established. The data obtained from this calibration are used in a linear regression to obtain an equation relating absorbance and concentration. This can then be used to calculate reactant concentration in real samples from their optical absorbance using the same chemical reaction.