Describing Solutions

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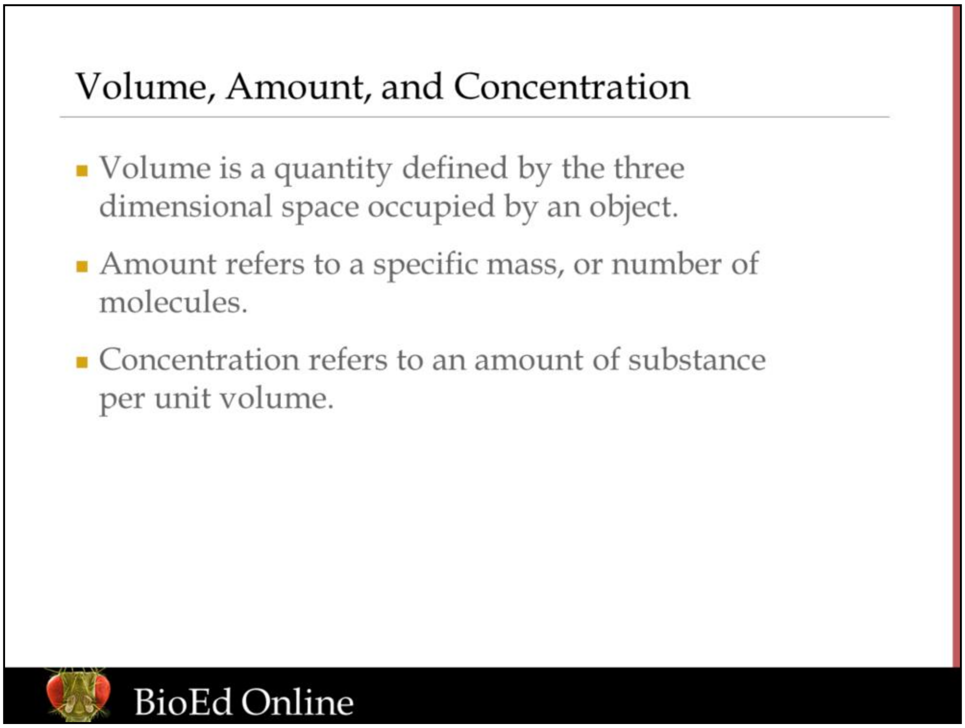

Volume, Amount, and Concentration

The common language of scientists uses units of measurement that are recognized in labs all over the world, appropriately called the International System of Units. The seven SI (for Systeme Internationale) base units are defined in terms of well established physical quantities or standards. Other units of measure were derived from the SI base units and are called SI-derived units. In solution making, we work with SI or SI-derived units for volumes, amounts, and concentrations.

Volume describes the space taken up by something. Since space is three-dimensional, a unit of volume is the cube of a unit of length, such as centimeters, feet, inches, or meters. An SI-derived unit for volume is the liter, defined as 1,000 cubic centimeters. One cubic centimeter, or one thousandth of a liter, is a milliliter. We use prefixes to simplify quantitative expressions of volume, amount, and concentration. For example, we might find it convenient to describe a volume in microliters (one millionth liter) or deciliters (1 tenth liter).

The SI base unit for an amount of substance is the mole. The mole and its use in defining solutions will be discussed later. A less specific unit with which to describe an amount of substance is a unit of mass. The SI base unit for mass is the kilogram, which is 1,000 grams. We may describe quantities using milligrams (one thousandth of a gram), micrograms (one millionth of a gram), or perhaps even nanograms (one billionth of a gram).

Concentration refers to the amount of substance within a specified volume. Solutions are defined by the type of solute or solutes, the type of solvent, the concentration of each solute, and often (for aqueous solutions), pH. Suspensions also can be defined as the amount of a minor component per unit total volume (volume of minor components plus volume of solvent). The concentration of a suspension can be described as the number of particles per unit volume, as often is the case for cell suspensions.

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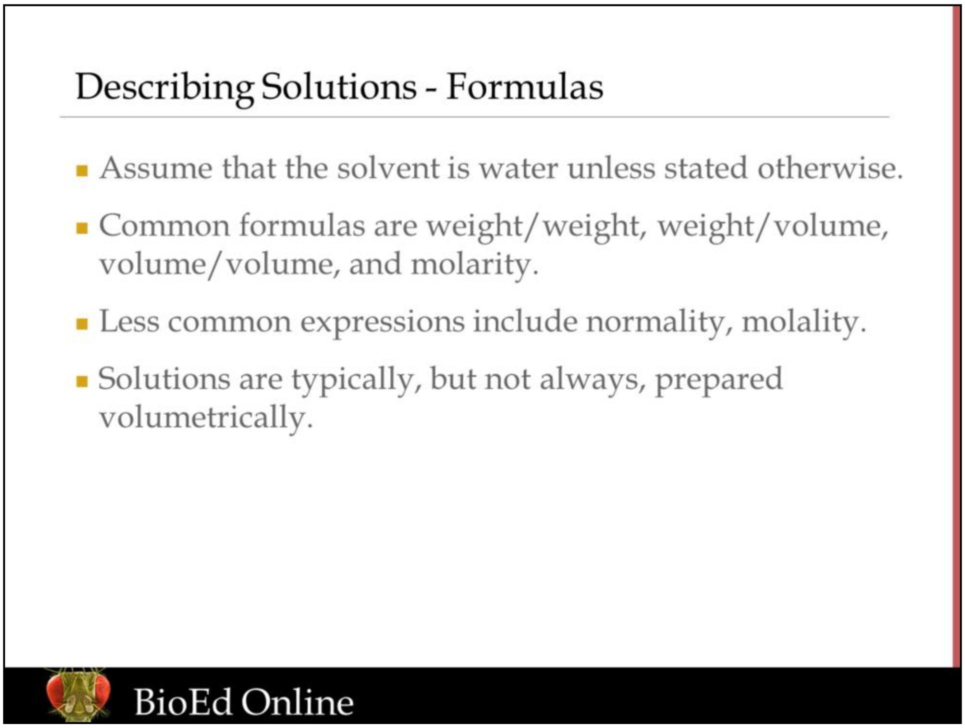

Describing Solutions - Formulas

Suppose that someone already has worked out the details, so all you have to do is read a formula and make a solution. We usually can assume that a solution is to be aqueous unless stated otherwise. What about the concentration of the substance to be added?

Common ways of describing the concentrations of solutions are weight-in-weight, weight-in-volume, volume-in-volume, and molarity. Less commonly used descriptions include normality and molality. We will go over them one at a time in part II, but first let's look at what they all have in common. A quantity of solute is measured out, mixed with solvent, and the volume is brought to some final value after the solute is completely dissolved. That is, solutions are typically prepared volumetrically. Because solutes add volume to a quantity of solvent, this method of preparation of solutions is necessary to ensure that an exact desired concentration is obtained.

There are exceptions, of course. For example, culture media for bacteria typically are made by adding a measured amount of powdered medium to a measured volume of water. In such cases, it isn't critical that a precise concentration be obtained, so a weight-to-volume method is appropriate, instead of weight-in-volume.

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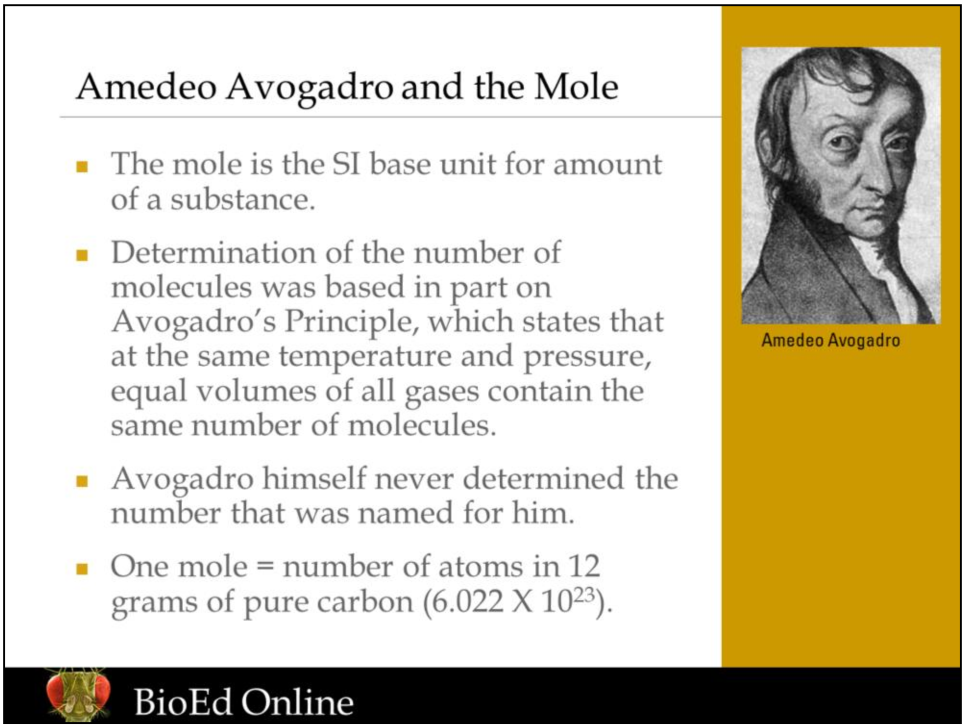

Amedeo Avogadro and the Mole

The masses of two different molecular species may be very different, so expressing a concentration as weight-to-volume or weight-to-weight tells us nothing about the proportions of actual reactants in a solution. Thus, it is therefore more useful to know the number of molecules of solute per unit volume than the mass of solute per unit volume. It would be extremely awkward to express concentrations as numbers of molecules per liter, because for a typical solution that number is staggeringly large. A convenient and physically meaningful unit for describing a quantity of molecules is the mole. The number of molecules in one mole of a pure substance, namely 6.022 x 10^{23}, is known as Avogadro's number, although Avogadro himself never calculated the number that bears his name.

Lorenzo Romano Amedeo Carlo Avogadro, conte di Quaregna e di Cerreto (1776 – 1856), was born in Turin, Italy. Avogadro's formal education was in law, and in fact, he had a successful legal career. In that era, many very basic scientific principles were as yet undiscovered, and it was possible for people such as Avogadro to pursue their interests in "natural philosophy" and actually make a lasting contribution to science.

In the early Nineteenth Century, John Dalton proposed that each atom of an element had a characteristic atomic weight, and that atoms were combined when chemical reactions took place. Around the same time, Gay-Lussac found that in chemical reactions involving gases, the ratios of volumes of the gases yielded small numbers that were integers not always equal to 1. If each substance was composed simply of a single atom, as Dalton had postulated, the ratios of volumes of reacting gases would have to be in unity.

In 1811, Avogadro made a distinction between atoms and molecules. For example, the "atoms" of nitrogen and oxygen are in reality "molecules, " each containing two atoms. Two molecules of hydrogen can combine with one molecule of oxygen to produce two molecules of water. Avogadro further suggested that at the same temperature and pressure, equal volumes of all gases contain the same number of molecules. This suggestion, which was borne out by later research, is known as Avogadro's principle.

Avogadro's number currently is based upon the definition of atomic mass, the atomic number for carbon, and the SI unit for mass. The atomic number for the common form of carbon is 12. A mole is defined as the number of atoms in 12 grams of pure carbon, which is 6.0221367 x 10^{23}, with some uncertainty about the seventh decimal place. One mole is an incredibly large number—counting at a pace of one per second it would take 20 million billion years to count the atoms in one mole.

References:

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