

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)

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

Molarity of a solution

▲ Definition for *e.g.* 1M NaCl

$$1M = \frac{1 \text{ mole NaCl}}{1L \text{ solution}}$$

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?

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

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?

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

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 percent weight to volume 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 percent volume to volume 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

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

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

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

Dilutions

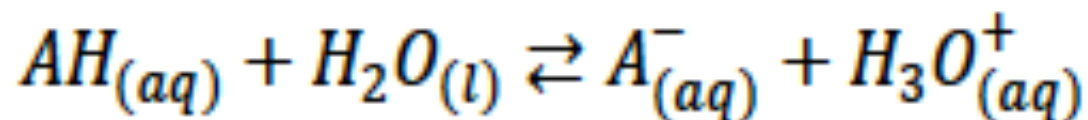
- ▲ When doing a dilution, you always start with a sample of concentrated solution, called **stock solution**, and add additional **solvent** to it, thus lowering the concentration. The main formula for dilutions is:

$$C_1V_1 = C_2V_2$$

- 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:



- ▲ 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:

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

The lower the pK_a , the higher the K_a and the stronger the acid!

- ▲ 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 :

$$[H_3O^+] = K_a \cdot \frac{[AH]}{[A^-]}$$
$$pH = pK_a - \log_{10} \frac{[AH]}{[A^-]}$$

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

- ▲ Thus, when

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