

ATOMIC AND MOLECULAR MASS

Individual atoms and molecules, however, are very small, and the masses of individual atoms and molecules are also very small. For macroscopic objects, we use units such as grams and kilograms to state their masses, but these units are much too big to comfortably describe the masses of individual atoms and molecules. Another scale is needed.

The atomic mass unit (u; some texts use amu, but this older style is no longer accepted) is defined as one-twelfth of the mass of a carbon-12 atom, an isotope of carbon that has six protons and six neutrons in its nucleus. By this scale, the mass of a proton is 1.00728 u, the mass of a neutron is 1.00866 u, and the mass of an electron is 0.000549 u. There will not be much error if you estimate the mass of an atom by simply counting the total number of protons and neutrons in the nucleus (i.e., identify its mass number) and ignore the electrons. Thus, the mass of carbon-12 is about 12 u, the mass of oxygen-16 is about 16 u, and the mass of uranium-238 is about 238 u. More exact masses are found in scientific references—for example, the exact mass of uranium-238 is 238.050788 u, so you can see that we are not far off by using the whole-number value as the mass of the atom.

What is the mass of an element? This is somewhat more complicated because most elements exist as a mixture of isotopes, each of which has its own mass. Thus, although it is easy to speak of the mass of an atom, when talking about the mass of an element, we must take the isotopic mixture into account.

The atomic mass of an element is a weighted average of the masses of the isotopes that compose an element. What do we mean by a weighted average? Well, consider an element that consists of two isotopes, 50% with mass 10 u and 50% with mass 11 u. A weighted average is found by multiplying each mass by its fractional occurrence (in decimal form) and then adding all the products. The sum is the weighted average and serves as the formal atomic mass of the element. In this example, we have the following:

$0.50 \times 10 \text{ u}$	$= 5.0 \text{ u}$
$0.50 \times 11 \text{ u}$	$= 5.5 \text{ u}$
Sum	$= 10.5 \text{ u} = \text{the atomic mass of our element}$

Note that no atom in our hypothetical element has a mass of 10.5 u; rather, that is the average mass of the atoms, weighted by their percent occurrence.

This example is similar to a real element. Boron exists as about 20% boron-10 (five protons and five neutrons in the nuclei) and about 80% boron-11 (five protons and six neutrons in the nuclei). The atomic mass of boron is calculated similarly to what we did for our hypothetical example, but the percentages are different:

$0.20 \times 10 \text{ u}$	$= 2.0 \text{ u}$
$0.80 \times 11 \text{ u}$	$= 8.8 \text{ u}$
Sum	$= 10.8 \text{ u} = \text{the atomic mass of boron}$

Thus, we use 10.8 u for the atomic mass of boron.

Virtually all elements exist as mixtures of isotopes, so atomic masses may vary significantly from whole numbers. [Table 3.5 “Selected Atomic Masses of Some Elements”](#) lists the atomic masses of some elements; a more expansive table is in [Chapter 17 “Appendix: Periodic Table of the Elements”](#). The atomic masses in [Table 3.5 “Selected Atomic Masses of Some Elements”](#) are listed to three decimal places where possible, but in most cases, only one or two decimal places are needed. Note that many of the atomic masses, especially the larger ones, are not very close to whole numbers. This is, in part, the effect of an increasing number of isotopes as the atoms increase in size. (The record number is 10 isotopes for tin.)

Table 3.5 Selected Atomic Masses of Some Elements

Element Name	Atomic Mass (u)	Element Name	Atomic Mass (u)
Aluminum	26.981	Molybdenum	95.94
Argon	39.948	Neon	20.180
Arsenic	74.922	Nickel	58.693

Element Name	Atomic Mass (u)	Element Name	Atomic Mass (u)
Barium	137.327	Nitrogen	14.007
Beryllium	9.012	Oxygen	15.999
Bismuth	208.980	Palladium	106.42
Boron	10.811	Phosphorus	30.974
Bromine	79.904	Platinum	195.084
Calcium	40.078	Potassium	39.098
Carbon	12.011	Radium	n/a
Chlorine	35.453	Radon	n/a
Cobalt	58.933	Rubidium	85.468
Copper	63.546	Scandium	44.956

Element Name	Atomic Mass (u)		Element Name	Atomic Mass (u)
Fluorine	18.998		Selenium	78.96
Gallium	69.723		Silicon	28.086
Germanium	72.64		Silver	107.868
Gold	196.967		Sodium	22.990
Helium	4.003		Strontium	87.62
Hydrogen	1.008		Sulfur	32.065
Iodine	126.904		Tantalum	180.948
Iridium	192.217		Tin	118.710
Iron	55.845		Titanium	47.867
Krypton	83.798		Tungsten	183.84

Element Name	Atomic Mass (u)	Element Name	Atomic Mass (u)
Lead	207.2	Uranium	238.029
Lithium	6.941	Xenon	131.293
Magnesium	24.305	Zinc	65.409
Manganese	54.938	Zirconium	91.224
Mercury	200.59	Molybdenum	95.94

Note: Atomic mass is given to three decimal places, if known.

Now that we understand that atoms have mass, it is easy to extend the concept to the mass of molecules. The molecular mass is the sum of the masses of the atoms in a molecule. This may seem like a trivial extension of the concept, but it is important to count the number of each type of atom in the molecular formula. Also, although each atom in a molecule is a particular isotope, we use the weighted average, or atomic mass, for each atom in the molecule.

For example, if we were to determine the molecular mass of dinitrogen trioxide, N_2O_3 , we would need to add the atomic mass of nitrogen two times with the atomic mass of oxygen three times:

2 N masses = $2 \times 14.007 \text{ u}$	= 28.014 u
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3 O masses = $3 \times 15.999 \text{ u}$	= 47.997 u
Total	= 76.011 u = the molecular mass of N_2O_3

We would not be far off if we limited our numbers to one or even two decimal places.

Example 5

What is the molecular mass of each substance?

1. NBr_3
2. C_2H_6

Solution

1. Add one atomic mass of nitrogen and three atomic masses of bromine:

1 N mass	= 14.007 u
3 Br masses = $3 \times 79.904 \text{ u}$	= 239.712 u
Total	= 253.719 u = the molecular mass of NBr_3

2. Add two atomic masses of carbon and six atomic masses of hydrogen:

2 C masses = $2 \times 12.011 \text{ u}$	= 24.022 u
6 H masses = $6 \times 1.008 \text{ u}$	= 6.048 u

Total	= 30.070 u = the molecular mass of C ₂ H ₆
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3. The compound C₂H₆ also has a common name—ethane.

Test Yourself

What is the molecular mass of each substance?

1. SO₂

2. PF₃

Answers

1. 64.063 u

2. 87.968 u