Atomic Mass

Atomic mass is based on a relative scale and the mass of ^{12}C (carbon twelve) is *defined* as 12 amu; so, this is an exact number.

Why do we specify ^{12}C? We do not simply state that the mass of a C atom is 12 AMU because elements exist as a variety of isotopes.

> Carbon exists as two major isotopes, ^{12}C, and ^{13}C (^{14}C exists and has a half life of 5730 y, ^{10}C and ^{11}C also exist and their half lives are 19.45 min and 20.3 days respectively). Each carbon atom has the same number of protons and electrons, 6. ^{12}C has 6 neutrons, ^{13}C has 7 neutrons, and ^{14}C has 8 neutrons and so on. So, we must specify which C atom defines the scale.

All the masses of the elements are determined relative to ^{12}C.

Average Atomic Mass

Since many elements have a number of isotopes, chemists use average atomic mass. On the periodic table the mass of carbon is reported as 12.011 amu. No single carbon atom has a mass of 12.011, but in a handful of C atoms the average mass of a carbon atom is 12.011.

Why 12.011?

If a sample of carbon was placed in a mass spectrometer the spectrometer would detect two different C atoms, ^{12}C and ^{13}C.

The *natural abundance* of ^{14}C, ^{10}C and ^{11}C in geologic (i.e. old) samples is so low that we cannot detect the effect these isotopes have on the average mass.

From the information collected from the mass spectrometer the average mass of a carbon atom is calculated.

The mass of ^{12}C is, of course, 12 amu.
^{13}C is 1.0836129 times heavier than ^{12}C; so, the mass of ^{13}C is 13.003355

amu.
98.89% of the sample is ^{12}C, and
1.11% of the sample is ^{13}C.

So, the natural abundance of ^{13}C is 1.11%, and the natural abundance of ^{12}C is 98.89%

The average mass is simply a weighted average.

If we know the *natural abundance* (the *natural abundance* of an isotope of an element is the percent of that isotope as it occurs in a sample on earth) of all the isotopes and the mass of all the isotopes we can find the average atomic mass. The average atomic mass is simply a weighted average of the masses of all the isotopes.

average mass C = (0.9889 x 12(exact) amu) + (0.0111 x 13.00335) amu

= 11.8668 + 0.144337

= 12.011137185

12.01 amu

Determine the average atomic mass of oxygen.

Isotope	**Atomic Mass (amu)**	**Natural Abundance**
$^{16}_{8}O$	15.99491	99.759%
$^{17}_{8}O$	16.99913	0.037%
$^{18}_{8}O$	17.99916	0.204%

average mass = (15.99491 amu x 0.99759) + (16.99913 amu x 0.00037) + (17.99916 amu x 0.00204)
average mass = 15.999 amu

$$\begin{array}{r} 15.95636 \\ 0.00629 \\ +\ 0.03672 \\ \hline 15.99937 \end{array}$$

Sig figs above are correct. After the multiplication step the significant figures which remain are in red. The last place that can be used after adding is the thousandths place.

Another kind of question could be asked…

Copper has two isotopes ^{63}Cu and ^{65}Cu. The atomic mass of copper is 63.54. If the atomic masses of ^{63}Cu and ^{65}Cu are 62.9296 and 64.9278 amu respectively what is the natural abundance of each isotope.

ave mass = (% ^{63}Cu /100 x mass ^{63}Cu amu) + (% ^{65}Cu /100 x mass ^{65}Cu amu)

^{63}Cu 62.9296 amu

^{65}Cu 64.9278 amu

63.54 amu = (% ^{63}Cu /100 x 62.9296 amu) + (% ^{65}Cu /100 x 64.9278 amu)

since (% ^{63}Cu) + (% ^{65}Cu) = 100

(% ^{63}Cu) = 100 – (% ^{65}Cu)

substitute

63.54 amu = ((100 – % ^{65}Cu)/100 x 62.9296 amu) + (% ^{65}Cu /100 x 64.9278 amu)

solve,

63.54 amu = 62.9296 amu – 0.629296 (% ^{65}Cu) + 0.649278 (% ^{65}Cu)

0.6104 = 0.019982 (% ^{65}Cu)

$31.0479 = (\%\ ^{65}Cu)$

$(\%\ ^{65}Cu) = 31.05$

$(\%\ ^{63}Cu) = 68.98$

Of course, a question like the one above could be turned around another way.

Gallium, atomic mass 69.72 amu, has two major isotopes, ^{69}Ga, atomic mass 68.9257 amu, and ^{71}Ga. If the natural abundance of each isotope is 60.00 and 40.00 % respectively, what is the mass (in amu) of ^{71}Ga.

$$69.72 \text{ amu} = (0.6000 \times 68.9257 \text{ amu}) + (0.4000 \times 71Ga)$$

$$^{71}Ga = 70.9249 \text{ amu}$$

The mole

Element	Average Mass of 1 Atom	Average Mass of 100 Atoms
C	12.01 amu	1,201 amu
H	1.0079 amu	100.79 amu
W	183.9 amu	18,390 amu

1 Atom (AMU) 100 Atoms (AMU)

$$\frac{C}{H} = \frac{12.01}{1.0079} = \frac{1201}{100.79} = \frac{1}{0.08400}$$

$$\frac{C}{W} = \frac{12.01}{183.9} = \frac{1201}{18390} = \frac{1}{15.31}$$

As long as we count the same number of atoms the ratio of the atomic masses stays the same.

Since atoms of C are so small, we could place enough of them on a balance so that the mass would be 12.01 g.

The same could be done with W; that is, 183.9 g of W could be placed on a balance.

$$\frac{\text{mass of W}}{\text{mass of C}} = \frac{183.9\text{ g}}{12.01\text{ g}} = \frac{15.31}{1}$$

Since the ratio of the masses is the same, each sample contains the same number of atoms.

We have given this special number of atoms a name. It is a **mole**. Even before a mole could be calculated it was used.

1 mole of C atoms ≈ **6.02214×10^{23}** atoms ≈ 12.01 g of C atoms

1 mole of eggs ≈ 6.02214×10^{23} eggs

The atomic mass of C is 12.01 amu. What is the mass of 1 C atom?

$$1\text{ C atom} \times \frac{1\text{ mol C atoms}}{6.02214 \times 10^{23}\text{ C atoms}} \times \frac{12.011\text{ g C}}{1\text{ mol C atoms}} = 1.9945 \times 10^{-23}\text{ g C}$$

Molar Mass/Molecular Mass

The molar mass of a molecule is simply the sum of the atomic masses.

CH_4 = 1 mole C atoms + 4 moles H atoms

$$\begin{array}{r} 12.01 \\ \underline{+\ 4 \times 1.0079} \end{array} \qquad \begin{array}{r} 12.01 \\ \underline{+\ 4.0316} \\ 16.0416 \end{array} = 16.04\text{ g/mole}$$