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# 1. SOLUTIONS. CONCENTRATIONS. IONIC FORCE OF THE STRONG ELECTROLYTE SOLUTIONS AND THEIR MIXTURES (PREPARATION OF PLASMA SUBSTITUTES)

Solutions are homogeneous mixtures, which contain dissolved substance, *solute* and *solvent*. The size of particles in the true solutions is less than  $10^{-9}$  m (or  $<1$  nm), which means the simplest small molecules or ions are present in a true solution, in contrast to colloidal dispersions. (For example, glucose solution contains molecules, sodium chloride solution — ions).

There are three states of matter, thus solutions may be solid (alloys), liquid (biological fluids) and gaseous (air). The component which does not change the aggregative state during the preparation of the solution is called a solvent. (For example, water is a solvent when we prepare the solution of sugar in water). If both components have the same aggregative state, we consider that the solvent is present in greater amount (e.g., acetic acid dissolved in water forms vinegar. In this case acetic acid is a solute, and water is a solvent).

Actually, solutions take an intermediate position between the barely mixtures and compounds, and the following effects may prove that an interaction between the molecules or ions of the solution takes place:

- 1) heat effect (when we dissolve ammonium nitrate, the beaker with freshly-prepared solution becomes cold, in case of sulphuric acid — vice versa);
- 2) change in volume (when we prepare the solution of alcohol in water, the resulting volume of solution is always less than expected; some substances, vice versa, may increase the final volume);
- 3) change in colour (compare the colour of  $\text{CuSO}_4$ , white if no water is present, and the solution of this salt, which is blue).

Interaction of solvent with solute leads to the formation of solvates (or hydrates in water solution) due to electrostatic forces, donor-acceptor and hydrogen bonds. Some hydrates are rather stable and popular in medicine. For example, we use the crystalline hydrate  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  to prepare solutions which are used as vasodilators (hypotensive effect), anticonvulsants, decongestants (anti-edematous action), for the labor pain relief; analgetic, sedative, laxative drugs and the part of polarizing mixtures for the various groups of patients.

Solubility of a substance depends on the nature of solvent, temperature, pressure (for gases). Solubility coefficient ( $k$ ) shows how many grams of matter may be dissolved in 1 L of water to reach the maximum degree of saturation at constant temperature (visual effect — the sediment of substance). We know 3 groups of substances by solubility:

- 1) soluble ( $k > 10$  gr/L, like  $\text{CaCl}_2$ , sugar, table salt);
- 2) slightly soluble ( $k = 0,01 - 10$  gr/L, like  $\text{CaSO}_4$ , silver sulfate);
- 3) insoluble ( $k < 0,01$  gr/L, like  $\text{Ca}_3(\text{PO}_4)_2$ , chalk).

Lots of solids increase solubility with temperature raise. Solubility of gases, vice versa, decreases at high temperatures. That is why, in cold sea water animals have a much greater supply of oxygen. Solubility of gases increases with pressure raise, and this is the reason for aeroembolism (caisson disease; the bends): solubility of nitrogen in the blood increases very sharply at high pressure, but when the conditions return to normal, excess of gas forms bubbles and blood-vessel occlusion follows.

The *concentration* is a measure of the amount of solute in a given amount of solution or solvent. A solution is *unsaturated* when more solute will dissolve in it. It is *saturated* when no more solute will dissolve (e.g., when solid sugar is added to a saturated solution of sugar, it falls to the bottom and no more seems to dissolve even if you stir it intensively). Thus, the term 'saturated' means 'full', whereas 'unsaturated' means that 'the solution could hold more'. There are also *supersaturated* solutions, but they are quite unstable and we need some special conditions to prepare them. A *concentrated* solution contains a large amount of solute per given amount of solvent or solution. A *dilute* solution contains a small amount of solute per given amount of solvent or solution (compare 50% and 1%  $\text{H}_2\text{SO}_4$ ).

The qualitative terms dilute versus concentrated or unsaturated versus saturated are descriptively useful. However, quantitative measures

of concentration are needed for many purposes in chemical laboratories, hospitals and pharmacies. For example, the effective administration of medicine usually requires a prescribed amount of the therapeutic agent.

**Concentration** can be expressed in various ways.

**1. Percent by mass,  $C_{\%}$ ,** means the number of grams of solute in 100 gr of solution:

$$C_{\%} = (m_{\text{solute}} / m_{\text{solution}}) \cdot 100\%. \quad (1.1)$$

A 10% NaCl solution may contain 10 grams of NaCl (and 90 grams of water) in 100 grams of solution or any other ratio of NaCl to solution for which the mass of NaCl is 10% of the total mass, for example, 20 milligrams of NaCl in 200 mg of the solution, 1.5 g of NaCl in 15 g of solution, or 7.4 kg of NaCl in 74 kg of solution. When using percent concentration by mass in calculations, it is convenient to use grams of solute per 100 grams of solution, since the number of grams of solute in 100 g of the solution is equal to the percent by mass of solute. To calculate the mass of solute in a given **volume** of solution, you must know the specific gravity (or **density**) of the solution (**d**). This is the mass of 1 ml of the solution:

$$m_{\text{solution}} = d \cdot V. \quad (1.2)$$

In case when you need *percent by volume* you are supposed to use «V» instead of «m» in formula (1.1) given above.

**2. Molarity,  $C_M$ ,** is defined as the number of moles of solute per 1 L of solution (1 L = 1 dm<sup>3</sup> = 1000 mL).

In accordance with this definition,  $C_M = \nu/V$ , where the number of moles  $\nu = m/M$ , and the volume V is put in liters. Finally, we get formula:

$$C_M = m / (M \cdot V), \quad (1.3)$$

where m — the mass of substance, gr; M — molar mass of substance, gr/mol; V — the volume of solution, L. In order to prepare the solution with a certain molar concentration, it is very convenient to use the numerator of formula (1.3):

$$m = C_M \cdot M \cdot V. \quad (1.4)$$

For example, if we need 100 ml of 1 M (1 mol/liter) NaCl solution, we should dissolve  $m(\text{NaCl}) = C_M \cdot M \cdot V = 1 \cdot 58.5 \cdot 0.1 = 5.85$  (gr), because

$$M_r(\text{NaCl}) = A_r(\text{Na}) + A_r(\text{Cl}) = 23 + 35.5 = 58.5,$$

and thus

$$M(\text{NaCl}) = 58.5 \text{ gr/mol.}$$

*Do not forget: if you want to get mass in grams, you take the volume in litres!*

Technically, you weigh 5.85 g of NaCl, put in the 100 mL flask using the funnel, and add enough distilled water to reach the mark level of the flask.

**3. Normality**,  $C_N$  is the number of equivalents of solute per 1 L of solution ( $1 \text{ L} = 1 \text{ dm}^3 = 1000 \text{ mL}$ ).

Similar to the formulae (1.3, 1.4) we have:

$$C_N = m / (\text{Eq} \cdot V), \quad (1.5)$$

$$m = C_N \cdot \text{Eq} \cdot V, \quad (1.6)$$

so we just use equivalent instead of molar mass.

*Note. And again, to calculate mass in grams, you take the volume in liters!*

An **equivalent** (Eq, or equivalent weight) of various reagents depends on the type of reaction, but for our purposes it is enough to know that the equivalent weight of acid is calculated by dividing the molecular weight ( $M_r$ ) of the acid by the number of hydrogen atoms ('H'):

$$\text{Eq}(\text{HCl}) = M(\text{HCl})/1 = 36.5 \text{ g/mol};$$

$$\text{Eq}(\text{H}_2\text{SO}_4) = M(\text{H}_2\text{SO}_4)/2 = 98 / 2 = 49 \text{ g/mol.}$$

The equivalent weight of a base is calculated by dividing the molecular weight of the base by the number of hydroxyl ('OH') groups:

$$\text{Eq}(\text{NaOH}) = M(\text{NaOH})/1 = 40 / 1 = 40 \text{ g/mol};$$

$$\text{Eq}(\text{Ca}(\text{OH})_2) = M(\text{Ca}(\text{OH})_2) / 2 = 74 / 2 = 37 \text{ g/mol.}$$

The equivalent weight of salt is calculated by dividing the molecular weight of the salt by the product of the number of the metal atoms and the oxidation number of this metal atom:

$$\overset{+3}{\text{Eq}(\text{Al}_2(\text{SO}_4)_3)} = M(\text{Al}_2(\text{SO}_4)_3) / (2 \cdot 3) = 342 / 6 = 57 \text{ g/mol.}$$

To prepare a 1 N solution ( $C_N = 1 \text{ N}$ , or one normal solution), it is necessary to add sufficient solvent to 1 equivalent of the solute to make a total volume of a solution of exactly 1 L. For example, a 1 N  $\text{Al}_2(\text{SO}_4)_3$  solution is prepared by dissolving 57 g of  $\text{Al}_2(\text{SO}_4)_3$  in a sufficient amount of water to produce 1 L of solution, because

$$m(\text{Al}_2(\text{SO}_4)_3) = C_N \cdot \text{Eq} \cdot V = 1 \cdot 57 \cdot 1 = 57 \text{ g.}$$

Of course, if we need only 100 ml of that solution, we should dissolve 5.7 g of  $\text{Al}_2(\text{SO}_4)_3$  in water to give 0.1 L of solution ( $m = C_N \cdot \text{Eq} \cdot V = 1 \cdot 57 \cdot 0.1 = 5.7 \text{ g}$ ).

There are some substances, for which  $\text{Eq} = M$  (division by 1), and consequently  $C_N = C_M$  (examples include acids like  $\text{HClO}_4$ ,  $\text{HNO}_3$ , etc.; bases like  $\text{KOH}$ ,  $\text{NaOH}$ , etc.; salts like  $\text{AgNO}_3$ ,  $\text{NaCl}$ , etc.). In other cases, we use the formulae (1.4, 1.6) to find a relationship between molarity and normality.

We get:  $m = C_N \cdot \text{Eq} \cdot V = C_M \cdot M \cdot V$  then cancel the volume:

$$C_N \cdot \text{Eq} = C_M \cdot M, \quad (1.7)$$

$\Rightarrow$  molarity  $C_M = C_N \cdot \text{Eq} / M$  or normality  $C_N = C_M \cdot M / \text{Eq}$ .

*Note. Sometimes we name normality "molar concentration of equivalent".*

**4. Molality,  $C_m$ ,** shows the number of moles of solute dissolved in 1 kg (1000 g) of the solvent.

$$C_m = \nu_{\text{solute}} / m_{\text{solvent}} = m_{\text{solute}} / (M \cdot m_{\text{solvent}}), \quad (1.8)$$

where  $\nu_{\text{solute}}$  — the number of moles of solute;  $m_{\text{solute}}$  — the mass of solute, gr;  $m_{\text{solvent}}$  — the mass of solvent, kg;  $M$  — the molar mass of solute, gr/mol.

For example, 0.2m solution of  $\text{H}_2\text{SO}_4$  means that there are 0.2 moles of  $\text{H}_2\text{SO}_4$  in 1000 g of water.

**5. Titer,  $T$ ,** is defined as the mass of solute (grams) per 1 ml (1  $\text{cm}^3$ ) of solution.

$$T = m / V, \quad (1.9)$$

where  $m$  — the mass of solute, gr;  $V$  — the volume of solution, mL.

In chemical analysis we use the following formula:

$$T = Eq \cdot C_N / 1000, \quad (1.10)$$

where  $Eq$  — the equivalent weight of solute, gr/mol;  $C_N$  — normality, N; 1000 — the coefficient for units translation (1 L = 1000 mL).

It is possible to calculate percent concentration if we know titer and consider that the solution is dilute, so its density  $d = 1$  g/ml:  $C\% = T \cdot 100$

**6. Mole fraction,  $X$ .** This is the ratio of the number of given substance moles to the total number of moles of all the substances, present in the solution. For the solution which contains only two components (solute “st”, and solvent “sv”)

$$X_{st} = n_{st} / (n_{st} + n_{sv}), \quad (1.11)$$

$$X_{sv} = n_{sv} / (n_{st} + n_{sv}), \quad (1.12)$$

where  $X_{st}$  — the mole fraction of solute,  $X_{sv}$  — the mole fraction of solvent,  $n_{st}$  — the number of solute moles,  $n_{sv}$  — the number of solvent moles.

$$\text{In sum, } X_{st} + X_{sv} = 1. \quad (1.13)$$

*Note. Percent by mass,  $C\%$ , is approximate concentration, the other ones (especially  $C_M$ ,  $C_N$ , and  $T$ ) are exact and thus useful in all types of analysis.*

### **Ionic force of strong electrolyte solutions and their mixtures (preparation of plasma substitutes)**

In a solution of a strong electrolyte, as well as in a mixture of strong electrolytes (like plasma substitutes), ionic force ( $I$ ) is used to characterize the electrostatic field created by ions. To calculate  $I$ , we sum up the molar concentration of each ion, multiplied by the squared charge, and find half of this sum:

$$I = 0.5 \cdot \sum_{(i)} C_i \cdot z_i^2 \quad (1.14)$$

As the ions organize the pairs and other units, their active concentration decreases, in comparison with the analytical one (for example, electrical conduction will be less than expected because the ions are “busy”). The real concentration (taking into consideration the interaction between ions) is called “activity ( $a_i$ )” of the certain type of ions and is mostly used in scientific researches.

For medical purposes, we use a lot of plasma substitutes. Calculating the ionic force of such mixtures, we are supposed to compare it with the ionic force of blood.

### I (blood serum) = 0.15 M

In the physiologic saline solution which is 0.9% NaCl,  $I = 0.154$  M. Ringer's solution, Hemodesum or other plasma substitutes (*example 8 with calculations is given below*) are more complicated by their composition and contain lots of different ions (so-called "polarizing mixture", Russian doctors often use a slang word "polyarka"). These solutions are very often used for clinical treatment.

## CONTROL QUESTIONS

1. The definition of the solution. What is the size of particles in a true solution? What kind of particles are there?
2. Give some examples of the solutions. Name solvent and solute.
3. Compare the mixtures (like Fe+S powder), aqueous solutions and compounds. What is the difference between them?
4. What is hydrate / crystalline hydrate? Examples.
5. How to distinguish a saturated solution from unsaturated?
6. What is solubility? Factors, which have influence on solubility.
7. Describe the mechanism of aeroembolism development.
8. What is percent by mass? Write the formula of  $C_{\%}$ . How many grams of  $\text{CaCl}_2$  do we need in order to prepare 200 gr of 10% solution?
9. What is molarity? Write the formula of  $C_M$ . Find the mass (gr) of  $\text{MgSO}_4$  ( $M = 120$  gr/mol) in 1 L of 0.1 M solution.
10. What is normality? Write the formula of  $C_N$ . Find the mass (gr) of  $\text{MgSO}_4$  ( $Mr = 120$ ) in 1 L of 0.1 N solution.
11. What is molality? Write the formula of  $C_m$ . What is the difference between molarity and molality? Find molality if 18 gr of glucose ( $M = 180$  gr/mol) are dissolved in 1 kg of water.
12. What is titer? Write the formula of T. Find this concentration if  $C_N(\text{NaCl}) = 0,1$  N.
13. What is the ionic force (I)? How to calculate the value of I?

14. Why do we calculate ionic force of plasma substitutes? What is the normal value of I for human blood?
15. Give some examples of plasma substitutes.

## EXAMPLES OF CALCULATIONS

### Example 1.

0.9% NaCl (saline solution) is called “physiological solution”. It is a popular plasma-substitute and solvent of drugs. How to prepare 200 mL of this solution (consider density  $d = 1 \text{ g/mL}$ )? Find its molarity (mol/L and mmol/L).

#### **Answer.**

*In accordance with definition 0.9% NaCl solution contains:*

*0.9 gr (NaCl) — in 100 gr (solution)  $\approx$  100 mL (solution),*

*x gr (NaCl) — in 200 mL.*

*Finding x, we get 1.8 gr.*

*Working in laboratory, we weigh 1.8 gr of NaCl, put in 200 mL flask, add distilled water. Solution for intravenous feeding must be sterilized.*

*There are different ways of  $C_M$  calculations. Let's use the formula (1.3), given above, in theory:*

$$C_M = 0.9 \text{ gr} / (58.5 \text{ gr/mol} \cdot 0.1 \text{ L}) = 0.15 \text{ mol/L} = 150 \text{ mmol/L},$$

*the same may be written like  $C_M = 0.15 \text{ M} = 150 \text{ mM}$*

*Note. Remember both concentrations ( $C_{\%}$  and  $C_M$ ) of physiological saline solution, it is necessary for your future studies.*

**So, 0.9% NaCl solution has molarity 0.15 M.**

### Example 2.

5%  $C_6H_{12}O_6$  (glucose solution) is also very often used for intravenous feeding and as a solvent of drugs. How to prepare 1 L of this solution ( $d = 1.06 \text{ g/mL}$ )? Find its molarity (mol/L and mmol/L).

#### **Answer.**

*Using the formulae (1.1) and (1.2), we have:*

$$m_{\text{solute}} = C_{\%} \cdot m_{\text{solution}} / 100\% = C_{\%} \cdot d \cdot V / 100 = 5\% \cdot 1.06 \text{ g/mL} \cdot 1000 \text{ mL} / 100\% = 53.0 \text{ g}.$$

*Working in laboratory, we weigh 53 gr of glucose, put in 1 L flask, add distilled water. Solution for intravenous feeding must be sterilized.*

*Of course, you can use a proportion like 5 gr (glucose) — in 100 gr (solution), then your actions will be similar to what we have in example 1.*

*Transforming this proportion, it is possible to find molarity as well.*

*Left part: 5 gr (glucose). How many moles?  $v = m/M = 5 / 180 = 0.028$  mole.*

*Right part: 100 gr (solution). What is it in liters?  $V = m_{\text{solution}} / d = 100 / 1.06 = 94.3$  mL.*

*Finally, we want to find how many moles in 1 L of solution in accordance with definition.*

*0.028 moles — in 94.3 mL,*

*x moles — in 1000 mL.*

*Finding x, we get 0.3 moles:*

$$C_M = 0.3 \text{ mol/L} = 300 \text{ mmol/L.}$$

*The same may be written like  $C_M = 0.3 \text{ M} = 300 \text{ mM}$ .*

*Note. Remember both concentrations ( $C_{\%}$  and  $C_M$ ), it is necessary for your future studies. So, **5% glucose solution has molarity 0.3 M.***

### **Example 3.**

Find  $C_M$  of 10%  $\text{CaCl}_2$ , which is used for intravenous injections. Consider  $d = 1.1$  g/mL.

**Answer.**

*10 gr ( $\text{CaCl}_2$ ) — in 100 gr (solution),*

↓

↓

$$v = 10 / 111 = \quad V = 100 / 1.1 = \\ = 0.09 \text{ moles} \quad = 90.9 \text{ mL,}$$

*0.09 moles — in 90.9 mL,*

*x moles — in 1000 mL.*

*$x = 0.99$  moles,  $C_M = 0.99 \text{ M} \approx 1 \text{ M}$ .*

### **Example 4.**

Find  $C_N$  and T of a 0.1 M  $\text{H}_2\text{SO}_4$ .

**Answer.**

*We know that  $\text{Eq}(\text{H}_2\text{SO}_4) = M/2$ . Let's put it in formula (1.7) and cancel "M":*

$$C_N \cdot M/2 = C_M \cdot M,$$

$$C_N = 2 \cdot C_M = 2 \cdot 0.1 = 0.2 \text{ (N).}$$

*Note. Remember that the value of  $C_N$  is always greater than  $C_M$ , excepting the cases when  $C_N = C_M$  (read theory).*

$$T = \text{Eq} \cdot C_N / 1000 = 49 \cdot 0.2 / 1000 = 0.0098 \text{ g/mL.}$$