3.3: Molar Mass

Mass is a basic physical property of matter. The mass of an atom or a molecule is referred to as the **atomic mass.** The atomic mass is used to find the average mass of elements and molecules and to solve stoichiometry problems.

Introduction

In chemistry, there are many different concepts of mass. It is often assumed that **atomic mass** is the mass of an atom indicated in **unified atomic mass units (u)**. However, the book *Quantities, Units and Symbols in Physical Chemistry* published by the IUPAC clearly states:

> "Neither the name of the physical quantity, nor the symbol used to denote it, should imply a particular choice of unit."

The name "atomic mass" is used for historical reasons, and originates from the fact that chemistry was the first science to investigate the same physical objects on macroscopic and microscopic levels. In addition, the situation is rendered more complicated by the isotopic distribution. On the macroscopic level, most mass measurements of pure substances refer to a mixture of isotopes. This means that from a physical stand point, these mixtures are not pure. For example, the macroscopic mass of oxygen (O_2) does not correspond to the microscopic mass of O_2. The former usually implies a certain isotopic distribution, whereas the latter usually refers to the most common isotope ($^{16}O_2$). Note that the former is now often referred to as the "molecular weight" or "atomic weight".

Mass Concepts in Chemistry

name in chemistry	physical meaning	symbol	units
atomic mass	mass on microscopic scale	m, m_a	Da, u, kg, g
molecular mass	mass of a molecule	m	Da, u, kg, g
isotopic mass	mass of a specific isotope		Da, u, kg, g
mass of entity	mass of a chemical formula	m, m_f	Da, u, kg, g
average mass	average mass of a isotopic distribution	m	Da, u, kg, g
molar mass	average mass per mol	$M = m/n$	kg/mol or g/mol
atomic weight	average mass of an element	$A_r = m / m_u$	unitless
molecular weight	average mass of a molecule	$M_r = m / m_u$	unitless
relative atomic mass	ratio of mass m and and the atomic mass constant m_u	$A_r = m / m_u$	unitless
atomic mass constant	$m_u = m(^{12}C)/12$	m_u = 1 Da = 1 u	Da, u, kg, g
relative molecular mass	ratio of mass m of a molecule and and the atomic mass constant m_u	$M_r = m / m_u$	unitless
relative molar mass	?	?	?
mass number	nucleon number	A	nucleons, or unitless
integer mass	nucleon number * Da	m	Da, u
nominal mass	integer mass of molecule consisting of most abundant isotopes	m	Da, u

name in chemistry	physical meaning	symbol	units
exact mass	mass of molecule calculated from the mass of its isotopes (in contrast of measured ba a mass spectrometer)		Da, u, kg, g
accurate mass	mass (not normal mass)		Da, u, kg, g

These concepts are further explained below.

Average Mass

Isotopes are atoms with the same atomic number, but different mass numbers. A different mass size is due to the difference in the number of neutrons that an atom contains. Although mass numbers are whole numbers, the actual masses of individual atoms are never whole numbers (except for carbon-12, by definition). This explains how lithium can have an atomic mass of 6.941 Da. The atomic masses on the periodic table take these isotopes into account, weighing them based on their abundance in nature; more weight is given to the isotopes that occur most frequently in nature. Average mass of the element E is defined as:

$$m(E) = \sum_{n=1} m(I_n) \times p(I_n) \tag{3.3.1}$$

where $\sum$ represents a n-times summation over all isotopes I_n of element E, and p(I) represents the relative abundance of the isotope I.

Example 1

Find the average atomic mass of boron using the Table 1 below:

Mass and abundance of Boron isotopes

n	isotope I_n		mass m (Da)	isotopic abundance p
1	^{10}B		10.013	0.199
2	^{11}B		11.009	0.801

Solution

The average mass of Boron is:

$$m(B) = (10.013\ Da)(0.199) + (11.009\ Da)(0.801) = 1.99\ Da + 8.82\ Da = 10.81\ Da \tag{3.3.2}$$

Relative Mass

Traditionally it was common practice in chemistry to avoid using any units when indicating atomic masses (e.g. masses on microscopic scale). Even today, it is common to hear a chemist say, "^{12}C has exactly mass 12". However, because mass is not a dimensionless quantity, it is clear that a mass indication needs a unit. Chemists have tried to rationalize the omission of a unit; the result is the concept of relative mass, which strictly speaking is not even a mass but a ratio of two masses. Rather than using a unit, these chemists claim to indicate the ratio of the mass they want to indicate and the atomic mass constant m_u which is defined analogous to the unit they want to avoid. Hence the relative atomic mass of the mass m is defined as:

$$A_r = \frac{m}{m_u} \tag{3.3.3}$$

The quantity is now dimensionless. As this unit is confusing and against the standards of modern metrology, the use of relative mass is discouraged.

Molecular Weight, Atomic Weight, Weight vs. Mass

Until recently, the concept of mass was not clearly distinguished from the concept of weight. In colloquial language this is still the case. Many people indicate their "weight" when they actually mean their mass. Mass is a fundamental property of objects, whereas weight is a force. Weight is the force F exerted on a mass m by a gravitational field. The exact definition of the weight is controversial. The weight of a person is different on ground than on a plane. Strictly speaking, weight even changes with location on earth.

When discussing atoms and molecules, the mass of a molecule is often referred to as the "molecular weight". There is no univerally-accepted definition of this term; however, mosts chemists agree that it means an average mass, and many consider it dimensionless. This would make "molecular weight" a synonym to "average relative mass".

Integer Mass

Because the proton and the neutron have similar mass, and the electron has a very small mass compared to the former, most molecules have a mass that is close to an integer value when measured in daltons. Therefore it is quite common to only indicate the **integer mass** of molecules. Integer mass is only meaningful when using dalton (or u) units.

Accurate Mass

Many mass spectrometers can determine the mass of a molecule with accuracy exceeding that of the integer mass. This measurement is therefore called the **accurate mass** of the molecule. Isotopes (and hence molecules) have atomic masses that are not integer masses due to a mass defect caused by binding energy in the nucleus.

Units

The atomic mass is usually measured in the units *unified atomic mass unit* (u), or dalton (Da). Both units are derived from the carbon-12 isotope, as 12 u is the exact atomic mass of that isotope. So 1 u is 1/12 of the mass of a carbon-12 isotope:

$$1 \text{ u} = 1 \text{ Da} = m(^{12}C)/12$$

The first scientists to measure atomic mass were John Dalton (between 1803 and 1805) and Jons Jacoband Berzelius (between 1808 and 1826). Early atomic mass theory was proposed by the English chemist William Prout in a series of published papers in 1815 and 1816. Known was Prout's Law, Prout suggested that the known elements had atomic weights that were whole number multiples of the atomic mass of hydrogen. Berzelius demonstrated that this is not always the case by showing that chlorine (Cl) has a mass of 35.45, which is not a whole number multiple of hydrogen's mass.

Some chemists use the atomic mass unit (amu). The amu was defined differently by physicists and by chemists:

- Physics: 1 amu = $m(^{16}O)/16$
- Chemistry: 1 amu = m(O)/16

Chemists used oxygen in the naturally occurring isotopic distribution as the reference. Because the isotopic distribution in nature can change, this definition is a moving target. Therefore, both communities agreed to the compromise of using $m(^{12}C)/12$ as the new unit, naming it the "unified atomic mass unit" (u). Hence, the amu is no longer in use; those who still use it do so with the definition of the u in mind. For this reason, the dalton (Da) is increasingly recommended as the accurate mass unit.

Neither u nor Da are SI units, but both are recognized by the SI.

Molar Mass

The molar mass is the mass of one mole of substance, whether the substance is an element or a compound. A mole of substance is equal to Avogadro's number (6.023×10^{23}) of that substance. The molar mass has units of g/mol or kg/mol. When using the unit g/mol, the numerical value of the molar mass of a molecule is the same as its average mass in daltons:

- Average mass of C: 12.011 Da
- Molar mass of C: 12.011 g/mol

This allows for a smooth transition from the microscopic world, where mass is measured in daltons, to the macroscopic world, where mass is measured in kilograms.

✔ Example 2

What is the molar mass of phenol, C_6H_5OH?

Average mass m = 6 × 12.011 Da + 6 × 1.008 Da + 1 × 15.999 Da = 94.113 Da

Molar mass = 94.113 g/mol = 0.094113 kg/mo

Measuring Masses in the Atomic Scale

Masses of atoms and molecules are measured by mass spectrometry. Mass spectrometry is a technique that measures the mass-to-charge ratio (m/q) of ions. It requires that all molecules and atoms to be measured be ionized. The ions are then separated in a mass analyzer according to their mass-to-charge ratio. The charge of the measured ion can then be determined, because it is a multiple of the elementary charge. The the ion's mass can be deduced. The average masses indicated in the periodic table are then calculated using the isotopic abundances, as explained above.

The masses of all isotopes have been measured with very high accuracy. Therefore, it is much simpler and more accurate to calculate the mass of a molecule of interest as a sum of its isotopes than measuring it with a commercial mass spectrometer.

Note that the same is not true on the nucleon scale. The mass of an isotope cannot be calculated accurately as the sum of its particles (given in the table below); this would ignore the mass defect caused by the binding energy of the nucleons, which is significant.

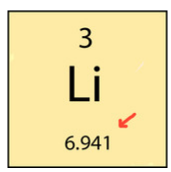
Table 2: Mass of three sub-atomic particles

Particle	SI (kg)	Atomic (Da)	Mass Number A
Proton	1.6726×10^{-27}	1.0073	1
Neutron	1.6749×10^{-27}	1.0087	1
Electron	9.1094×10^{-31}	0.00054858	0

As shown in Table 2, the mass of an electron is relatively small; it contributes less than 1/1000 to the overall mass of the atom.

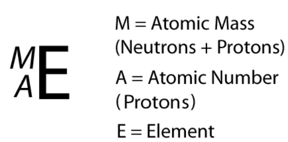
Where to Find Atomic Mass

The atomic mass found on the Periodic Table (below the element's name) is the average atomic mass. For example, for Lithium:



The red arrow indicates the atomic mass of lithium. As shown in Table 2 above and mathematically explained below, the masses of a protons and neutrons are about 1u. This, however, does not explain why lithium has an atomic mass of 6.941 Da where 6 Da is expected. This is true for all elements on the periodic table. The atomic mass for lithium is actually the average atomic mass of its isotopes. This is discussed further in the next section.

One particularly useful way of writing an isotope is as follows:



Applications

Applications Include:

1. Average Molecular Mass
2. Stoichiometry

Note: One particularly important relationship is illustrated by the fact that an atomic mass unit is equal to 1.66×10^{-24} g. This is the reciprocal of Avogadro's constant, and it is no coincidence:

$$\frac{\text{Atomic Mass (g)}}{1\text{g}} \times \frac{1\text{mol}}{6.022 \times 10^{23}} = \frac{\text{Mass (g)}}{1\text{atom}} \tag{3.3.4}$$

Because a mol can also be expressed as gram × atoms,

$$1\ u = \frac{M_u\ (molar\ mass\ unit)}{N_A\ (Avogadro's\ Number)} = 1\ \frac{g}{mol\ N_A} \tag{3.3.5}$$

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1u = Mu(molar mass unit)/NA(Avogadro's Number)=1g/mol/NA
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N_A known as Avogadro's number (Avogadro's constant) is equal to 6.023×10^{23} atoms.

Atomic mass is particularly important when dealing with stoichiometry.

Practice Problems

a. What is the molecular mass of radium bicarbonate, $Ra(HCO_3)_2$?
b. List the following, from least to greatest, in terms of their number of neutrons, and then atomic mass: ^{14}N, ^{42}Cl, ^{25}Na, ^{10}Be
c. A new element, Zenium, has 3 isotopes, ^{59}Ze, ^{61}Ze, and ^{67}Ze, with abundances of 62%, 27%, and 11% respectively. What is the atomic mass of Zenium?
d. An isotope with a mass number of 55 has 5 more neutrons than protons. What element is it?
e. How much mass does 3.71 moles of Fluorine have?
f. How many grams are there in 4.3×10^{22} molecules of $POCl_3$?
g. How many moles are there in 23 grams of sodium carbonate?

Solutions

a) Molecular mass of $Ra(HCO_3)_2$

= 226 + 2(1.01 u + 12.01 u + (16.00 u)(3)) = 348 u or g/mol

b) Number of neutrons: ^{10}Be, ^{14}N, ^{25}Na, ^{42}Cl

Atomic Mass: ^{10}Be, ^{14}N, ^{25}Na, ^{42}Cl

Note: It is the same increasing order for both number of neutrons and atomic mass because more neutrons means more mass.

c) Atomic mass of Zenium:

(59 u)(0.62) + (61 u)(0.27) + (67 u)(0.11)

= 37 u + 16 u + 7.4 u

= 60.4 u or g/mol

d) Mn

e) (3.71 moles F_2)(19 × 2 g/mol F_2)

= (3.71 mol F_2)(38 g/mol F_2)

= 140 g F_2

f) (4.3×10^{22} molecules $POCI_3$)(1 mol/6.022×10^{23} molecules POCI3)(30.97 + 16.00 + 35.45 x 3 g/mol $POCI_3$)

= (4.3×10^{22} molecules $POCI_3$)(mol/6.022×10^{23} molecules POCI3)(153.32 g/mol $POCI_3$)

= 11 g $POCI_3$

g) (23 g Na_2CO_3)(1 mol/22.99 × 2 + 12.01 + 16.00 × 3 g Na_2CO_3)

= (23 g Na_2CO_3)(1 mol/105.99 g Na_2CO_3)

= (0.22 mol Na_2CO_3)

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