

Isotopes and Average Atomic Mass

How do you calculate average atomic mass from isotopes?

●●○ Intermediate

Atomic Structure & Periodicity

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Before you start: it helps to understand [Inside the Atom: Protons, Neutrons and Electrons](#).

What is it?

Most elements exist naturally as a **mixture of isotopes**: atoms with the same number of protons but different numbers of neutrons, and therefore different masses.

The **atomic mass** printed on the periodic table is the **weighted average** mass of an element's atoms, taking into account how common each isotope is (its **natural abundance**).

KEY IDEA

The periodic-table atomic mass is an **average weighted by abundance**. That's why chlorine is 35.45 u and not a whole number: no single chlorine atom has a mass of 35.45 u.

Why does it matter?

- **Every calculation with molar mass uses it.** When you calculate moles from grams, you rely on average atomic masses, because real samples contain the natural mix of isotopes.
- **Isotopes are useful in their own right.** Carbon-14 is used to date ancient materials, and other isotopes are used in medical imaging and as tracers to follow chemical reactions.
- **It explains the periodic table.** Seeing that atomic masses are averages removes the mystery of the decimals.

How does it work?

1. Measuring isotopes

A **mass spectrometer** separates atoms (as ions) by mass. It measures the **mass** of each isotope and its **relative abundance**: what fraction of all the atoms it makes up.

2. Calculating the weighted average

Multiply each isotope's mass by its abundance written as a **decimal fraction** (75.76% → 0.7576), then add the results:

$$\text{average} = \sum (\text{fraction} \times \text{mass})$$

where the sum ($\sum$) runs over every isotope of the element.

The abundances of all the isotopes must add up to 100% (a total fraction of 1).

3. The average leans toward the most common isotope

Because chlorine-35 is about three times as common as chlorine-37, the average (35.45 u) is much closer to 35 than to 37.

∞ THINK OF IT LIKE THIS

Your final course grade might be an exam worth 75% and coursework worth 25%. If you score 35 on the exam and 37 on the coursework, your overall mark isn't the simple average (36). It's pulled toward the exam score because the exam counts more. Isotope abundances work the same way: they're the "weights".

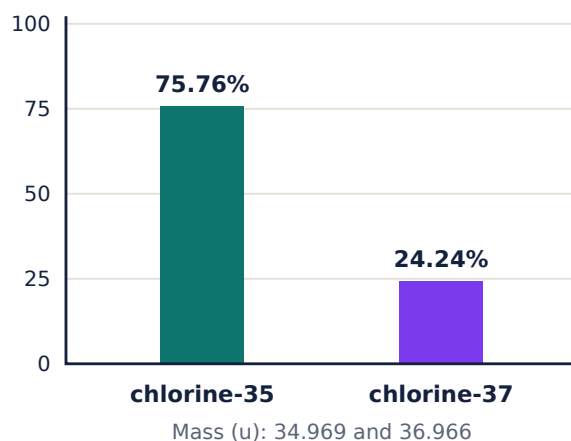
🔍 MORE PRECISELY

Chlorine exists naturally as Cl_2 molecules, so a real mass spectrum of chlorine gas also shows molecular peaks (around 70, 72 and 74 u). The diagram below is simplified to show single atoms. Isotope abundances also vary very slightly between natural sources, which is why IUPAC gives some standard atomic weights as a small range. Values such as $\text{Cl} = 35.45$ are conventional rounded values for everyday use.

Visualise it

Simplified mass spectrum of chlorine

Relative abundance of each isotope (%)



Weighted average: 35.45 u, the value on the periodic table. It is closer to 35 than to 37.

Chlorine-35 is about three times as common as chlorine-37, so the average sits near 35.

Worked example

WORKED EXAMPLE: Average atomic mass of chlorine

Question: Chlorine-35 (34.969 u) makes up 75.76% of chlorine atoms, and chlorine-37 (36.966 u) makes up 24.24%. Calculate the average atomic mass.

Step 1 Convert to fractions: 0.7576 and 0.2424 (they add up to 1).

Step 2 Multiply and add: $(0.7576 \times 34.969) + (0.2424 \times 36.966)$

Step 3 $= 26.493 + 8.961 = 35.45 \text{ u}$

Answer: 35.45 u, the value on the periodic table.

WORKED EXAMPLE: An element with three isotopes

Question: Magnesium consists of magnesium-24 (23.985 u, 78.99%), magnesium-25 (24.986 u, 10.00%) and magnesium-26 (25.983 u, 11.01%). Calculate its average atomic mass.

$$(0.7899 \times 23.985) + (0.1000 \times 24.986) + (0.1101 \times 25.983) = 24.31 \text{ u}$$

WORKED EXAMPLE: Finding an abundance from the average

Question: Gallium has two isotopes, gallium-69 (68.926 u) and gallium-71 (70.925 u). Its average atomic mass is 69.723 u. What percentage of gallium atoms are gallium-69?

Step 1 Let x = fraction of gallium-69, so gallium-71 is $1 - x$.

Step 2 $68.926x + 70.925(1 - x) = 69.723$

Step 3 $70.925 - 1.999x = 69.723$, so $x = \frac{1.202}{1.999} = 0.601$

Answer: about 60.1% of gallium atoms are gallium-69 (and 39.9% are gallium-71).

Common mistake

COMMON MISTAKE: Taking a simple average

Averaging 34.969 and 36.966 gives 35.97 u, which is wrong. The isotopes are not equally common, so each mass must be **weighted** by its abundance.

 COMMON MISTAKE: Using percentages instead of fractions

Multiplying by 75.76 instead of 0.7576 gives an answer 100 times too big. Convert percentages to decimals first (or divide your final sum by 100).

 COMMON MISTAKE: Confusing mass number with isotope mass

The **mass number** (35) is a whole-number count of protons and neutrons. The **isotope mass** (34.969 u) is the measured mass, which is slightly different. Use the isotope masses for accurate averages.

 NOTATION NOTE

- Periodic-table values are also called **relative atomic mass** (A_r , no units) or **atomic weight**, especially in UK courses and in IUPAC terminology.
- Abundance may be given as a percentage or as a fraction; check which before calculating.

Remember this

 REMEMBER THIS

- Atomic mass on the periodic table = **weighted average** of the isotope masses.
- Average = Σ (fraction $\times$ isotope mass), with fractions adding up to 1.
- The average is closest to the most abundant isotope.
- Never use a simple average unless the isotopes are equally common.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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