

## Solutions: Definitions and Examples

### So, What is a Solution?

- A homogeneous mixture of two or more components
- “Mixture” implies non-stoichiometric, that is, a solution is “free” to have any composition (relative amounts of components) you desire!
- The most abundant component is the solvent (usually)
- The less abundant (or chemically more interesting) component(s) is(are) called the solute(s)
- The phase of the solution is determined by the phase of the solvent. When the solvent is H<sub>2</sub>O, the solute(s) is(are) said to be in the aqueous phase (*aq*)

### Putting a Number on Concentration: Solution Stoichiometry

#### 1. Mixing Ratio:

$$\frac{\text{mass of solute}}{\text{mass of solvent}} \times 10^n$$

The masses of solute and solvent must be measured in the same units

If  $n = 6$ , the mixing ratio is in parts per million (ppm)

If  $n = 9$ , the mixing ratio is in parts per billion (ppb)

If  $n = 12$ , the mixing ratio is in parts per trillion (ppt)

This is commonly used in environmental chemistry (*cf.* the EPA standard for ClO<sub>4</sub><sup>-</sup>), but only for dilute solutions. If the mixing ratio of a solute is greater than, say, 100 ppm, a chemist will usually employ another unit to quantify its concentration.

#### 2. Mass (or Weight) Percent: **(Know for Test 1)**

$$\frac{\text{mass of solute}}{\text{total mass of solution}} \times 100\% \quad \text{or} \quad \frac{\text{mass of solvent}}{\text{total mass of solution}} \times 100\%$$

This unit is especially common for concentrated solutions.

#### 3. Density: **(Know for Test 1)**

$$\rho = \frac{\text{mass of solution}}{\text{volume of solution}}$$

Mass of solution is usually measured in g, and volume of solution is usually measured in mL

The density of a solution will be very close to the density of pure solvent (*e.g.* 1.0 g/mL for H<sub>2</sub>O(*l*)) unless the solution is rather concentrated.

4. Molarity: (Know for Test 1)

$$[\text{solute}] = \frac{\text{mol of solute}}{\text{volume of solution in L}}$$

Note that molarity (abbreviated M and read as “molar”) must be in units of moles per liter!

Molarity is the most common way to measure concentration, since it is moles of solute that are most relevant in a chemical reaction, and it is convenient to measure out volumes of a solution accurately.

5. Number Density: (Know how to calculate for Test 1)

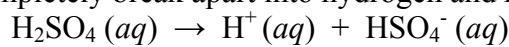
$$[\text{solute}] = \frac{\text{number of solute particles}}{\text{volume of solution}}$$

You do not need to memorize this definition, but you should know how to convert from a molarity (in mol/L) to a number density in units of (# of particles/mL). Number density is the most common concentration unit in atmospheric chemistry.

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Example: The concentrated sulfuric acid used in the sugar demonstration has a density of  $1.82 \text{ g mL}^{-1}$  and contains 5.0%  $\text{H}_2\text{O}$  by mass.

- Calculate the molarity of the sulfuric acid.
- An experiment requires 50 mL of 1.0 M  $\text{H}_2\text{SO}_4$ . Describe how you would prepare this solution using the given supply of concentrated  $\text{H}_2\text{SO}_4$ .
- As we will learn in greater detail in Chapter 18,  $\text{H}_2\text{SO}_4$  is a “strong” acid, which means that the  $\text{H}_2\text{SO}_4$  will completely break apart into hydrogen and hydrogen sulfate ions:



(The exception is in extremely concentrated solutions that contain more than ~25%  $\text{H}_2\text{SO}_4$  by mass.) Calculate the total number of ions in the solution prepared in part (b).