

2.3 Molar Mass - Counting Atoms by Weighing Them

Skills to Develop

After reading this section you should be able to

- Define *Avogadro's number* and explain why it is important to know.
- Define the *mole*.
- Be able to calculate the number of moles in a given mass of an element, or the mass corresponding to a given number of moles.
- Define *atomic weight* and *molar mass*; explain how the latter differs from the former.
- Be able to find the number of atoms in a given mass of a substance.

The chemical changes we observe always involve *discrete numbers of atoms* that rearrange themselves into new configurations. These numbers are HUGE—far too large in magnitude for us to count or even visualize, but they are still *numbers*, and we need to have a way to deal with them. We also need a bridge between these numbers, which we are unable to measure directly, and the weights of substances, which we do measure and observe. The **mole concept** provides this bridge, and is central to all of quantitative chemistry.

Counting Atoms: Avogadro's Number

Owing to their tiny size, atoms and molecules cannot be counted by direct observation. But much as we do when "counting" beans in a jar, we can estimate the number of particles in a sample of an element or compound if we have some idea of the volume occupied by each particle and the volume of the container. Once this has been done, we know the number of formula units (to use the most general term for any combination of atoms we wish to define) in any arbitrary weight of the substance. The number will of course depend both on the formula of the substance and on the mass of the sample. However, if we consider a mass of substance that is the same as its atomic or molecular mass, expressed in grams, we have only one number to know: **Avogadro's number**, abbreviated as N_A .

Avogadro's number

Avogadro's number is known to nine significant digits:

$$N_A = 6.02214076 \times 10^{23}. \quad (1)$$

However, for most calculations you only need to know it to four significant figures:

$$N_A \approx 6.022 \times 10^{23}. \quad (2)$$

So 6.022×10^{23} of *what*? Well, of anything you like: apples, stars in the sky, burritos. However, the only *practical* use for N_A is for expressing the huge numbers of the tiny particles such as atoms or molecules that we deal with in chemistry. Avogadro's number is a *counting number*, just like a dozen (12 things), or a gross (144 things), or a ream (500 things). Students can think of 6.022×10^{23} as the "chemist's dozen".

Before getting into the use of Avogadro's number in problems, take a moment to convince yourself of the reasoning embodied in the following examples.

Example 1: Mass ratio from atomic weights

The atomic weights of oxygen and carbon are 16.0 and 12.0 atomic mass units (*amu*), respectively. How much heavier is the oxygen atom in relation to carbon?

Solution

Atomic weights represent the relative masses of different kinds of atoms. This means that the atom of oxygen has a mass that is

$$\frac{16 \text{ amu}}{12 \text{ amu}} = \frac{4}{3} \approx 1.33$$

as great as the mass of a carbon atom.

Example 2: Mass of a single atom

The absolute mass of a carbon atom is 12.0 atomic mass units (*amu*). How many grams will a single oxygen atom weigh?

Solution

The absolute mass of a carbon atom is 12.0 *amu* or

$$12 \text{ } \cancel{\text{amu}} \times \frac{1.6605 \times 10^{-24} \text{ g}}{1 \text{ } \cancel{\text{amu}}} = 1.99 \times 10^{-23} \text{ g (per carbon atom)}$$

The mass of the oxygen atom will be 4/3 greater (from Example 1):

$$\left(\frac{4}{3}\right) 1.99 \times 10^{-23} \text{ g} = 2.66 \times 10^{-23} \text{ g (per oxygen atom)}$$

Alternatively we can do the calculation directly like with carbon:

$$16 \text{ } \cancel{\text{amu}} \times \frac{1.6605 \times 10^{-24} \text{ g}}{1 \text{ } \cancel{\text{amu}}} = 2.66 \times 10^{-23} \text{ g (per oxygen atom)}$$

Example 3: Relative masses from atomic weights

Suppose that we have N carbon atoms, where N is a number large enough to give us a pile of carbon atoms whose mass is 12.0 grams. How much would the same number, N , of oxygen atoms weigh?

Solution

We use the results from Example 1 again. The collection of N oxygen atoms would have a mass of

$$\frac{4}{3} \times 12 \text{ g} = 16.0 \text{ g}.$$

Exercise 1

What is the numerical value of N in Example 3?

Answer

Using the results of Examples 2 and 3.

$$N \times 1.99 \times 10^{-23} \text{ g (per carbon atom)} = 12 \text{ g}$$

or

$$N = \frac{12 \text{ } \cancel{\text{g}}}{1.99 \times 10^{-23} \text{ } \cancel{\text{g}} \text{ (per carbon atom)}} = 6.03 \times 10^{23} \text{ atoms}$$

There are a lot of atoms in 12 g of carbon.

Things to understand about Avogadro's number

- It is a *number*, just as is "dozen", and thus is *dimensionless*.
- It is a *huge* number, far greater in magnitude than we can visualize
- Its practical use is limited to counting tiny things like atoms, molecules, "formula units", electrons, or photons.
- The value of N_A can be known only to the precision that the number of atoms in a measurable weight of a substance can be estimated. Because large numbers of atoms cannot be counted directly, a variety of ingenious indirect measurements have been made involving such things as [Brownian motion](#) and [X-ray scattering](#).
- The current value was determined by measuring the distances between the atoms of silicon in an ultrapure crystal of this element that was shaped into a perfect sphere. (The measurement was made by X-ray scattering.) When combined with the measured mass of this sphere, it yields Avogadro's number. However, there are two problems with this:

- The silicon sphere is an artifact, rather than being something that occurs in nature, and thus may not be perfectly reproducible.
- The standard of mass, the kilogram, is not precisely known, and its value appears to be changing. For these reasons, there are proposals to revise the definitions of both N_A and the kilogram.

Moles and their Uses

The *mole* (abbreviated mol) is the the SI measure of *quantity of a "chemical entity"*, which can be an atom, molecule, formula unit, electron or photon. One mole of anything is just Avogadro's number of that something. Or, if you think like a lawyer, you might prefer the official SI definition:

Definition: The Mole

The mole is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.01200 kilogram of carbon 12

Avogadro's number (Equation 2) like any pure number, is dimensionless. However, it also defines the mole, so we can also express N_A as $6.022 \times 10^{23} \text{ mol}^{-1}$; in this form, it is properly known as **Avogadro's constant**. This construction emphasizes the role of Avogadro's number as a *conversion factor* between number of moles and number of "entities".

Example 4: number of moles in N particles

How many moles of nickel atoms are there in 80 nickel atoms?

Solution

$$80 \text{ Ni atoms} \times \frac{1 \text{ mole Ni atoms}}{6.022 \times 10^{23} \text{ Ni atoms}} = 1.33 \times 10^{-22} \text{ mole Ni atoms}$$

Is this answer reasonable? Yes, because 80 is an extremely small fraction of N_A .

Molar Mass

The mass of one mole of the fundamental units (atoms, molecules, or groups of atoms that correspond to the formula of a pure substance) is the ratio of its mass to 1/12 the mass of one mole of C-12 atoms, and being a ratio, is dimensionless. But at the same time, this *molar mass* (as many now prefer to call it) is also the observable mass of one mole (N_A) of the substance, so we frequently emphasize this by stating it explicitly as so many grams (or kilograms) per mole: g mol^{-1} .

It is important always to bear in mind that the mole is a number and not a mass. But each individual particle of a pure substance has a mass of its own, so a mole of any specific substance will always correspond to a certain mass, the molar mass, of that substance.

Example 5: Manganese atoms

The molar mass of manganese (Mn) is 54.938 grams per mole

- How many moles of manganese atoms are present in 20.0 g of manganese?
- How many atoms of manganese are present in 20.0 g of manganese?
- How many grams of manganese atoms are present in 3.47×10^{24} manganese atoms?

Solution

- $20.0 \text{ grams Mn atoms} \times \frac{1 \text{ mole Mn atoms}}{54.938 \text{ grams of Mn atoms}} = 0.364 \text{ moles of Mn atoms}$
- $20.0 \text{ grams Mn atoms} \times \frac{1 \text{ mole Mn atoms}}{54.938 \text{ grams of Mn atoms}} \times \frac{6.022 \times 10^{23} \text{ Mn atoms}}{1 \text{ mole Mn atoms}} = 2.19 \times 10^{23} \text{ Mn atoms}$
- $3.47 \times 10^{24} \text{ Mn atoms} \times \frac{1 \text{ mole Mn atoms}}{6.022 \times 10^{23} \text{ Mn atoms}} \times \frac{54.938 \text{ grams of Mn atoms}}{1 \text{ mole Mn atoms}} = 317 \text{ grams Mn atoms}$

Exercise 1

The molar mass of ruthenium (Ru) is 101.07 grams per mole

- How many moles of ruthenium atoms are present in 200.0 g of manganese?
- How many atoms of ruthenium are present in 200.0 g of manganese?
- How many grams of ruthenium atoms are present in 3.47×10^{24} manganese atoms?

Answer

$$200.0 \text{ grams Ru atoms} \times \frac{1 \text{ mole Ru atoms}}{101.07 \text{ grams of Ru atoms}} = 1.979 \text{ moles of Ru atoms}$$

$$200.0 \text{ grams Ru atoms} \times \frac{1 \text{ mole Ru atoms}}{101.07 \text{ grams of Ru atoms}} \times \frac{6.022 \times 10^{23} \text{ Ru atoms}}{1 \text{ mole Ru atoms}} = 1.192 \times 10^{24} \text{ Ru atoms}$$

$$3.47 \times 10^{24} \text{ Ru atoms} \times \frac{1 \text{ mole Ru atoms}}{6.022 \times 10^{23} \text{ Ru atoms}} \times \frac{101.07 \text{ grams of Ru atoms}}{1 \text{ mole Ru atoms}} = 582.4 \text{ grams Ru atoms}$$

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