Solution Concentration

**solution**: homogeneous mixture of substances present as atoms, ions, and/or molecules

**solute**: component present in smaller amount

**solvent**: component present in greater amount

Note: Unless otherwise stated, the solvent for most solutions considered in this class will almost always be water!

**Aqueous solutions** are solutions in which water is the solvent.
A concentrated solution has a large quantity of solute present for a given amount of solution.

A dilute solution has a small quantity of solute present for a given amount of solution.

The more solute in a given amount of solution → the more concentrated the solution

Example: Explain the difference between the density of pure ethanol and the concentration of an ethanol solution.
How do we measure concentration?

Concentration can be measured a number of ways:

- **ppm** (parts per million) – one part in a million parts
- **ppb** (parts per billion) – one part in a billion parts
- **g/kg** (grams per kilogram) – one gram solute per one kilogram of solvent

The chemical standard most used is **Molarity**

\[
\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}
\]

units: \( M \) (molar = mol/L)
Ion Concentrations

• When an ionic compound is dissolved in water, the concentration on the individual ions is based on their molecular formula...

• For example:
  – 1 M NaCl solution contains 1 M Na$^+$ and 1 M Cl$^-$
  – 2 M NaCl solution contains 2 M Na$^+$ and 2 M Cl$^-$
  – 1 M CaCl$_2$ solutions contains 1 M Ca$^{2+}$ and 2 M Cl$^-$
  – 2 M CaCl$_2$ solutions contains 2 M Ca$^{2+}$ and 4 M Cl$^-$
Solving Concentration Problems

Keep in mind that if molarity and volume are both given, you can calculate # of moles since:

\[
\text{volume} \times \text{molarity} = \text{volume (in L)} \times \frac{\text{moles of solute}}{\text{liters of solution}}
\]

so volume units will cancel \(\rightarrow\) # of moles!

If you are given \textit{volume and molarity} for a solution, multiply them together to \textit{get # of moles}!
Preparing Solutions

How to prepare a 1.00 molar NaCl solution

First add 1.00 mole of NaCl.

Water

Add water until solid is dissolved. Then add additional water until the 1-liter mark is reached.

A 1.00 molar NaCl solution

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More concentrated standard solutions (with accurately known concentrations) are often diluted with deionized water to get a solution with a specific concentration.

1. A 100.0-mL volumetric flask has been filled to the mark with a 0.100 M $\text{K}_2\text{Cr}_2\text{O}_7$ solution.
2. This is transferred to a 1.000-L volumetric flask.
3. All of the initial solution is rinsed out of the 100.0-mL flask.
4. The 1.000-L flask is then filled with distilled water to the mark on the neck, and shaken thoroughly. The concentration of the now-diluted solution is 0.0100 M.
Dilutions

- A determined volume of a more concentrated solution is measured out in a small flask.
- The more concentrated solution is then transferred to a larger empty volumetric flask.
- The solution is diluted with deionized water to obtain the calibrated volume of the flask.

**Dilution Equation:** \[ M_1 V_1 = M_2 V_2 \]
\[ M_1 V_1 = M_2 V_2 \]

\( M_1 \) = initial molarity, \( V_1 \) = initial volume  
\( M_2 \) = final molarity, \( V_2 \) = final volume

The dilution equation is used to determine the new molarity of the dilute solution (\( M_2 \)) given the molarity of the more concentrated solution (\( M_1 \)) or the volume of the more concentrated solution to use (\( V_1 \)) given the new total volume of the dilute solution (\( V_2 \)) required.
Calculate the molarity of a hydrochloric acid solution prepared by diluting 15.0 mL of 6.00M hydrochloric acid to give a total volume of 100.0 mL.

\[ M_1 = 6.00 \text{ M} \quad V_1 = 15.0 \text{ mL} \]
\[ M_2 = ? \quad V_2 = 100.0 \text{ mL} \]

\[ M_1 V_1 = M_2 V_2 \]
\[ M_2 = \frac{M_1 V_1}{V_2} \]

\[ M_2 = \frac{(6.00 \text{ M}) \times (15.0 \text{ mL})}{100 \text{ mL}} \]

\[ = 0.900 \text{ M} \]
One important property of oxalic acid, $\text{H}_2\text{C}_2\text{O}_4(aq)$, is its ability to remove rust, $\text{Fe}_2\text{O}_3$, as shown in the following equation:

$$\text{Fe}_2\text{O}_3(s) + 6 \text{H}_2\text{C}_2\text{O}_4(aq) \rightarrow 2 \text{Fe(C}_2\text{O}_4)_3^{-3}(aq) + 3 \text{H}_2\text{O}(l) + 6 \text{H}^+(aq)$$

What volume of a 0.500M oxalic acid solution is required to remove 25.0 g of rust?

$\text{MW of } \text{Fe}_2\text{O}_3 = 159.69 \text{ g/mol}$

$25.0 \text{ g of } \text{Fe}_2\text{O}_3 = 0.157 \text{ mol}$

Need 6 mol of oxalic acid for 1 mol of $\text{Fe}_2\text{O}_3$

$0.0157 \text{ mol } \text{Fe}_2\text{O}_3 \times 6 = 0.939 \text{ mol oxalic acid}$

$0.939 \text{ mol}/0.500 \text{ M} = 1.88 \text{ L oxalic acid solution}$
Acid-Base Titrations

• Definitions:
  – **standard solution:** an acid or base solution where the concentration is known, generally to at least 3 or more sig figs
    • used to analyze properties of substances, such as the neutralizing power of commercial antacids, the tartness of wine, etc.
  – **acid-base indicators:**
    • Solutions that are pH sensitive and change color
    • Generally have color changes occurring for pH close to 7 since reactions monitored are neutralization reactions which occur near pH=7
Acid-Base Indicators

methyl red

bromthymol blue

phenolphthalein
• Definitions (cont’d)
  – **titration**: The gradual addition of a solution from a buret to another solution in a flask or beaker until the reaction between the two is complete, as signaled by the indicator changing color.
  – **titrant**: the solution in the buret
Acid-Base Titrations

• Definitions (cont’d)
  – **analyte**: the solution for which a property (e.g. molar concentration) is being determined

• In some titrations, the titrant is also the analyte, in which case a known amount of acid or base is present in the flask, and the amount of titrant necessary to neutralize it will be used to determine the concentration of the titrant.

• In other titrations, the titrant’s concentration is known, and the amount of titrant used to neutralize it will determine the concentration of the solution in the flask.
Acid-Base Titrations

• Definitions (cont’d)
  – **equivalence point**: Theoretical point in the titration when the amount of base added is exactly equal to the acid present, so the base completely neutralizes the acid.
  – **endpoint**: The moment when the acid-base indicator changes color.

• Note: in most acid-base titrations, the phenolphthalein indicator does not change color until the solution is basic, so in reality, there is a slight excess of hydroxide ion present in the solution when it turns pink.

• Ideally, the endpoint is reached with one drop or a fraction of a drop of titrant, so the endpoint is very close to the equivalence point.
Endpoint