LABORATORY NOTES:

Preparing chemical solutions

A solution is a homogeneous mixture of one or more solute(s) dissolved in a solvent. A solute is a substance that is dissolved in a liquid solvent to produce a solution.

There are a wide range of chemicals used in solution form for various experimental activities and demonstrations in the school science curriculum. These solutions are required in different strengths or concentrations based on intended activity.

The concentration of a solution refers to the quantity of solute dissolved in a particular quantity of solvent or solution. There are several ways of expressing solution concentrations such as moles per litre, percentage concentration, grams per litre, saturated solutions etc.

In a school science laboratory there are 5 main methods used for the calculation of and preparation of solutions.

1. Concentration in moles per litre, molar concentration or molarity (mol/L or mol L^{-1} or M)
2. Concentration by percentage (either %w/v or % v/v or sometimes %w/w)
3. Concentration in grams per litre (g/L or g L^{-1})
4. Preparing solutions by dilution
5. Preparing saturated solutions

1. Concentration in moles per litre, molar concentration or molarity

The mole

The mole is a unit of measurement used to describe the amount of a chemical species. It can be used to describe the number of atoms, molecules, ions, electrons, etc. The abbreviation of mole or moles is *mol*.

One mole contains 6.022×10^{23} particles (atoms, molecules, ions, electrons). This is known as Avogadro's number.

The use of the term *mole* in chemistry is analogous to how the word *dozen* is used in everyday language. For example, one dozen apples is 12 apples, while one mole of apples would be 6.022 x 10^{23} apples.

The weight in grams of one mole of a substance is the molecular weight (MW), or molar mass, of that substance. To determine the number of moles, n, in a given quantity of a substance, you divide the given quantity of the substance by the molecular weight:

$$n = \frac{mass\ of\ substance\ (\ g)}{molecular\ weight\ (g)mol} = \frac{m}{MW}$$

Molarity

In chemistry, molarity is the most frequently used method of expressing concentration of a solution. **Molarity** indicates the number of moles of solute dissolved in a litre of a solution; has the symbol M, and the unit, *moles per litre (mol/L).*

The concentration of a solution in mol/L can be calculated using the formula

$$c \;=\; \frac{n}{v} \;\; (moles\ per\ litre)$$

$$concentration \;= c = \frac{number\ of\ moles\ of\ solute\ (mol)}{volume\ (L)}$$

Where *c* = concentration of the solution in moles per litre (mol/L)

n = amount of moles of solute (mol)

V = volume of solution in litres (L)

The equation $c \;= \frac{n}{V}$ can be rearranged to find the number of moles of solute in a certain volume of a solution of known concentration

$$n \;=\; c\,x\,V$$

To find the volume of a solution of known concentration which will give you a certain number of moles of the solute

$$V \;=\; n/c$$

A molar solution

The symbol M is pronounced 'molar'. Molar solutions use the molecular weight of a solute to calculate molar concentration in a litre of solution.

The molecular weight can be found on the chemical bottle label, in a data book or safety data sheet (SDS), or by adding together the atomic weights of all of the atoms, which appear in the chemical formula of the substance.

e.g. NaCl

1 sodium atom 1 x 22.9 g = 22.99 g

1 chloride atom 1 x 35.45 g = 35.45 g

Molecular weight = 58.44 g

Therefore, a 1M solution of sodium chloride consists of 58.44 g of NaCl dissolved in 1 L of water.

Note: If you are using a hydrated salt, the water(s) of hydration must be included in the calculation of the molecular weight.

Once the molecular weight of a chemical is known the following formula is used to calculate the weight of chemical to dissolve in solution for varying molar solutions.

Combining 2 equations $c = \frac{n}{V}$ and $n = \frac{m}{MW}$ the following formula is obtained

$$m = c \; x \; V \; x \; MW$$

Where:

m = mass of solute in grams (g)

c = concentration of solution in moles per litre) (mol/L)

V = volume of solution in litres (L)

MW = molecular weight of solute in grams (g)

'1M' is pronounced 'one molar' and contains the molecular weight of a chemical dissolved in one litre of water. A '2M solution' is pronounced 'a two molar solution' and has twice the molecular weight of a chemical dissolved in one litre of water.

Example 1: Calculation for preparing 1 litre of a 0.5 M copper (II) sulfate solution

Note that we are using copper (II) sulfate **pentahydrate**, ($CuSO_4.5H_2O$)

$$Molecular\ weight = MW = 63.55 + 32.06 + (4x15.99) + \mathbf{5}((\mathbf{2x1.008}) + \mathbf{15.99})) = 249.68\ g$$
$$Concentration = c = 0.5\ M$$
$$Volume = V = 1\ L$$

The quantity of the solid copper sulfate pentahydrate required to make 1L of 0.5M solution

$$\begin{aligned} m &= c \; x \; V \; x \; MW \\ &= 0.5 \; x \; 1 \; x \; 249.68 \\ &= 124.84\ g \end{aligned}$$

Examples 2 and 3 below compares the use of anhydrous, meaning no water, and monohydrate, meaning having one water of crystallisation, solids of the same compound for preparation of 250 mL of 0.2 M solutions. Note the difference in molecular weights.

Example 2: Calculation for preparing 250 mL of a 0.2 M solution of sodium carbonate using the anhydrous salt (Na_2CO_3).

Before performing the calculation, we convert the volume, 250 mL, to litres, i.e. 0.25 L.

$$Molecular\ weight = MW = 105.99\ g$$
$$Concentration = c = 0.2\ M$$
$$Volume = V = 0.25\ \mathrm{L}$$

The quantity of the solid sodium carbonate required to make 250 ml solution of 0.2 M solution

$$\begin{aligned} m &= c \; x \; V \; x \; MW \\ &= 0.2 \; x \; 0.25 \; x \; 105.99 \\ &= 5.30\ g \end{aligned}$$

Example 3: Calculation for preparing 250 mL of a 0.2 M solution of sodium carbonate using sodium carbonate monohydrate (Na2CO3.H2O).

$$Molecular\ weight = MW = 124.00\ g$$
$$Concentration = c = 0.2\ M$$
$$Volume = V = 0.25\ L$$

The quantity of the solid sodium carbonate monohydrate required to make 250 mL solution of 0.2M

$$\begin{aligned} m &= c \, x \, V \, x \, MW \\ &= 0.2 \, x \, 0.25 \, x \, 124.00 \\ &= 6.20 \, g \end{aligned}$$

Method to prepare a solution using a solid chemical in

1 L of distilled or deionised water

1. Calculate the amount of chemical required to make 1 L of solution at the required molarity.
2. Weigh the amount of chemical using an electronic balance onto a clean, dry watch glass or weighing boat.
3. Carefully transfer the weighed chemical to a beaker containing about two thirds of the final solution volume of distilled or deionised water (about 500-650 mL). When preparing solutions, the solute is dissolved in a portion of the total volume required and then the volume is made up to the required volume.
4. Using a wash bottle containing either distilled or deionised water, wash the watch glass or weighing boat into the beaker to remove all traces of weighed chemical.
5. Stir to dissolve with a stirring rod or on a magnetic stirring platform.
6. It may be necessary to gently heat the solution to speed up the dissolution of the salt.
7. Once dissolved transfer the solution to a 1 L measuring cylinder or volumetric flask.
8. Rinse the beaker, stirring rod and filter funnel using a wash bottle and transfer the washings to the measuring cylinder or volumetric flask.
9. Make up to 1 L with distilled or deionised water. Make sure that the vessel and solution is at room temperature and that the bottom of the meniscus is in line with the mark on the neck of the volumetric flask or the 1 L mark on the measuring cylinder.
10. Stopper and mix thoroughly.
11. Transfer the solution to a labelled reagent bottle.

Note: If preparing a solution directly using a volumetric flask the final volume should be measured with the solution and vessel at room temperature, as this is the temperature at which volumetric glassware is calibrated. If heating the solution, a beaker or conical flask should be used rather than a volumetric flask, as heating volumetric glassware may affect its calibration.

2. Concentration by percentage

The concentration of a solution can be expressed as a percentage concentration, usually as %w/v or %v/v and also sometimes as %w/w.

Where the solute is a solid:

Percentage weight per volume (%w/v) is used. This is the mass of solute (solid) in grams dissolved in 100 ml of solution.

$$\%w/v = \frac{(mass\ of\ solute\ (g))}{volume\ of\ solution\ (mL)} x\ 100\%$$

Alternatively,

%w/v = mass of solute (g) in 100 mL of solution.

Example: a 2%w/v solution of sodium chloride would be prepared from 2 g of sodium chloride dissolved in water and made up to a volume of 100 mL.

Where the solute is a liquid:

Percentage volume per volume (%v/v) is used. This is the volume of solute (liquid) in millilitres per 100 mL of solution.

$$\%\ v/v = \frac{(volume\ of\ solute)}{volume\ of\ solution} x\ 100\%$$

Alternatively,

%v/v = volume of solute (mL) in 100 mL of solution.

Example: a 5%v/v aqueous solution of ethanol would be prepared by taking 5 mL of pure ethanol and diluting this with water to a volume of 100 mL.

Weight percent (%w/w):

This is the mass of solute in grams per 100 g of solution. It is often used in aqueous commercial preparations, for example in concentrated solutions of acids. A weight percent concentration has the advantage that the solution can be prepared independently of temperature considerations. Generally not used in school science.

$$\%w/w\ = \frac{(mass\ of\ solute)}{(mass\ of\ solution)}\ \ x\ 100\%$$

Alternatively,

%w/w = mass of solute (g) in 100 g of solution

3. Concentration in grams per litre

A solution can be prepared by dissolving a known mass or volume of solute in a known amount of solvent.

Concentration is expressed as grams of solute dissolved in one litre of solution.

Example: Calculation for preparing 300 mL of a sucrose solution at a concentration of 5 g/L.

As only 300 mL (0.3 L) of solution is required, only a fraction of the 5 g will be needed.

To find the quantity of sucrose required, the concentration is multiplied by the fraction of litres required:

$$m = 5\ g/L\ x\ 0.3\ L$$

$$m = 1.5g$$

This amount of sucrose is weighed out and dissolved in enough water to make up the volume to a total of 300 mL.

4. Preparing solutions by dilution

Dilution is the process of adding more solvent to a solution.

Solutions can be prepared by diluting a solution of a known higher concentration to produce solutions of lower concentration

Often a stock solution is a concentrated solution that is diluted to a lower concentration for use, called a working solution.

When carrying out a dilution, a definite volume of the concentrated solution is measured out and placed in a volumetric flask of required volume and sufficient solvent is then added to make up to the calibration mark.

It is useful to remember that the dilution does not alter the number of moles of solute present. Only the volume of the final solution changed due to addition of extra solvent. Calculation of the required volume of the initial concentrated solution to produce the diluted solution is based on this fact that the number of moles of solute is the same before and after the dilution.

From the equation $c = \frac{n}{V}$, $n = c\ V$ is obtained.

Where, *n* = amount of moles of solute (mol)

c = concentration of solution in moles per litre (mol/L)

V = volume of solution in litres (L)

Since number of moles of solute is not changed, volume of concentrated solution can be calculated as below.

$$n1(number\ of\ moles\ before\ dilution) = n2(number\ of\ moles\ after\ dilution)$$

$$c1V1 = c2V2$$

$$V1 = \frac{c2\,V2}{c1}$$

Where:

V1 = initial volume or the volume of concentrated solution (in litres)

c1 = concentration of the initial solution or concentrated solution

*V*2 = final volume or the volume of diluted solution (in litres)

c2 = concentration of the final or diluted solution

Example 1: Preparation of 500 mL of 0.5 M Hydrochloric acid (HCl) from a 2 M solution of HCl.

$$Cl)before\ dilution = n2\ (HCL)after\ dilution$$

$$c1V1 = c2V2$$

$$2\ x\ V1 = 0.5\ x\ .5$$

$$V1 = \frac{0.5\ x\ .5L}{2}$$

$$V1 = 0.125\ or\ 125mL$$

This volume of 2 M HCl is measured and placed in a 500 mL volumetric flask containing about 250 mL distilled water. Then enough distilled water is added to make up to the 500 mL mark. The solution should be mixed well to obtain a homogeneous solution of 0.5 M HCl.

Example 2: Diluting hydrogen peroxide (H_2O_2) solution

The strength of hydrogen peroxide is often expressed in volumes. This relates to the volume of hydrogen peroxide and the number of volumes of oxygen gas it can produce upon decomposition.

Hydrogen peroxide (H_2O_2) decomposes to water and oxygen.

$$2H_2O_2 \qquad\qquad 2H_2O + O_2$$

1 volume (1 mL) of 20-volume hydrogen peroxide will produce 20 volumes (20 mL) of oxygen as a gas.

For example, taking 10 mL of 20 volume hydrogen peroxide, decomposition of the hydrogen peroxide will produce 20 x 10 mL = 200 mL of oxygen gas.

Both 100 volume (30% solution) and 120 volume (35% solution) concentrations are commercially available to purchase. 20 volume (6%) concentration is the common strength of hydrogen peroxide used for science investigations in schools.

Calculation for the preparation of 1L of a 20 volume (6%) solution of hydrogen peroxide from 100 volume (30%) solution. 100 volume is the initial concentration and 20 volume is the final concentration.

$$1 = c2V2$$

$$V1 = \frac{c2V2}{c1}$$

$$V1 = \frac{20\ x\ 1}{100}$$

$$V1 = 0.2L\ or\ 200\ mL$$

200 mL of 100 volume hydrogen peroxide is added to approximately 500 mL of distilled water and made up to 1L. The diluted solution should be mixed well.

Alternatively, the calculation using percentage concentrations also can be used. 30% is the initial concentration and 6% is the final concentration.

$$c2V2$$

$$V1 = \frac{c2V2}{c1}$$

$$V1 = \frac{6\%\ x\ 1}{30\%}$$

$$V1 = 0.2L\ or\ 200\ mL$$

5. Saturated solutions

The solubility of a solute is the maximum amount that can dissolve at a specified temperature in a specified volume of a particular solvent. Solubility depends on type of solute, type of solvent and temperature.

A saturated solution is where at a particular temperature no more solute can be dissolved in the solvent. This is where despite stirring no more of the solute can be dissolved and the excess settles in the bottom of the beaker.

The solubility of most solids increases with increasing temperature, so a saturated solution prepared at a high temperature will contain more dissolved solute than it would contain at a lower temperature. If a saturated solution prepared at a high temperature is cooled then the solute will come out of solution relative to the temperature to which it cools.

Before preparing a saturated solution of a chemical it is necessary to determine its solubility in a given amount of solute at a particular temperature.

The solubility of most chemicals (how much will dissolve in a solvent at a given temperature) can be obtained from a chemical data book, the safety data sheet or from a solubility curve of the chemical.

A solubility curve compares the amount of solute that will dissolve in a given amount of solvent at various temperatures. Generally the solvent is water and the concentration is provided in grams of solute in 100 g of solvent. Solubility curves are different for different chemicals.

Example: A saturated solution of sucrose

The solubility of sucrose at 20^0C is 203.9 g / 100 mL of water. The solubility of sucrose at 100^0C is around 500 g / 100 mL of water.

If you try to dissolve 220 g of sucrose in 100 mL of water at 20^0C then 203.9 g will dissolve forming a saturated solution and the remaining 16.1 g will settle to the bottom of the beaker. The undissolved solute can be separated from the saturated solution by filtration.

References and further reading

Dungey, Barbara. 2003. *The laboratory – A Science reference and preparation manual for schools*. Rev. ed. R O & B Dungey: Traralgon, Victoria

'Laboratory solution preparation', Flinn Scientific website, https://www.flinnsci.com/laboratory-solution-preparation/dcat016/ (Accessed November 2018)