Solutions & Dilutions

Basic concepts and formula

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Moles

- Number of moles of a substances contained in a given quantity (*e.g.* in 25 g NaCl)

$$Number\ of\ moles = \frac{weight\ in\ grams}{weight\ of\ one\ mole} = \frac{25g}{58.4g \cdot mol^{-1}} = 0.43 moles\ NaCl$$

Molarity of a solution

- Definition for *e.g.* 1M NaCl

$$1M = \frac{1\ mole\ NaCL}{1L\ solution}$$

If you need to prepare a certain volume of solution at a given concentration, you need to know the number of moles required.

For instance, if you need to prepare 200 mL of a 0.25 M solution of NaCl, how many grams of NaCl do you need?

$$\frac{0.25moles\ NaCl}{1L} \cdot 0.2L = 0.05moles\ NaCl \cdot \frac{58.4g\ NaCl}{mole\ NaCl} = 2.92g\ NaCl$$

Molarity of a solution

- On the contrary: If you want to know the final molar concentration *for e.g.* when you add 0.33 moles of a substance in a volume of 0.5, which molarity do we obtain?

$$c = \frac{0.33 moles}{0.5L} = \frac{0.67 moles}{L} = 0.67M$$

Solutions

- **% Solutions**: the amount (weight or volume) of a solute is expressed as a percentage of the total solution weight or volume.

Easy to be calculated, because they do not depend on knowledge of the molecular weight.

- **%w/v** is <u>percent weight to volume</u> with units g/100 mL.

Therefore a 1% w/v solution has 1 g of solute in a total of 100 ml of solution.

- **%v/v** is <u>percent volume to volume</u> with units mL/100 mL.

Therefore a 1% v/v solution of ethanol has 1 ml of pure ethanol in 100 mL of total solution.

Solutions

$$\frac{weight}{volume}\% = \frac{weight\ of\ solute}{volume\ of\ solution}\times 100 \qquad \text{(Eq. 1)}$$

$$\frac{weight}{weight}\% = \frac{weight\ of\ solute}{weight\ of\ solution}\times 100 \qquad \text{(Eq. 2)}$$

$$\frac{volume}{volume}\% = \frac{volume\ of\ solute}{volume\ of\ solution}\times 100 \qquad \text{(Eq. 3)}$$

Dilutions

- When doing a dilution, you always start with a sample of concentrated solution, called **stock solution**, and add additional **solvent** to it, thus lowering the concentration. The main formula for dilutions is:

$$C_1V_1 = C_2V_2$$

- Where C_1 and V_1 are the concentration and the volume, respectively, of the **stock solution**
- C_2 and V_2 are the concentration and the volume, respectively, of the **diluted final solution**.

pH and pK_a

- When an acid dissolves in water, at the equilibrium we have:

$$AH_{(aq)} + H_2O_{(l)} \rightleftarrows A^-_{(aq)} + H_3O^+_{(aq)}$$

- The position of the equilibrium is measured by the equilibrium constant

$$K_{eq} = \frac{[H_3O^+][A^-]}{[AH][H_2O]}$$

- In dilute solutions of acid, $[H_2O]$ stays roughly constant (about 56 mol dm^{-3}). We therefore define a new equilibrium constant, the acidity constant K_a:

$$K_a = \frac{[H_3O^+][A^-]}{[AH]}$$

pH and pK_a

- in logarithmic form is expressed as:

$$pK_a = -\log_{10} K_a$$

<u>The lower the pK_a, the higher the K_a and the stronger the acid!</u>

- It turns that the pK_a of an acid is the pH at which it is exactly half-dissociated. This can be shown by rearranging the expression for K_a:

$$[H_3O^+] = K_a \cdot \frac{[AH]}{[A^-]}$$

$$pH = pK_a - \log_{10} \frac{[AH]}{[A^-]}$$

$$\mathrm{pH} = -\log_{10}(a_{\mathrm{H}^+})$$

- Thus, when

$$[AH] = [A^-] \rightarrow pH = pK_a$$