EXPRESSIONS OF CONCENTRATION

▶ The concentration of a solution may be expressed using a variety of units. These units include:

- **% m/m**, which means the ***mass in grams*** of solute in ***100 g of solution.***
- **% m/V**, which means the ***mass in grams of solute*** in ***100 mL of solution***
- **% V/V**, which means the ***volume in mL*** of solute in ***100 mL of solution.***
- **ppm**, which means the mass in gram of solute in 10^6 g of solution or the ***mass of a solute in mg*** per ***kg of solution*** (equivalent to mg L^{-1} for dilute solutions).
- **concentration in mol L^{-1}**, which means the ***amount in mole of solute in 1.0 L*** of solution.
- **concentration in g L^{-1}**, which means the mass in ***gram of solute in 1.0 L*** of solution. With very small solute concentrations the mass of solute may be measured in **mg L^{-1}** (10^{-3} gram) or **µg L^{-1}** (10^{-6} gram).

EXAMPLE 4:

The % mass of hydrogen peroxide in a solution is 0.3% (**% m/m**; % by mass). Given that the density of peroxide solution is 1.03 g mL^{-1} calculate:

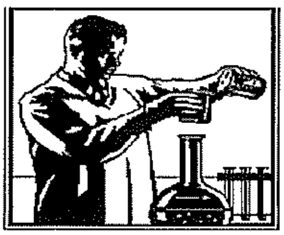
(a) The concentration in **mol L^{-1}**.
(b) The concentration in **g L^{-1}**.
(c) The concentration in **µg mL^{-1}**.
(d) The concentration in **mg L^{-1}**.
(e) The concentration in **ppm**.

(a) **MOLAR MASS:*

(H_2O_2)

1.008 X 2
16.00 X 2

34.02 g/ mol



$m(H_2O_2)$ per mL $= \frac{0.3}{100} \times 1.03$

$= 0.003 \times 1.03$

$= 0.00309$ g

$m(H_2O_2)$ per L $= 1{,}000 \times 0.00309$

$= 3.09$ g

$n(H_2O_2)$ per L $=$ m/M

$= 3.09 / 34.02$

$= 0.09082 = 0.0908$ mol

▶ **$[H_2O_2] = 0.0908$ M**

(b) $m(H_2O_2)$ per mL = $\frac{0.3}{100} \times 1.03$

= 0.003×1.03

= <u>0.00309 g</u>

$m(H_2O_2)$ per L = $1{,}000 \times 0.00309$

= <u>3.09 g</u>

▶ **$[H_2O_2] = 3.09$ g L^{-1}**

(c) ▶ **$[H_2O_2] = 3.09$ g L^{-1}**

■ **There are 1,000,000 µg in a gram!**

■ **There are 1,000 mL in a Litre!**

▶ $[H_2O_2]$ = $3.09 \times 1{,}000{,}000$ µg L^{-1}

= 3,090,000 µg L^{-1}

= $\frac{3{,}090{,}000}{1{,}000}$ µg mL^{-1}

▶ **$[H_2O_2]$ = 3,090 µg mL^{-1}**

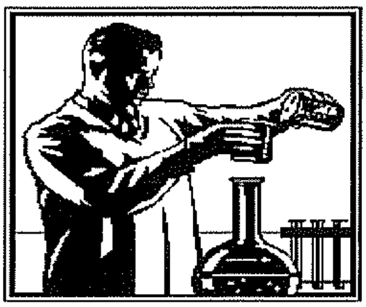
(d) ▶ **$[H_2O_2] = 3.09$ g L^{-1}**

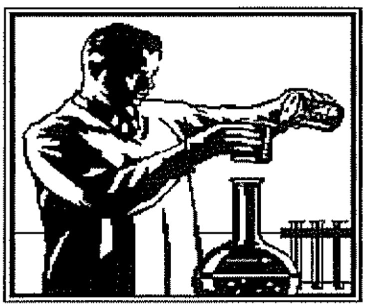
■ **There are 1,000 mg in a gram!**

▶ $[H_2O_2]$ = $3.09 \times 1{,}000$ mg L^{-1}

= 3,090 mg L^{-1}

▶ **$[H_2O_2]$ = 3,090 mg L^{-1}**





(e) ▶ <u>$[H_2O_2]$ = 3.09 g L^{-1}</u>

■ **The density of the solution is 1.03 g mL^{-1}**

■ **We need to know the MASS of a LITRE of solution in kilograms to be able to calculate the ppm!**

Mass 1 Litre = 1,000 x 1.03

<u>= 1,030 g Or 1.03 kg</u>

■ **We need to know the mass of solute in milligrams per Litre.**

▶ **$[H_2O_2]$ = 3.09 g L^{-1}**

= 3.09 x 1,000 mg L^{-1}

= 3,090 mg L^{-1}

▶ ***There are 3,090 milligams in 1.03 kilograms of solution***

▶ **$[H_2O_2] = \frac{3{,}090}{1.03}$**

$[H_2O_2]$ = <u>3,000 ppm</u>

PARTS PER MILLION / DENSITY

- It is important to note that parts per million concentration is a **MASS based comparison** between the **mass of a solute** and the **mass of the solution.** If the solute is measured in milligrams (mg) and the solution in kilograms (kg) then the masses will be in a 1: 1,000,000 ratio (1 kg = 1,000,000 mg). This unit is used when the solute is in small amount.

- It is not possible to calculate the mass of a solution given only its volume. If a density value is given it allows us to work out what a certain volume of a solution will weigh! As we work in molar concentration most of the time, which is a mole amount per litre of solution, it is best to always work with the mass of a **LITRE of solution** for conversion to ppm!

- It should be noted that a gas mixture is also a solution so **ppm** is used to determine the amount of a pollutant gas in a larger mass of air.

EXAMPLE 5:

Sulfur dioxide is used as a preservative in white wine. A sample of a white wine was analysed in a titration and found to have a molar concentration of 0.000125 M of SO_2. Given that the **density** of the wine is 1.02 gmL^{-1} . Determine the concentration of SO_2 in **parts per million**.

> ***To convert grams to kilograms you simply divide bv 1.000!***

MASS of 1 Litre $=$ *Density* (gml^{-1}) *x Volume (mL)* $= 1.02 \times 1,000$

$= \underline{1,020\ g\ \text{or}\ 1.02\ kg}$

> **** $\underline{SO_2}$***
>
> ***32.06***
>
> ***$\underline{+ 2 \times 16.00}$***
>
> ***$\underline{64.06\ g\ mol^{-1}}$***

$[SO_2] = 0.000125\ M$

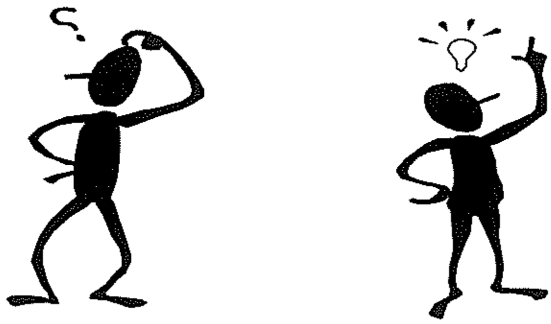
$m(SO_2)\ /\ Litre = n \times M$

$= 0.000125 \times 64.06$

> ***To convert grams to milligrams you simply multiply by 1,000!***

$\underline{= 0.008008\ g} = \underline{8.008\ mg}$

$\therefore [SO_2] = \frac{8.008}{1.02} = 7.850\ ppm = \underline{7.85\ ppm}$ (3 S.F. !)



EXAMPLE 6:

If the sulphur dioxide concentration in a 10,000 kg air sample is 1,500 ppm. What mass of the gas would there be in the air sample?

▶ The best approach to these problems is to represent the ppm concentration as a fraction or decimal and relate to the total mass.

$$1,500\ ppm = \frac{1,500}{1,000,000} = 0.0015$$

▶ $\underline{m(SO_2) = 0.0015 \times 10,000\ kg = 15kg}$

DILUTION PROBLEMS

► If extra water is added to a solution of known concentration it is important to realise that the ***moles of dissolved chemical does not change,*** only the concentration does.

- A range of problems concern situations where it is necessary to find out how much water must be added to change the concentration of a solution from its original value to a different value. This sort of problem may be approached two ways:

EXAMPLE 7:

How much water must be added to 300 mL of a 0.150 M salt solution (NaCl) to produce a solution with a concentration of 0.0745 M?

APPROACH 1: MOLES APPROACH

**** Calculate the number of moles of salt (NaCl) solute and relate that to the desired solution concentration to find out the volume of the new solution in which this many moles may produce the right concentration value. Once this is arrived at you can calculate the volume differential with the original solution and work out how much water to add by subtraction!***

**** n(NaCl) is the SAME in both solutions***

$$^{*}n(NaCl) = C \times V = 0.300 \times 0.150 = \underline{\underline{0.0450\ mol}}$$

⇨

$$[NaCl]_{new} = \frac{^{*}n}{V}$$

$$\blacktriangleright\ V(NaCl)_{new} = \frac{^{*}n}{[NaCl]_{new}} = \frac{0.0450}{0.0745} = \underline{\underline{0.604\ L\ \text{or}\ 604\ mL}}$$

- Volume of new solution is 604 mL.
- Volume of old solution was 300 mL.

► Volume of Water Added = 604 – 300 = **304 mL**

APPROACH 2: EQUATE CONCENTRATION EQUATIONS

**** If n(NaCl) is the SAME in both solutions then the product of concentration and volume must be the same for both solutions (*** $n = C_1V_1 = C_2V_2$ ***).***

$$C_1V_1 = C_2V_2$$

$$V_2 = \frac{C_1\ V_1}{C_2} = \frac{0.150 \times 0.300}{0.0745} = 0.604\ L\ \text{or}\ 604\ mL$$

► *Volume of Water Added* = $V_2 - V_1$ = 604 – 300 = **304 mL**