

Two atoms with the same number of protons but different number of neutrons are called isotopes. For instance, naturally occurring chlorine occurs as two isotopes. ^{37}Cl which has 17 protons and 20 neutrons, and ^{35}Cl which has 17 protons and 18 neutrons. A few elements, such as fluorine, phosphorus, and iodine occur naturally with just one isotope, but most elements contain a mixture of isotopes. Another example is element number 22, titanium (Ti), a light and strong metal used in jet engines and in artificial human joints. There are five naturally occurring isotopes of Ti. Each one has 22 protons in its nucleus, but the number of neutrons varies from 24 to 28. In chemical reactions, all isotopes of an element behave nearly identically since the number of neutrons does not affect its electron or chemical behavior. Tin has the largest number of naturally occurring stable isotopes, ten.

Each isotope of a particular element has a different mass, therefore the mass of one mole of a particular isotope has its own unique value. Since a sample of a chemical element is usually a mixture of all its naturally occurring stable isotopes, the atomic mass that is reported in the periodic chart must reflect the masses of all its naturally occurring stable isotopes. Since the isotopes occur in differing proportions, the average mass must proportionately reflect the masses of the different isotopes. Thus the average atomic mass of an element is a weighted average.

Calculating the weighted averages

For example, the analysis of the four isotopes of iron give the following data:

isotope	isotopic molar mass (g/mole)	naturally occurring abundance
^{54}Fe	53.940	5.82 %
^{56}Fe	55.935	91.66 %
^{57}Fe	56.935	2.19 %
^{58}Fe	57.933	0.33 %

To calculate the weighted average atomic mass (which is the average atomic mass given in the periodic chart), multiply the percent of each isotope (in decimal form) by its particular molar mass, then add those quantities together.

$$(0.0582 \times 53.940) + (0.9166 \times 55.935) + (0.0219 \times 56.935) + (0.0033 \times 57.933) = 55.85 \text{ g/mole}$$

Note that the average molar mass will usually round to the most abundant isotope. While this is not always true, it is a close enough approximation.

Calculating % abundances

The reverse calculation can also be made, since the average atomic mass is in the periodic chart, it can be used together with information about the mass of each isotope, ^{35}Cl and ^{37}Cl to calculate the percent natural abundance.

Using chlorine and its two isotopes as an example.

Knowing that % always adds up to 100, and when used as a decimal, the two values will always add up to 1. Thus you can name one isotope as x and the other as $1 - x$. It doesn't matter which one is labeled as which.

$$(x \times 35) + ([1 - x] \times 37) = 35.45$$

Distributing and gathering terms will yield a fractional value for x which allows the determination of the two percentages.

$$35x + 37 - 37x = 35.45$$

$$-2x = -1.55$$

$x = 0.775$ since x was defined as ^{35}Cl there must be approximately 77.5 % and by subtraction from 100 since there are only two isotopes, there must be approximately 22.5 % of ^{37}Cl .

This is only approximate because the actual isotopic molar masses are not 35 and 37, but rather 34.96885 and 36.9659 respectively.

You would only be asked to do this calculation for a element that has only two naturally occurring isotopes.