**Calculate the average atomic weight when given isotopic weights and abundances**

[Calculate the isotopic abundances, given the atomic weight and isotopic weights](http://www.chemteam.info/Mole/AvgAtomicWt-Reverse.html)

To do these problems you need some information: the exact atomic weight for each naturally-occurring stable isotope and it’s percent abundance. These values can be looked up in a standard reference book such as the "Handbook of Chemistry and Physics." The unit associated with the answer can be either amu or g/mol, depending on the context of the question. If it is not clear from the context that g/mol is the desired answer, go with amu (which means atomic mass unit).

[This problem can also be reversed, as in having to calculate the isotopic abundances when given the atomic weight and isotopic weights.](http://www.chemteam.info/Mole/AvgAtomicWt-Reverse.html) Study the tutorial below and then look at the problems done in the reverse direction.

**Problem #1: Carbon**

|  |  |  |
| --- | --- | --- |
| mass number | exact weight | percent abundance |
| 12 | 12.000000 | 98.90 |
| 13 | 13.003355 | 1.10 |

To calculate the average atomic weight, each exact atomic weight is multiplied by its percent abundance (expressed as a decimal). Then, add the results together and round off to an appropriate number of significant figures.

This is the solution for carbon:

(12.000000) (0.9890) + (13.003355) (0.0110) = 12.011 amu

**Problem #2: Nitrogen**

|  |  |  |
| --- | --- | --- |
| mass number | exact weight | percent abundance |
| 14 | 14.003074 | 99.63 |
| 15 | 15.000108 | 0.37 |

This is the solution for nitrogen:

(14.003074) (0.9963) + (15.000108) (0.0037) = 14.007 amu

[Video: How to Calculate an Average Atomic Weight.](http://blip.tv/file/2699681)

|  |  |  |
| --- | --- | --- |
| **Problem #3: Chlorine** |   | **Problem #4: Silicon** |
| mass number | exact weight | percent abundance |   | mass number | exact weight | percent abundance |
| 35 | 34.968852 | 75.77 |   | 28 | 27.976927 | 92.23 |
| 37 | 36.965903 | 24.23 |   | 29 | 28.976495 | 4.67 |
|   |   |   |   | 30 | 29.973770 | 3.10 |
| The answer for chlorine: 35.453 |   | The answer for silicon: 28.086 |

This type of calculation can be done in reverse, where the isotopic abundances can be calculated knowing the average atomic weight. [Go to tutorial on reverse direction.](http://www.chemteam.info/Mole/AvgAtomicWt-Reverse.html)

**Problem #5:** In a sample of 400 lithium atoms, it is found that 30 atoms are lithium-6 (6.015 g/mol) and 370 atoms are lithium-7 (7.016 g/mol). Calculate the average atomic mass of lithium.

**Solution:**

1) Calculate the percent abundance for each isotope:

Li-6: 30/400 = 0.075
Li-7: 370/400 = 0.925

2) Calculate the average atomic weight:

x = (6.015) (0.075) + (7.016) (0.925)

x = 6.94 g/mol

**Problem #6:** A sample of element X contains 100 atoms with a mass of 12.00 and 10 atoms with a mass of 14.00. Calculate the average atomic mass (in amu) of element X.

**Solution:**

1) Calculate the percent abundance for each isotope:

X-12: 100/110 = 0.909
X-14: 10/110 = 0.091

2) Calculate the average atomic weight:

x = (12.00) (0.909) + (14.00) (0.091)

x = 12.18 amu (to four sig figs)

3) Here's another way:

100 atoms with mass 12 = total atom mass of 1200

10 atoms with mass 14 = total atom mass of 140

1200 + 140 = 1340 (total mass of all atoms)

Total number of atoms = 100 + 10 = 110

1340/110 = 12.18 amu

The first way is the standard technique for solving this type of problem. That's because we do not generally know the specific number of atoms in a given sample. More commonly, we know the percent abundances, which is different from the specific number of atoms in a sample.

**Problem #7:** Boron has an atomic mass of 10.81 amu according to the periodic table. However, no single atom of boron has a mass of 10.81 amu. How can you explain this difference?

**Solution:**

10.81 amu is an average, specifically a weighted average. It turns out that there are [two stable isotopes of boron:](https://secure.wikimedia.org/wikipedia/en/wiki/Isotopes_of_boron) boron-10 and boron-11.

Neither isotope weighs 10.81 amu, but you can arrive at 10.81 amu like this:

x = (10.013) (0.199) + (11.009) (0.801)

x = 1.99 + 8.82 = 10.81

**Problem #8:** Copper occurs naturally as Cu-63 and Cu-65. Which isotope is more abundant?

**Solution:**

Look up the atomic weight of copper: 63.546 amu

Since our average value is closer to 63 than to 65, we conclude that Cu-63 is the [more abundant isotope.](https://secure.wikimedia.org/wikipedia/en/wiki/Isotopes_of_copper)

**Problem #9:** Copper has two naturally occurring isotopes. Cu-63 has an atomic mass of 62.9296 amu and an abundance of 69.15%. What is the atomic mass of the second isotope? What is its nuclear symbol?

**Solution:**

1) Look up the atomic weight of copper:

63.546 amu

2) Set up the following and solve:

(62.9296) (0.6915) + (x) (0.3085) = 63.546

43.5158 + 0.3085x = 63.546

0.3085x = 20.0302

x = 64.9277 amu

3) The nuclear symbol is:

29-Cu-65

The correct symbol would have the 29 subscripted left of the Cu and the 65 would be superscripted left of the Cu.

**Problem #10:** Naturally occurring iodine has an atomic mass of 126.9045. A 12.3849 g sample of iodine is accidentally contaminated with 1.0007 g of I-129, a synthetic radioisotope of iodine used in the treatment of certain diseases of the thyroid gland. The mass of I-129 is 128.9050 amu. Find the apparent "atomic mass" of the contaminated iodine.

**Solution:**

1) Calculate mass of contaminated sample:

12.3849 g + 1.0007g = 13.3856 g

2) Calculate percent abundances of (a) natural iodine and (b) I-129 in the contaminated sample:

(a) 12.3849 g / 13.3856 g = 0.92524
(b) 1.0007 g / 13.3856 g = 0.07476

3) Calculate "atomic mass" of contaminated sample:

(126.9045) (0.92524) + (128.9050) (0.07476) = x

x = 127.0540 amu

**Practice Problems**

Calculate the average atomic weight for:

1) Magnesium

|  |  |  |
| --- | --- | --- |
| mass number | exact weight | percent abundance |
| 24 | 23.985042 | 78.99 |
| 25 | 24.985837 | 10.00 |
| 26 | 25.982593 | 11.01 |

2) Molybdenum

|  |  |  |
| --- | --- | --- |
| mass number | exact weight | percent abundance |
| 92 | 91.906808 | 14.84 |
| 94 | 93.905085 | 9.25 |
| 95 | 94.905840 | 15.92 |
| 96 | 95.904678 | 16.68 |
| 97 | 96.906020 | 9.55 |
| 98 | 97.905406 | 24.13 |
| 100 | 99.907477 | 9.63 |

3) Tin …… WOW!

|  |  |  |
| --- | --- | --- |
| mass number | exact weight | percent abundance |
| 112 | 111.904826 | 0.97 |
| 114 | 113.902784 | 0.65 |
| 115 | 114.903348 | 0.36 |
| 116 | 115.901747 | 14.53 |
| 117 | 116.902956 | 7.68 |
| 118 | 117.901609 | 24.22 |
| 119 | 118.903310 | 8.58 |
| 120 | 119.902200 | 32.59 |
| 122 | 121.903440 | 4.63 |
| 124 | 123.905274 | 5.79 |

The answers? Look on a periodic table!! Remember that the above is the method by which the average atomic weight for the element is computed. No one single atom of the element has the given atomic weight because the atomic weight of the element is an average, specifically called a "weighted" average.

Given the previous paragraph, here is a question you could be asked:

Silver has an atomic mass of 107.868 amu. Does any atom of any isotope of silver have a mass of 107.868 amu? Explain why.

The specific question is about silver, but it could be any element.

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