Average Atomic Mass

Unit: Atomic Structure

NGSS Standards/MA Curriculum Frameworks (2016): N/A

Mastery Objective(s): (Students will be able to…)

- Calculate the average atomic mass of an atom from percent abundance data.

Success Criteria:

- Solutions correctly turn masses into percentages.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: abundance

Language Objectives:

- Explain the laws of conservation of mass, definite proportions, and multiple proportions.

Notes:

mass number: the mass of *one individual atom* (protons + neutrons). Always a whole number.

abundance: the percentage of atoms of an element that are one specific isotope.

average atomic mass: the estimated weighted *average* of the mass numbers *of all of the atoms* of a particular element on Earth.

Use this space for summary and/or additional notes:

Calculating an Average

Suppose we want to calculate the class average on a test.

1. Multiply each score times the number of students who got that score.
2. Add up the number for each score to get the total points.
3. Divide the total by the number of students to get class average.

Suppose 20 students took a test and the scores were:

test score	50	60	70	80	90	100
# students who earned it	3	4	4	7	2	1

We would calculate class average by:

1. Multiply each score by the number of students who earned it:

   $50 \cdot 3 = 150$

   $60 \cdot 4 = 240$

   $70 \cdot 4 = 280$

   $80 \cdot 7 = 560$

   $90 \cdot 2 = 180$

   $100 \cdot 1 = 100$

2. Add up the totals from step #1.

   $150 + 240 + 280 + 560 + 180 + 100 = 1510$

3. Divide the total from step #2 by the number of students.

   $\frac{1510}{20} = \boxed{75.5}$

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Calculating an Average from Percentages

If we have percentages instead of the number of data points, convert those percentages to decimal fractions (divide by 100) and add the decimal fractions directly. (This is equivalent to pretending that you have a total of 100 items.)

This is how we calculate the average atomic mass on the periodic table for each element.

For example, zinc has five stable isotopes. We would calculate the average atomic mass of zinc as follows:

Mass Number	Actual Mass of Isotope (amu)	Relative Abundance
64	63.929	48.63 %
66	65.926	27.90 %
67	66.927	4.10 %
68	67.925	18.75 %
70	69.925	0.62 %

1. Convert percent abundances to fractions (divide by 100).

2. Multiply the fractional abundance times the atomic mass for each isotope.

Mass Number	Actual Mass of Isotope (amu)	Fractional Abundance	Contribution (amu)
64	63.929	0.4863	(63.929) (0.4863) = 31.089
66	65.926	0.2790	(65.926) (0.2790) = 18.393
67	66.927	0.0410	(66.927) (0.0410) = 2.774
68	67.925	0.1875	(67.925) (0.1875) = 12.636
70	69.925	0.0062	(69.925) (0.0062) = 0.434

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3. Add up the contribution from each isotope to get the total atomic mass.

$$31.089 + 18.393 + 2.774 + 12.636 + 0.434 = \boxed{65.396}$$

   The average atomic mass of zinc is 65.396 amu.

4. Check that your average atomic mass is in between the smallest and largest.

   Yes, 65.396 is between the smallest (63.929) and the largest (69.925).

Note that ***you can't assume the most common isotope from the average atomic mass***. The average atomic mass of 65.396 suggests that the most common isotope of zinc should be zinc-65, but that isotope doesn't actually exist.

Sample problem:

The atomic mass and abundance of the two stable isotopes of carbon are:

Isotope	Atomic Mass (amu)	Relative Abundance
$^{12}_{6}C$	12.000 000	98.93 %
$^{13}_{6}C$	13.003 355	1.07 %
$^{14}_{6}C$	14.003 242	< 0.000 001 %

Calculate the average atomic mass of carbon.

Solution:

1. Convert abundances to fractions

   98.93 % ÷ 100 = 0.9893   1.07 % ÷ 100 = 0.0107

   (We can ignore carbon-14 because the percentage is so small that it doesn't affect the average.)

2. Multiply abundance x mass # for each isotope

   0.9893 x 12.000 000 = 11.8716

   0.0107 x 13.003 355 = 0.1391

3. Add up the number from each isotope to get the total

   11.8716 + 0.1391 = 12.0107

4. Check that your answer is in between the mass of the smallest isotope and the mass of the largest one.

   Yes, 12.0107 is between 12.000 and 13.004.

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Homework Problems

Calculate the average atomic mass of each of the following elements, based on the percent abundance of their isotopes. For each element, your answers should agree with the atomic mass listed on the periodic table.

Because you could just look up the answers on the periodic table, ***you must show how to set up the calculations in order to receive credit***.

1. bromine

isotope	**atomic mass (amu)**	**relative abundance**
$^{79}_{35}Br$	78.9184	50.69 %
$^{81}_{35}Br$	80.9163	49.31 %

2. boron

isotope	**atomic mass (amu)**	**relative abundance**
$^{10}_{5}B$	10.0129	19.9 %
$^{11}_{5}B$	11.0093	80.1 %

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3. chlorine

isotope	atomic mass (amu)	relative abundance
$^{35}_{17}Cl$	34.9689	75.78 %
$^{37}_{17}Cl$	36.9659	24.22 %

4. magnesium

isotope	atomic mass (amu)	relative abundance
$^{24}_{12}Mg$	23.9850	78.99 %
$^{25}_{12}Mg$	24.9858	10.00 %
$^{26}_{12}Mg$	25.9826	11.01 %

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