CONCENTRATIONS OF SOLUTIONS

Lec.7 & 8

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We describe the relative amounts of solute and solvent in a solution by means of units of concentration. There are several such units, and we will examine the most commonly used ones.

Weight / Weight Percent

One way to specify the concentration of a solute in a solution is as a *percent by weight*. The concentration of the solute is given by the following equation :

$$\text{Percent by weight solute} = \frac{\text{weight of solute, in g}}{\text{weight of solute, in g} + \text{weight of solvent, in g}} \times 100$$

Example 8-1 : What is the percent by weight of sugar in a solution made by dissolving 10 g of sugar in 90 g of water?

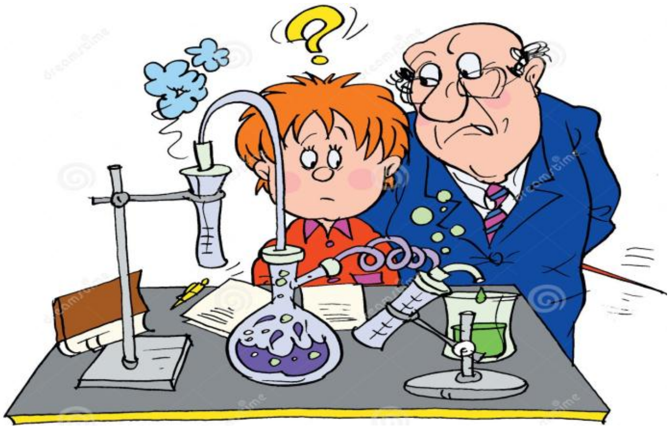
Solution:

$$\text{Percent by weight sugar} = \frac{10\ \text{g}}{10\ \text{g} + 90\ \text{g}} \times 100 = \frac{10\ \cancel{\text{g}}}{100\ \cancel{\text{g}}} \times 100 = 10\%$$

•EXERCISE 8-1 Determine the percent by weight of the solute in each of the following solutions:

(a) 1.50 g of sodium chloride in 100 g of water

(b) 3.50 g of glucose in 250 g of water



Volume/Volume Percent

A convenient way of expressing the concentration of a liquid solute dissolved in a liquid is as a *percent by volume.* This unit of concentration is similar to percent by weight except that volumes in milliliters are used instead of weights in grams. The equation is as follows:

$$\text{Percent by volume solute} = \frac{\text{volume of solute, in mL}}{\text{total volume of solution, in mL}} \times 100$$

Example 8-2 : What is the percent by volume of ethyl alcohol in a solution made by diluting 10 mL of ethyl alcohol to 100 mL with water?

Solution:

$$\text{Percent by volume ethyl alcohol} = \frac{10\,\cancel{\text{mL}}}{100\,\cancel{\text{mL}}} \times 100 = 10\%$$

EXERCISE 8-2: Determine the percent by volume of the solute in each of the following solutions:

(a) 5.0 mL of rubbing alcohol diluted to 150 ml with water

(b) 15 mL of ethyl alcohol diluted to 500 mL with water

Weight / Volume Percent

This widely used method of expressing concentrations is a combination of weight and volume. The weight is usually that of the solid solute and the volume is that of the total solution. This unit is defined as follows:

Example 8-3 ; What is the percent by weight/volume of sodium chloride in a solution made by diluting 1.5 g of sodium chloride to 100 mL with water?

Solution:

$$\text{Percent by weight/volume solute} = \frac{\text{weight of solute, in g}}{\text{total volume of solution, in mL}} \times 100$$

$$\text{Percent by weight/volume NaCl} = \frac{1.5\text{ g}}{100\text{ mL}} \times 100 = 1.5\%$$

Low concentrations of solute are often expressed in ***milligrams per 100 mL.*** This weight/volume percent unit is defined as follows:

$$\text{Milligrams per 100 mL} = \text{mg/100 mL} = \frac{\text{weight of solute, in mg}}{\text{100 mL of solution}}$$

The unit mg/100 mL is sometimes called **mg percent.**

This unit is often used to express the concentrations of solute in blood and urea, as shown in the following example:

Example 8-4 A 1-mL sample of blood plasma is found to contain 3.3 mg of sodium ions. Express this concentration in mg/100 mL.

Solution:

STEP 1. Express the amount of solute in mg per total volume of solution:

Solute = 3.3 mg

Solution = 1 mL of blood plasma

$$\frac{3.3\text{ mg}}{1\text{ mL}}$$

Sorry Professor, you're right:
I DID skip a line of the instructions...

STEP 2. The definition of mg/100 mL is the weight of solute in 100 mL of solution. We know how many mg are in 1 mL of blood plasma. To find out the number of mg in 100 mL of blood plasma, we set up the following proportion:

$$\frac{3.3\ \text{mg}}{1\ \text{mL}} = \frac{X}{100\ \text{mL}}$$

STEP 3. Rearrange and solve for X.

$$X = \frac{3.3\ \text{mg} \times 100\ \cancel{\text{mL}}}{1\ \cancel{\text{mL}}} = 330\ \text{mg}$$

There are 330 mg of sodium ions in 100 mL of blood plasma. Therefore, according to equation 8-4,

$$\text{Milligrams per 100 mL} = \frac{330\ \text{mg}}{100\ \text{mL}}$$

Exercise 8-4 Determine the concentration of solute in each of the following solutions in mg/100 mL.

(a) 32.0 mg of sugar diluted to 10.0 mL with water

(b) 5.00 mL of solution that contains 1.00 g of sodium ion

<u>Parts Per Million and Parts Per Billion</u>

**These units of concentration are widely used to report very small amounts of solute in a solution. The concentration of pollutants in water and air are usually reported in these units.

--One part per million, abbreviated ppm, contains 1 part of solute per 1 million (10^6) parts of solution. By parts we mean any unit of measure such as grams, liters, or anything else we choose.

For example, the concentration of solid pollutants in solid food is given in ppm expressed as mg of pollutant (the solute) in **1 million mg of solid food (the solution). Because 1 million mg is equivalent to 1 kg, ppm is usually defined as follows:

$$\text{ppm} = \frac{\text{weight of solute, in mg}}{\text{weight of solution, in kg} = 10^6\ \text{mg of solution}}$$

To express the concentrations of small quantities of solid solutes in water, the unit **<u>ppm</u>** is usually *defined* as **mg of solute per liter of solution.** This change from weight to volume of solvent can be made because 1 million mg (1 kg) of water occupies approximately 1 L. This definition of ppm is also frequently used even though the solution may weigh somewhat more or less than 1 kg.

$$ppm = \frac{wt.\ of\ solute\ (g) * 10^6}{wt.of\ solution\ (g)}$$

$$ppm = \frac{wt.\ of\ solute\ (g) * 10^6}{Vol.of\ solution\ (ml)}$$

**Air pollution is measured in ppm on the basis of measurements of volume rather than weight. Thus, 1 ppm means that there is 1 µL of pollutant (the solute) per 1 million (10^6) μL (1 L) of air (the solution).

** The sensitivity of analytical methods has improved so much that **p**arts **p**er **b**illion, abbreviated **ppb**, has become a common unit of concentration.

**Its use and definition are similar to those of ppm. Thus, 1 ppb contains 1 part of solute per 1 billion (10^9) parts of solution. Again, the parts refer to weight or volume, depending on whether the solution is a gas, liquid, or solid.

Example 8-5 The maximum Food and Drug Administration (FDA) tolerance of mercury in fish is 0.5 ppm. A 10-g sample of fish is found to contain 72 μg of mercury. Does the amount of mercury in the fish exceed the FDA maximum tolerance?

<u>Solution:</u>

$$ppm = \frac{wt.\ of\ solute\ (g) * 10^6}{wt.of\ solution\ (g)}$$

$$ppm = \frac{72 * 10^{-6}\ (g) * 10^6}{10\ (g)} = 7.2$$

Another solution :

Example 8-5 The maximum Food and Drug Administration (FDA) tolerance of mercury in fish is 0.5 ppm. A 10-g sample of fish is found to contain 72 μg of mercury. Does the amount of mercury in the fish exceed the FDA maximum tolerance?

Solution:

STEP 1. Calculate the mg of mercury in 1 kg of this fish using the following proportion:

$$\frac{72\ \mu g}{10\ g} = \frac{X}{1\ kg}$$

STEP 2. Rearrange and solve for X:

$$X = \frac{72\ \cancel{\mu g}}{10\ \cancel{g}} \times \cancel{1\ kg} \times \frac{10^3\ \cancel{g}}{1\ \cancel{kg}} \times \frac{1\ mg}{10^3\ \cancel{\mu g}} = 7.2\ mg$$

STEP 3. Determine ppm from its definition:

$$ppm = \frac{\text{weight of solute, in mg}}{\text{1 kg of solution}} = \frac{\text{7.2 mg of mercury}}{\text{1 kg of fish}} = 7.2\ ppm$$

The amount of mercury in the fish exceeds the maximum FDA tolerance. The fish is contaminated with mercury.

Molar Concentrations (Molarity) Molar concentration, or molarity, which is defined as ***the number of moles of solute per liter of solution, is designated by the capital letter*** **M.** This definition is given in the form of an equation, as follows:

$$M = \frac{\text{number of moles of solute}}{\text{number of liters of solution}}$$

$$M = \frac{wt.\ of\ solute\ (g) * 1000}{M.wt. * Vol.(ml)}$$

Example 8-6 Determine the molar concentration of a solution that contains 25.0 g of glucose, $C_6H_{12}O_6$, in 500 mL of solution.

Solution:

$$M = \frac{wt.\ of\ solute\ (g) * 1000}{M.wt. * Vol.(ml)}$$

$$M = \frac{25 * 1000}{180 * 500} = 0.278$$

H.W (3-5) Many fish are contaminated with the insecticide DDT. One such fish is the Coho salmon from Lake Michigan. A 20-g sample of Coho salmon was found to contain 0.010 mg of DDT. Express this amount of DDT in ppm. **(p.173)**

***Notice that the molecular weight of the solute is not needed when we express the concentration of the solute in any of the units of percent, ppm, or ppb. Thus, we can specify the concentration of a substance without knowing anything about its chemical composition.

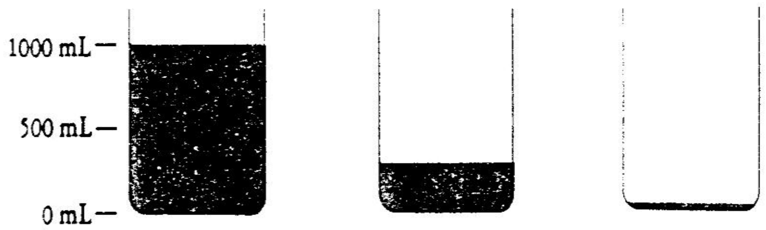
H.W :EXERCISE 8-6 Determine the molar concentration of each of the following

solutions:

(a.) 40.0 g of NaOH made up to 1.0'L with water

(b.) 250 mL of a solution that contains 5.40g of NaCl

** Molar concentration expresses the ratio of solute to solution. Two solutions that have the same molar concentrations have the same ratios of solute to solution even though the total volumes of the two solutions may be different. We can demonstrate this important fact by means of the three solutions of glucose in water shown in Figure 8-1.



	Solution A	Solution B	Solution C
Weight of glucose, g	90.0	22.5	0.450
Amount of glucose, mol	0.500	0.125	0.00250
Volume of solution, mL	1000	250	5.00
Volume of solution, L	1.00	0.250	0.00500
Molar concentration	$\frac{0.500\ \text{mol}}{1.00\ \text{L}} = 0.500\ \text{M}$	$\frac{0.125\ \text{mol}}{0.250\ \text{L}} = 0.500\ \text{M}$	$\frac{0.00250\ \text{moi}}{0.00500\ \text{L}} = 0.500\ \text{M}$

Fig. 8-1. Three different volumes of solutions of the same concentration.

**Thus, a bottle labeled 0.500 M glucose may contain 10 L or as little as 1 mL. No matter how much solution there is in the bottle, every drop of it has a glucose concentration of 0.500 M.

**The importance of molar concentration is that we can determine the weight of the solute contained in any volume of solution. This fact is in the following examples.

Example 8-7:A patient is fed intravenously 0.50 L of a 1.0 M glucose solution. How many grams of glucose has the patient received?

Solution:

$$M = \frac{wt.\ of\ solute\ (g) * \ 1000}{M.wt. * Vol.(ml)}$$

$$1 = \frac{wt. * \ 1000}{180 * 0.50 * 1000} = 90\ grams$$

H.W. EXERCISE *8-7 ;*How many grams are needed to make each of the following solutions?

1.00 L of a 0.100 M NaCl solution

a. 250 mL of a 1.50 M glucose solution

b. 500 mL of a 0.150M sucrose ($C_{12}H_{22}O_{1:l}$) solution

EXERCISE 8-8 ;How many grams of solute are there in each of the following quantities of solution?

a. 25.0 mL of a 1.00 M LiBr solution

b. 100 mL of a 0.500 M NaOH solution

c. 250 mL of a 0.100 M NaHCO, solution

Milliequivalents Per Liter

This unit is used to ***express low concentrations of ions in body fluids***. To use this unit, we must learn about ***equivalents***. **One equivalent** of an ion, abbreviated **Eq**, is defined ***as 1 mole of that ion multiplied by the absolute value of its charge.***

** For example, 1 mole of sodium ions contains one equivalent of sodium ions. One mole of chloride ions contains one equivalent of chloride ions. One mole of magnesium ions contains two equivalents of magnesium ions.

EXERC1SE 8-9 : How many equivalents are there of each underlined ion in the following quantities?

a. 1.00 mol $\underline{Na}HCO_3$

b. 0.150 mol $Na_2\underline{CO_3}$

c. 2.00 mol $Li\underline{Br}$

d. 0.350 mol $\underline{Fe}Cl_3$

e.1.50 mol $Mg\underline{Cl_2}$

d. 0.250 mol $Ca_3\underline{(PO_4)_2}$

The unit **milliequivalent per liter**, abbreviated **mEq/L**, is defined as follows:

$$\text{Milliequivalent per liter (mEq/L)} = \frac{\text{number of milliequivalents of ion}}{\text{volume of solution, in L}} \quad (8\text{-}6)$$

Example 8-9 Express the concentrations of sodium and chloride ions in a 0.001 M NaCl solution in terms of milliequivalents per liter.

Solution:

STEP 1. A 0.001 M NaCl solution contains $\frac{0.001 \text{ mol NaCl}}{\text{L}}$.

$$\frac{0.001 \text{ mol NaCl}}{\text{L}} = \frac{0.001 \text{ mol Na}^+}{\text{L}}$$

and

$\frac{0.001 \text{ mol Cl}^-}{\text{L}}$ (See Section 8.5.)

STEP 2. $1 \text{ mol Na}^+ = 1 \text{ Eq Na}^+ \text{ or } 1 = \frac{1 \text{ Eq Na}^+}{1 \text{ mol Na}^+} = \frac{1 \text{ mEq Na}^+}{0.001 \text{ mol Na}^+}$

$$1 \text{ mol Cl}^- = 1 \text{ Eq Cl}^- \text{ or } 1 = \frac{1 \text{ Eq Cl}^-}{1 \text{ mol Cl}^-} = \frac{1 \text{ mEq Cl}^-}{0.001 \text{ mol Cl}^-}$$

STEP 3. $\frac{\cancel{0.001 \text{ mol Na}^+}}{\text{L}} \times \frac{1 \text{ mEq Na}^+}{\cancel{0.001 \text{ mol Na}^+}} = \frac{1 \text{ mEq Na}^+}{\text{L}}$

$$\frac{\cancel{0.001 \text{ mol Cl}^-}}{\text{L}} \times \frac{1 \text{ mEq Cl}^-}{\cancel{0.001 \text{ mol Cl}^-}} = \frac{1 \text{ mEq Cl}^-}{\text{L}}$$

Another solutions

$$M = \frac{\text{number of moles of solute}}{\text{number of liters of solution}}$$

$$0.001 = \frac{no.\,of\ moles}{1\ L}$$

no. of moles = 0.001

so

no. of moles of **Na^+** = 0.001

no. of mole of **Cl^-** = 0.001

no. of Eq of ion = 1 mole of ion * | its charge |

no. of Eq of **Na^+** = 0.001 * | 1^+ | = 0.001

no. of Eq of **Cl^-** = 0.001 * | 1^- | = 0.001

no. of mEq of **Na^+** = 0.001 * 10^3 = 1 (1 mole = 10^3 mmole)

no. of mEq **of** **Cl^-** = 0.001 * 10^3 = 1

$$\text{Milliequivalent per liter (mEq/L)} = \frac{\text{number of milliequivalents of ion}}{\text{volume of solution, in L}}$$

$$mEq/L\ of\ \text{Na} = \frac{1\ mEq\ Na}{L}$$

$$mEq/L\ of\ \text{Cl} = \frac{1\ mEq\ \text{Cl}}{L}$$

EXERCISE 8-10 Determine the concentration of each ion, in mEq/L, in each of the following solutions:
(a) 0.001 M LiBr (b) 0.05 M Na_2CO_3

EXERCISE 8-11 If enough water is added to 25.0 g of $NaHCO_3$ to make 5.00 L of solution, what is the concentration of HCO_3^- ion, in mEq/L?

The various units of concentration are summarized in Table 8-3.

We have learned in this section how to express the concentration of a solution in several different units. The practical use of these units in clinical

Table 8-3. Summary of Concentration Units

Units	Definition
Weight/weight percent	$\dfrac{\text{Weight of solute, in g}}{\text{Weight of solute, in g + weight of solvent, in g}} \times 100$
Volume/volume percent	$\dfrac{\text{Volume of solute, in mL}}{\text{Total volume of solution, in mL}} \times 100$
Weight/volume percent	$\dfrac{\text{Weight of solute, in g}}{\text{Total volume of solution, in mL}} \times 100$
Milligram/100 mL	$\dfrac{\text{Weight of solute, in mg}}{\text{100 mL of solution}}$
Parts per million	$\dfrac{\text{Weight of solute, in mg}}{\text{Weight of solution, in kg}}$ or $\dfrac{\text{mg of solute}}{\text{L of solution}}$ or $\dfrac{\mu L}{L}$
Parts per billion	$\dfrac{\text{Weight of solute, in } \mu g}{\text{Weight of solution, in kg}}$ or $\dfrac{\mu g \text{ of solute}}{\text{L of solution}}$ or $\dfrac{\mu L}{10^3\ L}$
Molar concentration (molarity)	$\dfrac{\text{Number of moles of solute}}{\text{Number of liters of solution}}$
Milliequivalents per liter	$\dfrac{\text{Number of milliequivalents}}{\text{Volume of solution, in L}}$

Table of concentration measures

Frequently used standards of concentration			
Measurement	**Notation**	**Generic formula**	**Typical units**
Mass percentage	wt%	$\left(\frac{\text{grams solute} \times 100}{\text{grams solution}}\right)$	%
Mass-volume percentage	-	$\left(\frac{\text{grams solute} \times 100}{\text{milliliters solution}}\right)$	% *though strictly* %g/mL
Volume-volume percentage	vol%	$\left(\frac{\text{milliliters solute} \times 100}{\text{milliliters solution}}\right)$	%
Molarity	*M*	$\left(\frac{\text{moles solute}}{\text{liters solution}}\right)$	mol/L (or M or mol/dm^3)
Parts per million	ppm	$\left(\frac{\text{milligrams solute}}{\text{liters solution}}\right)$	mg/L
Parts per billion	ppb	$\left(\frac{\text{micrograms solute}}{\text{kilograms solution}}\right)$	µg/kg

الكتاب المنهجي:

The Chemical Basis of Life: General Organic and Biological Chemistry

By:George H. Schmid