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AL- MUSTAQBAL UNIVERSITY

College Of Health And Medical Techniques

Prosthetic Dental Techniques Department

Second Grade

Second Semester

Advanced chemistry

Lecture 4 (The theoretical part)

(Solutions and Their Concentrations)

By:

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2. **Solutions and Their Concentrations**

2.1.The Mole

The **mole** (abbreviated mol) is the SI unit for the amount of a chemical substance. It is always associated with specific microscopic entities such as atoms, molecules, ions, electrons, other particles, or specified groups of such particles as represented by a chemical formula. It is the amount of the specified substance that contains the same number of particles as the number of carbon atoms in exactly 12 grams of 12C. This important number is Avogadro's number *N*A = 6.022 * 10^{23}. The **molar mass** *M* of a substance is the mass in grams of 1 mole of that substance. We calculate molar masses by summing the atomic masses of all the atoms appearing in a chemical formula. For example, the molar mass of formaldehyde CH2O is

$$\mathcal{M}_{CH_2O} = \frac{1\ \cancel{\text{mol C}}}{\text{mol } CH_2O} \times \frac{12.0\ \text{g}}{\cancel{\text{mol C}}} + \frac{2\ \cancel{\text{mol H}}}{\text{mol } CH_2O} \times \frac{1.0\ \text{g}}{\cancel{\text{mol H}}}$$
$$+ \frac{1\ \cancel{\text{mol O}}}{\text{mol } CH_2O} \times \frac{16.0\ \text{g}}{\cancel{\text{mol O}}}$$
$$= 30.0\ \text{g/mol } CH_2O$$

and that of glucose, $C_6H_{12}O_6$, is

$$\mathcal{M}_{C_6H_{12}O_6} = \frac{6\ \cancel{\text{mol C}}}{\text{mol } C_6H_{12}O_6} \times \frac{12.0\ \text{g}}{\cancel{\text{mol C}}} + \frac{12\ \cancel{\text{mol H}}}{\text{mol } C_6H_{12}O_6} \times \frac{1.0\ \text{g}}{\cancel{\text{mol H}}}$$
$$+ \frac{6\ \cancel{\text{mol O}}}{\text{mol } C_6H_{12}O_6} \times \frac{16.0\ \text{g}}{\cancel{\text{mol O}}} = 180.0\ \text{g/mol } C_6H_{12}O_6$$

Thus, 1 mole of formaldehyde has a mass of 30.0 g, and 1 mole of glucose has a mass of 180.0 g.

2.2. The Millimole

Sometimes it is more convenient to make calculations with millimoles (mmol) rather than moles. The millimole is 1/1000 of a mole, and the mass in grams of a millimole, the millimolar mass (m*M*), is likewise 1/1000 of the molar mass.

2.3. Calculating the Amount of a Substance in Moles or Millimoles

The two examples that follow illustrate how the number of moles or millimoles of a species can be determined from its mass in grams or from the mass of a chemically related species .

EXAMPLE 4-1

Find the number of moles and millimoles of benzoic acid ($\mathcal{M}$ = 122.1 g/mol) that are contained in 2.00 g of the pure acid.

Solution

If we use HBz to represent benzoic acid, we can write that 1 mole of HBz has a mass of 122.1 g. Therefore,

$$\text{amount HBz} = n_{HBz} = 2.00\ \text{g}\ \cancel{\text{HBz}} \times \frac{1\ \text{mol HBz}}{122.1\ \text{g}\ \cancel{\text{HBz}}} \tag{4-1}$$

$$= 0.0164\ \text{mol HBz}$$

To obtain the number of millimoles, we divide by the millimolar mass (0.1221 g/mmol), that is,

$$\text{amount HBz} = 2.00\ \text{g}\ \cancel{\text{HBz}} \times \frac{1\ \text{mmol HBz}}{0.1221\ \text{g}\ \cancel{\text{HBz}}} = 16.4\ \text{mmol HBz}$$

EXAMPLE 4-2

What is the mass in grams of Na^+ (22.99 g/mol) in 25.0 g of Na_2SO_4 (142.0 g/mol)?

Solution

The chemical formula tells us that 1 mole of Na_2SO_4 contains 2 moles of Na^+, that is,

$$\text{amount Na}^+ = n_{Na^+} = \cancel{\text{mol Na}_2\text{SO}_4} \times \frac{2\ \text{mol Na}^+}{\cancel{\text{mol Na}_2\text{SO}_4}}$$

To find the number of moles of Na_2SO_4, we proceed as in Example 4-1:

$$\text{amount Na}_2\text{SO}_4 = n_{Na_2SO_4} = 25.0\ \text{g}\ \cancel{\text{Na}_2\text{SO}_4} \times \frac{1\ \text{mol Na}_2\text{SO}_4}{142.0\ \text{g}\ \cancel{\text{Na}_2\text{SO}_4}}$$

Combining this equation with the first leads to

$$\text{amount Na}^+ = n_{Na^+} = 25.0\ \text{g}\ \cancel{\text{Na}_2\text{SO}_4} \times \frac{1\ \cancel{\text{mol Na}_2\text{SO}_4}}{142.0\ \text{g}\ \cancel{\text{Na}_2\text{SO}_4}} \times \frac{2\ \text{mol Na}^+}{\cancel{\text{mol Na}_2\text{SO}_4}}$$

To obtain the mass of sodium in 25.0 g of Na_2SO_4, we multiply the number of moles of Na^+ by the molar mass of Na^+, or 22.99 g. And so,

$$\text{mass Na}^+ = \cancel{\text{mol Na}^+} \times \frac{22.99\ \text{g Na}^+}{\cancel{\text{mol Na}^+}}$$

Substituting the previous equation gives the mass in grams of Na^+:

$$\text{mass Na}^+ = 25.0\ \text{g}\ \cancel{\text{Na}_2\text{SO}_4} \times \frac{1\ \cancel{\text{mol Na}_2\text{SO}_4}}{142.0\ \text{g}\ \cancel{\text{Na}_2\text{SO}_4}} \times \frac{2\ \cancel{\text{mol Na}^+}}{\cancel{\text{mol Na}_2\text{SO}_4}} \times \frac{22.99\ \text{g Na}^+}{\cancel{\text{mol Na}^+}}$$

$$= 8.10\ \text{g Na}^+$$

2.4. Concentration of Solutions

we describe the four fundamental ways of expressing solution concentration: molar concentration, percent concentration, solution-diluent volume ratio, and p-functions.

2.4.1. *Molar Concentration*

The **molar concentration** c_X of a solution of a solute species X is the number of moles of that species that is contained in 1 liter of the solution (*not 1 L of the solvent*). In terms of the number of moles of solute, *n*, and the volume, *V*, of solution, we write

$$c_x = \frac{n_X}{V}$$

$$\text{molar concentration} = \frac{\text{no. moles solute}}{\text{volume in liters}}$$

The unit of molar concentration is **molar**, symbolized by **M**, which has the dimensions of mol/L, or mol L^{-1}. Molar concentration is also the number of millimoles of solute per milliliter of solution.

$$1\ \text{M} = 1\ \text{mol L}^{-1} = 1\ \frac{\text{mol}}{\text{L}} = 1\ \text{mmol L}^{-1} = 1\ \frac{\text{mmol}}{\text{L}}$$

EXAMPLE 4-3

Calculate the molar concentration of ethanol in an aqueous solution that contains 2.30 g of C_2H_5OH (46.07 g/mol) in 3.50 L of solution.

Solution

To calculate molar concentration, we must find both the amount of ethanol and the volume of the solution. The volume is given as 3.50 L, so all we need to do is convert the mass of ethanol to the corresponding amount of ethanol in moles.

$$\text{amount } C_2H_5OH = n_{C_2H_5OH} = 2.30\ \text{g}\ \cancel{C_2H_5OH} \times \frac{1\ \text{mol}\ C_2H_5OH}{46.07\ \text{g}\ \cancel{C_2H_5OH}}$$
$$= 0.04992\ \text{mol}\ C_2H_5OH$$

To obtain the molar concentration, $c_{C_2H_5OH}$, we divide the amount by the volume. Thus,

$$c_{C_2H_5OH} = \frac{2.30\ \text{g}\ \cancel{C_2H_5OH} \times \dfrac{1\ \text{mol}\ C_2H_5OH}{46.07\ \text{g}\ \cancel{C_2H_5OH}}}{3.50\ \text{L}}$$
$$= 0.0143\ \text{mol}\ C_2H_5OH/\text{L} = 0.0143\ \text{M}$$

We will see that there are two ways of expressing molar concentration: molar analytical concentration and molar equilibrium concentration. The distinction between these two expressions is in whether the solute undergoes chemical change in the solution process.

1- **Molar analytical concentration**, or for the sake of brevity, just **analytical concentration**, of a solution gives the *total* number of moles of a solute in 1 liter of the solution (or the total number of millimoles in 1 mL).
2- The **molar equilibrium concentration**, or just **equilibrium concentration**, refers to the molar concentration of a *particular species* in a solution at equilibrium.

EXAMPLE 4-4

Calculate the analytical and equilibrium molar concentrations of the solute species in an aqueous solution that contains 285 mg of trichloroacetic acid, Cl_3CCOOH (163.4 g/mol), in 10.0 mL (the acid is 73% ionized in water).

Solution

As in Example 4-3, we calculate the number of moles of Cl_3CCOOH, which we designate as HA, and divide by the volume of the solution, 10.0 mL, or 0.0100 L. Therefore,

$$\text{amount HA} = n_{HA} = 285 \text{ mg} \cancel{\text{HA}} \times \frac{1 \text{ g} \cancel{\text{HA}}}{1000 \text{ mg} \cancel{\text{HA}}} \times \frac{1 \text{ mol HA}}{163.4 \text{ g} \cancel{\text{HA}}}$$
$$= 1.744 \times 10^{-3} \text{ mol HA}$$

The molar analytical concentration, c_{HA}, is then

$$c_{HA} = \frac{1.744 \times 10^{-3} \text{ mol HA}}{10.0 \cancel{\text{mL}}} \times \frac{1000 \cancel{\text{mL}}}{1 \text{ L}} = 0.174 \frac{\text{mol HA}}{\text{L}} = 0.174 \text{ M}$$

In this solution, 73% of the HA dissociates, giving H^+ and A^-:

$$HA \rightleftharpoons H^+ + A^-$$

The equilibrium concentration of HA is then 27% of c_{HA}. Thus,

$$[HA] = c_{HA} \times (100 - 73)/100 = 0.174 \times 0.27 = 0.047 \text{ mol/L}$$
$$= 0.047 \text{ M}$$

The equilibrium concentration of A^- is equal to 73% of the analytical concentration of HA, that is,

$$[A^-] = \frac{73 \text{ mol } A^-}{100 \cancel{\text{mol HA}}} \times 0.174 \frac{\cancel{\text{mol HA}}}{\text{L}} = 0.127 \text{ M}$$

(continued)

Because 1 mole of H^+ is formed for each mole of A^-, we can also write

$$[H^+] = [A^-] = 0.127 \text{ M}$$

and

$$c_{HA} = [HA] + [A^-] = 0.047 + 0.127 = 0.174 \text{ M}$$

EXAMPLE 4-5

Describe the preparation of 2.00 L of 0.108 M $BaCl_2$ from $BaCl_2 \cdot 2H_2O$ (244.3 g/mol).

Solution

To determine the number of grams of solute to be dissolved and diluted to 2.00 L, we note that 1 mole of the dihydrate yields 1 mole of $BaCl_2$. Therefore, to produce this solution we will need

$$2.00\ \cancel{L} \times \frac{0.108\ \text{mol}\ BaCl_2 \cdot 2H_2O}{\cancel{L}} = 0.216\ \text{mol}\ BaCl_2 \cdot 2H_2O$$

The mass of $BaCl_2 \cdot 2H_2O$ is then

$$0.216\ \cancel{\text{mol}\ BaCl_2 \cdot 2H_2O} \times \frac{244.3\ \text{g}\ BaCl_2 \cdot 2H_2O}{\cancel{\text{mol}\ BaCl_2 \cdot 2H_2O}} = 52.8\ \text{g}\ BaCl_2 \cdot 2H_2O$$

Dissolve 52.8 g of $BaCl_2 \cdot 2H_2O$ in water and dilute to 2.00 L.

EXAMPLE 4-6

Describe the preparation of 500 mL of 0.0740 M Cl^- solution from solid $BaCl_2 \cdot 2H_2O$ (244.3 g/mol).

Solution

$$\text{mass}\ BaCl_2 \cdot 2H_2O = \frac{0.0740\ \cancel{\text{mol}\ Cl}}{\cancel{L}} \times 0.500\ \cancel{L} \times \frac{1\ \cancel{\text{mol}\ BaCl_2 \cdot 2H_2O}}{2\ \cancel{\text{mol}\ Cl}}$$

$$\times \frac{244.3\ \text{g}\ BaCl_2 \cdot 2H_2O}{\cancel{\text{mol}\ BaCl_2 \cdot 2H_2O}} = 4.52\ \text{g}\ BaCl_2 \cdot 2H_2O$$

Dissolve 4.52 g of $BaCl_2 \cdot 2H_2O$ in water and dilute to 0.500 L or 500 mL.

2.4.2.Percent Concentration

Chemists frequently express concentrations in terms of percent (parts per hundred). Unfortunately, this practice can be a source of ambiguity because percent composition of a solution can be expressed in several ways. Three common methods are

$$\text{weight percent (w/w)} = \frac{\text{weight solute}}{\text{weight solution}} \times 100\%$$

$$\text{volume percent (v/v)} = \frac{\text{volume solute}}{\text{volume solution}} \times 100\%$$

$$\text{weight/volume percent (w/v)} = \frac{\text{weight solute, g}}{\text{volume solution, mL}} \times 100\%$$

2.4.2.1.Parts per Million and Parts per Billion

For very dilute solutions, **parts per million** (ppm) is a convenient way to express concentration:

$$c_{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 \text{ ppm}$$

Example :

What is the molar concentration of K1 in a solution that contains 63.3 ppm of K3Fe(CN)6 (329.3 g/mol)?

Solution

Because the solution is so dilute, it is reasonable to assume that its density is 1.00 g/mL. Therefore, according to Equation 4-2,

$$63.3 \text{ ppm } K_3Fe(CN)_6 = 63.3 \text{ mg } K_3Fe(CN)_6/L$$

$$\frac{\text{no. mol } K_3Fe(CN)_6}{L} = \frac{63.3 \cancel{\text{ mg } K_3Fe(CN)_6}}{L} \times \frac{1 \cancel{\text{ g } K_3Fe(CN)_6}}{1000 \cancel{\text{ mg } K_3Fe(CN)_6}}$$

$$\times \frac{1 \text{ mol } K_3Fe(CN)_6}{329.3 \cancel{\text{ g } K_3Fe(CN)_6}} = 1.922 \times 10^{-4} \frac{\text{mol}}{L}$$

$$= 1.922 \times 10^{-4} \text{ M}$$

$$[K^+] = \frac{1.922 \times 10^{-4} \cancel{\text{ mol } K_3Fe(CN)_6}}{L} \times \frac{3 \text{ mol } K^+}{1 \cancel{\text{ mol } K_3Fe(CN)_6}}$$

$$= 5.77 \times 10^{-4} \frac{\text{mol } K^+}{L} = 5.77 \times 10^{-4} \text{ M}$$

2.4.2.2.p-Functions

Scientists frequently express the concentration of a species in terms of its **p-function**, or **p-value**. The p-value is the negative logarithm (to the base 10) of the molar concentration of that species. Thus, for the species X,

$$\mathbf{pX = - \log [X]}$$

EXAMPLE 4-8

Calculate the p-value for each ion in a solution that is 2.00×10^{-3} M in NaCl and 5.4×10^{-4} M in HCl.

Solution

$$pH = -\log [H^+] = -\log (5.4 \times 10^{-4}) = 3.27$$

To obtain pNa, we write

$$pNa = -\log[Na^+] = -\log (2.00 \times 10^{-3}) = -\log (2.00 \times 10^{-3}) = 2.699$$

The total Cl^- concentration is given by the sum of the concentrations of the two solutes:

$$[Cl^-] = 2.00 \times 10^{-3}\ M + 5.4 \times 10^{-4}\ M$$
$$= 2.00 \times 10^{-3}\ M + 0.54 \times 10^{-3}\ M = 2.54 \times 10^{-3}\ M$$
$$pCl = -\log[Cl^-] = -\log 2.54 \times 10^{-3} = 2.595$$

Note that in Example 4-8, and in the one that follows, the results are rounded according to the rules listed on page 117.

EXAMPLE 4-9

Calculate the molar concentration of Ag^+ in a solution that has a pAg of 6.372.

Solution

$$pAg = -\log [Ag^+] = 6.372$$
$$\log [Ag^+] = -6.372$$
$$[Ag^+] = 4.246 \times 10^{-7} \approx 4.25 \times 10^{-7}\ M$$

2.4.2.3. Density and Specific Gravity of Solutions

Density and specific gravity are related terms often found in the analytical literature. The **density** of a substance is its mass per unit volume, and its **specific gravity** is the ratio of its mass to the mass of an equal volume of water at 4°C. Density has units of kilograms per liter or grams per milliliter in the metric system. Specific gravity is dimensionless and so is not tied to any particular system of units. For this reason, specific gravity is widely used in describing items of commerce(see **Figure 10**).Since the density of water is approximately 1.00 g/mL and since we use the metric system throughout this text, we use density and specific gravity interchangeably. The specific gravities of some concentrated acids and bases are given in **Table 1**.

EXAMPLE 4-10

Calculate the molar concentration of HNO_3 (63.0 g/mol) in a solution that has a specific gravity of 1.42 and is 70.5% HNO_3 (w/w).

Solution

Let us first calculate the mass of acid per liter of concentrated solution

$$\frac{\text{g HNO}_3}{\text{L reagent}} = \frac{1.42\ \cancel{\text{kg reagent}}}{\text{L reagent}} \times \frac{10^3\ \cancel{\text{g reagent}}}{\cancel{\text{kg reagent}}} \times \frac{70.5\ \text{g HNO}_3}{100\ \cancel{\text{g reagent}}} = \frac{1001\ \text{g HNO}_3}{\text{L reagent}}$$

Then,

$$c_{\text{HNO}_3} = \frac{1001\ \cancel{\text{g HNO}_3}}{\text{L reagent}} \times \frac{1\ \text{mol HNO}_3}{63.0\ \cancel{\text{g HNO}_3}} = \frac{15.9\ \text{mol HNO}_3}{\text{L reagent}} \approx 16\ \text{M}$$

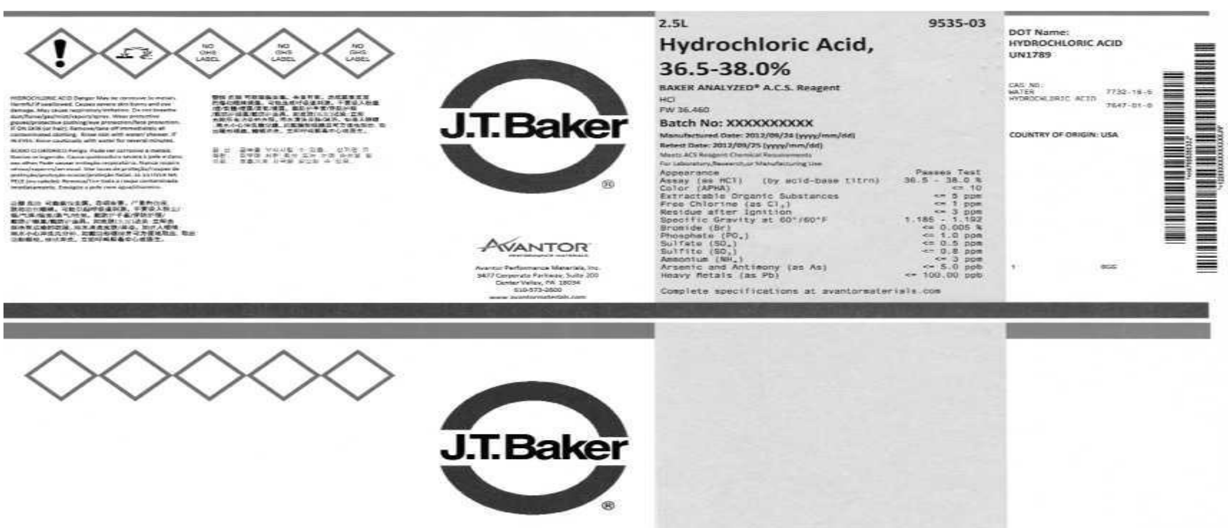

Figure 10 Label from a bottle of reagent-grade hydrochloric acid. Note that the specific gravity of the acid over the temperature range of 60° to 80°F is specified on the label.

TABLE 4-3

Specific Gravities of Commercial Concentrated Acids and Bases

Reagent	Concentration, % (w/w)	Specific Gravity
Acetic acid	99.7	1.05
Ammonia	29.0	0.90
Hydrochloric acid	37.2	1.19
Hydrofluoric acid	49.5	1.15
Nitric acid	70.5	1.42
Perchloric acid	71.0	1.67
Phosphoric acid	86.0	1.71
Sulfuric acid	96.5	1.84

EXAMPLE 4-11

Describe the preparation of 100 mL of 6.0 M HCl from a concentrated solution that has a specific gravity of 1.18 and is 37% (w/w) HCl (36.5 g/mol).

Solution

Proceeding as in Example 4-10, we first calculate the molar concentration of the concentrated reagent. We then calculate the number of moles of acid that we need for the diluted solution. Finally, we divide the second figure by the first to obtain the volume of concentrated acid required. Thus, to obtain the concentration of the reagent, we write

$$c_{HCl} = \frac{1.18 \times 10^3 \cancel{\text{g reagent}}}{\text{L reagent}} \times \frac{37 \cancel{\text{g HCl}}}{100 \cancel{\text{g reagent}}} \times \frac{1 \text{ mol HCl}}{36.5 \cancel{\text{g HCl}}} = 12.0 \text{ M}$$

The number of moles HCl required is given by

$$\text{no. mol HCl} = 100 \cancel{\text{mL}} \times \frac{1 \cancel{\text{L}}}{1000 \cancel{\text{mL}}} \times \frac{6.0 \text{ mol HCl}}{\cancel{\text{L}}} = 0.600 \text{ mol HCl}$$

Finally, to obtain the volume of concentrated reagent, we write

$$\text{vol concd reagent} = 0.600 \cancel{\text{mol HCl}} \times \frac{1 \text{ L reagent}}{12.0 \cancel{\text{mol HCl}}} = 0.0500 \text{ L or } 50.0 \text{ mL}$$

Therefore, dilute 50 mL of the concentrated reagent to 600 mL.

The solution to above Example is based on the following useful relationship, which we will be using countless times:

$$V_{concd} * c_{concd} = V_{dil} * c_{dil}$$

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Lecture (2)

(المحاليل ووحدات التركيز)

(Solutions and Their Units of Concentrations)

- **The Mole.**
- **The Millimole.**
- **Concentration of Solutions.**
- **Molar Concentration.**
- **Percent Concentration.**
- **p-Functions.**
- **Density and Specific Gravity of Solutions.**

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