SOLUTIONS-2
SOLUTION CONCENTRATIONS

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SOLUTION CONCENTRATIONS

- There are a number of ways one can measure concentration of solute in solutions.
  - Molarity – commonly used since volumes of solutions are easy to measure
  - Mole fraction – used for gases and vapor pressures of solutions
  - Molality – temperature independent
  - Percent by Mass – temperature independent and need not know molar masses
  - Conversion between units requires the use of density if any mass to volume or volume to mass conversion in needed.

VARIOUS UNITS OF SOLUTIONS

- Molarity (mol/L)

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

- Mass percent (mass/mass; mass/vol; vol/vol – the denominator is solution)

$$\text{Mass percentage of solute} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$$

- Molality (mol/Kg)

$$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

- Mole fraction

$$X = \frac{\text{moles of solute}}{\text{total moles of solution}}$$

- Parts per million

$$ppm = \frac{\text{mass* of solute}}{\text{mass* of solution}} \text{x}10^6$$

- Parts per billion

$$ppb = \frac{\text{mass* of solute}}{\text{mass* of solution}} \text{x}10^9$$

EXAMPLE: MASS PERCENT

An experiment calls for 36.0 g of a 5.00% aqueous solution of potassium bromide. Describe how you would make up such a solution.

Solution:

A 5.00% aqueous solution of KBr has 5.00 g KBr per 100. g solution. The remainder of the 100. g is water: 95 g.

We can use this ratio to determine the mass of KBr in 36.0 g solution:

$$36.0\ \text{g solution} \times \frac{5.00\ \text{g KBr}}{100.\ \text{g solution}} = 1.80\ \text{g KBr}$$

Since 1.80 g KBr is required for 36.0 g of solution, the remainder consists of 34.2 g water.

We make the solution by mixing 1.8 g KBr in 34.2 g water.

EXAMPLE: MOLALITY

Iodine dissolves in a variety of organic solvents. For example, in methylene chloride, it forms an orange solution. What is the molality of a solution of 5.00 g iodine, I_2, in 30.0 g of methylene chloride, CH_2Cl_2?

Solution:

Mass of solute = 5.00 g I_2

Mass of solvent = 30.0 g CH_2Cl_2

Calculate mols of the solute and convert the mass of solvent into Kg.

$$m = \frac{5.00\ \cancel{g}\ I_2}{30.0\ \cancel{g}\ \text{solvent}} \times \frac{1\,\text{mol}\ I_2}{253.8\ \cancel{g}\ I_2} \times \frac{10^3\ \cancel{g}}{1\,\text{kg}} = 0.657\,mol/Kg$$

EXAMPLE: MOL FRACTION

A solution of iodine, I_2, in methylene chloride, CH_2Cl_2, contains 5.00 g I_2 and 56.0 g CH_2Cl_2. What is the mole fraction of each component in this solution?

Solution:

Mass of solute = 5.00 g I_2

Mass of solvent = 56.0 g CH_2Cl_2

Find moles of both; add the moles and then calculate mol fraction of each.

$$\text{Moles solute} = 5.00\ \text{g}\ I_2 \times \frac{1\ \text{mol}\ I_2}{253.8\ \text{g}\ I_2} = 0.01970\ \text{mol}$$

$$\text{Moles solvent} = 56.0\ \text{g}\ I_2 \times \frac{1\ \text{mol}\ CH_2Cl_2}{84.93\ \text{g}\ CH_2Cl_2} = 0.6594\ \text{mol}$$

$$\text{Total moles} = 0.01970\ \text{mol} + 0.6594\ \text{mol} = 0.6791\ \text{mol}$$

$$X_{I_2} = \frac{0.01970\ \text{mol}\ I_2}{0.6791\ \text{mol total}} = 0.0290$$

$$X_{CH_2Cl_2} = \frac{0.6594\ \text{mol}\ CH_2Cl_2}{0.6791\ \text{mol total}} = 0.971$$

EXAMPLE: UNIT CONVERSIONS; MASS % TO MOL FRACTION

A bottle of bourbon is labeled 94 proof, meaning that it is 47% by volume of alcohol in water. What is the mole fraction of ethyl alcohol, C_2H_5OH, in the bourbon? The density of ethyl alcohol is 0.80 g/mL.

Solution:

When no volume is given – assume it is 1L. So there is 470 mL alcohol and 530 mL water. Find masses of both; convert to mols and find mol fractions.

$$470\,\text{mL}\,C_2H_5OH \times \frac{0.80\,\text{g}}{1\,\text{mL}} \times \frac{1\,\text{mol}}{46.08\,\text{g}} = 8.16\,\text{mol}$$

$$530\,\text{mL}\,H_2O \times \frac{1.00\,\text{g}}{1\,\text{mL}} \times \frac{1\,\text{mol}}{18.02\,\text{g}} = 29.4\,\text{mol}$$

$$\text{Total moles} = 8.16\,\text{mol} + 29.4\,\text{mol} = 37.6\,\text{mol}$$

$$X_{\text{ethanol}} = \frac{8.16\,\text{mol}\,C_2H_5OH}{37.6\,\text{mol total}} = 0.22$$

EXAMPLE: UNIT CONVERSION; MOLALITY TO MOLARITY

Citric acid, $HC_6H_7O_7$, is often used in fruit beverages to add tartness. An aqueous solution of citric acid is 2.331 *m* $HC_6H_7O_7$. What is the molarity of the solution? The density of the solution is 1.1346 g/mL.

Solution:

2.331 *m* means – 2.331 mols of citric acid in 1.000 Kg ***SOLVENT***; molarity is mols over L of ***SOLUTION***.

Strategy: convert mols citric acid to mass -> mass of solution -> use d to find L of solution.

$$\text{Mass solute} = 2.331\,\text{mol} \times \frac{192.14\,\text{g}}{1\,\text{mol}} = 447.88\,\text{g}$$

$$\text{Mass of solution} = 447.88\,\text{g} + 1000.00\,\text{g} = 1447.88\,\text{g}$$

$$\text{Liters solution} = 1447.88\,\text{g} \times \frac{1\,\text{mL}}{1.1346\,\text{g}} \times \frac{10^{-3}\,\text{L}}{\text{mL}} = 1.267L$$

$$M = \frac{2.331\,\text{mol}}{1.2761\,\text{L}} = 1.827\,\frac{\text{mol}}{\text{L}}$$

KEY CONCEPTS

- Know all the units of solutions
- Convert one unit of concentration to another.