



Average Atomic Mass Calculator

Companion worksheet — use the interactive tool to complete each part.

Name: _____ Date: _____

Part A — Explore real elements

Using "Explore Real Elements" mode, choose 3 different elements. For each, record its isotope data, show the weighted-average calculation, and compare to the tool's answer.

Tip: The formula is: $(mass_1 \times abundance_1 \text{ as a decimal}) + (mass_2 \times abundance_2 \text{ as a decimal}) + \dots$ — remember to divide each percent by 100 first!

Element 1

Element chosen: _____

Mass number	Abundance (%)

Show your work:

My calculated average: _____ Actual atomic mass (from app): _____

Element 2

Element chosen: _____

Mass number	Abundance (%)

Show your work:

My calculated average: _____ Actual atomic mass (from app): _____



Element 3

Element chosen: _____

Mass number	Abundance (%)

Show your work:

My calculated average: _____ Actual atomic mass (from app): _____

Part A wrap-up

Which of your three elements' average atomic mass ended up closest to a whole number? What does that tell you about the abundance of its most common isotope?



Part B — Practice by hand

Calculate each average atomic mass on paper first. Then switch to "Build Your Own" mode and enter the same numbers to check yourself.

Problem 1

Element Q has two isotopes: Q-24 (60% abundance) and Q-26 (40% abundance).

Show your work:

Average atomic mass: _____ Matched app? Y / N: _____

Problem 2

Element Z has two isotopes: Z-10 (25% abundance) and Z-11 (75% abundance).

Show your work:

Average atomic mass: _____ Matched app? Y / N: _____

Problem 3

Element M has three isotopes: M-50 (5% abundance), M-52 (83% abundance), and M-54 (12% abundance).

Show your work:

Average atomic mass: _____ Matched app? Y / N: _____



Part C — Work backward (challenge)

These problems give you the target average instead of the abundances. Set up an algebra equation to solve for the missing percentages — then check your setup using "Challenge: Hit the Target" mode.

Tip: Let x = the abundance (as a decimal) of the lighter isotope. Then $(1 - x)$ is the abundance of the heavier one. Set up: $(\text{mass}_1 \cdot x) + (\text{mass}_2 \cdot (1 - x)) = \text{target average}$.

Problem 1

An element has isotopes with mass numbers 10 and 12. Its average atomic mass is 10.80. Find the percent abundance of each isotope.

Show your work:

Lighter isotope %: _____ Heavier isotope %: _____

Problem 2

An element has isotopes with mass numbers 6 and 7. Its average atomic mass is 6.20. Find the percent abundance of each isotope.

Show your work:

Lighter isotope %: _____ Heavier isotope %: _____

Quick reference

Isotope	atoms of the same element with different numbers of neutrons
Abundance	the percentage of naturally occurring atoms that are a given isotope
Weighted average	an average where some values count more than others, based on how common they are



Reflection

Why is average atomic mass a "weighted" average rather than a simple average of the isotope masses? What would happen to an element's atomic mass if one of its isotopes suddenly became far more common than the others?



TEACHER ANSWER KEY — remove or skip this page before printing for students

Part B — Practice by hand

Problem 1: $(24 \times 0.60) + (26 \times 0.40) = 14.4 + 10.4 = 24.8$

Problem 2: $(10 \times 0.25) + (11 \times 0.75) = 2.5 + 8.25 = 10.75$

Problem 3: $(50 \times 0.05) + (52 \times 0.83) + (54 \times 0.12) = 2.5 + 43.16 + 6.48 = 52.14$

Part C — Work backward

Problem 1: $10x + 12(1-x) = 10.80 \rightarrow 10x + 12 - 12x = 10.80 \rightarrow -2x = -1.20 \rightarrow x = 0.60$. So 60% is mass 10, 40% is mass 12.

Problem 2: $6x + 7(1-x) = 6.20 \rightarrow 6x + 7 - 7x = 6.20 \rightarrow -x = -0.80 \rightarrow x = 0.80$. So 80% is mass 6, 20% is mass 7.

Reflection — sample answer

It's weighted because each isotope contributes to the average in proportion to how common it actually is in nature — a rare isotope should barely move the average, while a common one should dominate it. If one isotope suddenly became far more abundant, the average atomic mass would shift noticeably toward that isotope's mass number.

Teaching notes

- Problems in Part B (Q, Z, M) use fictional element letters intentionally, so students can't just look up the answer — they must calculate.
- Part C's "work backward" problems use clean, round target values by design (10.80 and 6.20) so the algebra resolves to whole-number percentages — good for building confidence before tackling real, messier isotope data.
- This worksheet pairs with all three modes of the Average Atomic Mass Calculator at learnsightsystems.com/resources/chemistry — Explore mode for Part A, Build Your Own for checking Part B, and Challenge mode for checking Part C's setups.