RELATIVE ATOMIC MASS

The use of Carbon-12

There are about molecules in one drop of water. To count these molecules individually would be nearly impossible. Therefore, atoms and molecules are counted by their mass. This is possible because all atoms of different elements have different masses. For this to work all atoms of elements need to be measured against something, the standard atom of reference is:

Calculate the number of subatomic particles in C-12:

To make it easier an atom of C-12 is said to weight exactly ( ).
The atomic mass units in C-12 is roughly equal to the number of and in one atom of C-12.
This means that one atomic unit (1u) is roughly equal to proton or neutron.

Definitions:

Relative isotopic mass: The mass of an atom of an relative to the mass of an atom of C- .

Relative atomic mass ( ): The mass of an atom of an relative to the mass of an atom of C- .

Relative molecular mass ( ): The mass of a relative to the mass of an atom of C- .

Relative formula mass ( ): The mass of a of a relative to the mass of an atom of C- .

Relative Molecular mass is used when the smallest unit is a .

Relative formula unit is used when the smallest unit is a of an .

Why 1/12th of Carbon?

It is important to include the phrase ' ' when describing relative masses. The reason being that it doesn't make sense otherwise!

We know that 1u is equal to roughly the mass one proton or one neutron. This means that we are effectively measuring the mass of one neutron or proton against the mass of one of an element.

e.g.

Hydrogen has proton. If we measure it compared to 1/12th of C-12 ( ) it is . Therefore, we should expect a value of . The Mr H is .

Helium has protons and neutrons. If I compare it to (1 atomic unit) it is roughly times as big. Therefore, the Mr of He should be , which it is.

Explaining the definitions

Relative isotopic mass does not contain the words . This is because we are comparing atom of an isotope directly with of a C-12 atom.

Relative atomic mass, formula mass and molecular mass all contain the phrase . This is because most elements have . Isotopes are of the same with the same number of but different number of . This means that isotopes of elements have different . When we are comparing these different isotopes with 1/12th of C-12 we have to take these masses into account.

To complicate matters different isotopes of an element, occur in different naturally in the environment.

e.g. Chlorine has two isotopes ^{35}Cl and ^{37}Cl it naturally occurs in abundances of 75% for ^{35}Cl and 25% for ^{37}Cl.

Therefore, if there was one hundred molecules of Chlorine in a jar, how many molecules of each isotope would there be?

How would you calculate the mass of all of the one hundred molecules in the jar?

How would you calculate the average mass of the chlorine molecules in the jar?

This is how we calculate the relative atomic mass of an element:

Calculate the relative atomic mass of Magnesium to three significant figures given that the three naturally occurring isotopes of Magnesium; ^{24}Mg, ^{25}Mg and ^{26}Mg have percentage abundancies of 79.0%, 10.0% and 11.0%.

A sample of Rubidium contains 3.61×10^{23} atoms of ^{85}Rb and 1.39×10^{23} atoms of ^{87}Rb, calculate the relative atomic mass of Rb in the sample to three significant figures.

Boron exists as two isotopes ^{10}B and ^{11}B, given the relative atomic mass of Boron is 10.8 what are the percentage abundancies of the two isotopes?