

Parts of an Atom

Atom - is electrically neutral ($\# p^+ = \# e^-$)
 - is composed of a nucleus containing protons and neutrons, and electrons that surround the nucleus.

Atomic Number - is the number of protons found in the nucleus of an atom.

Protons - are subatomic particles possessing a positive charge.

Neutrons - are subatomic particles possessing a neutral charge.

Electrons - are subatomic particles possessing a negative charge.
 For an atom, the electrons are equal to the atomic number.

Isotope - is a form of an element in which the atoms have the same number of protons as all other forms of that element, but it has a **different number of neutrons and therefore a different atomic mass**.

Mass Number - is the sum of the number of protons and neutrons.

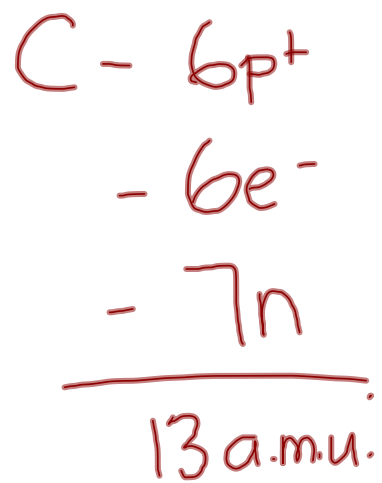
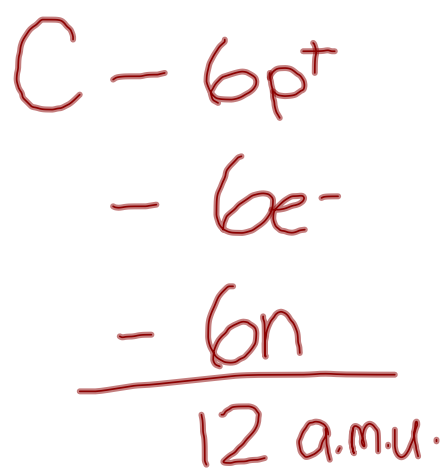
Carbon - 6 protons and 6 neutrons has a mass number of 12.

Another isotope of ^{12}C is ^{13}C , which has 6 protons and 7 neutrons.



MAIN SUBATOMIC PARTICLES

Particle	Location	Relative Mass	Charge
proton	nucleus	1 a.m.u.	+
neutron	nucleus	1 a.m.u.	none
electron	outside nucleus	small	-




isotopes

	Charge	Location	Relative Size
protons	+ 'ive	nucleus	1 a.m.u.
neutrons	neutral	nucleus	1 a.m.u.
electrons	- 'ive	outside nucleus	"massless"

Isotopes of Carbon

always has 6

changes

Isotope	p	n
${}^8\text{C}$	6	2
${}^9\text{C}$	6	3
${}^{10}\text{C}$	6	4
${}^{11}\text{C}$	6	5
${}^{12}\text{C}$	6	6
${}^{13}\text{C}$	6	7
${}^{14}\text{C}$	6	8
${}^{15}\text{C}$	6	9
${}^{16}\text{C}$	6	10
${}^{17}\text{C}$	6	11
${}^{18}\text{C}$	6	12
${}^{19}\text{C}$	6	13
${}^{20}\text{C}$	6	14
${}^{21}\text{C}$	6	15

most common

Calculating Atomic Mass

To calculate the atomic mass of an element, multiply the mass of each isotope by its natural abundance, expressed as a decimal, and then add the products.

atomic mass \rightarrow weighted average

Ex. Carbon has two stable isotopes: carbon - 12 (12.000 amu) which has natural abundance of 98.89%, and carbon - 13 (13.003 amu), which has natural abundance of 1.11%.
What is the atomic mass of carbon?

$$12.000 \text{ a.m.u.} - 98.89\%$$

$$13.003 \text{ a.m.u.} - 1.11\%$$

$$(12.000)(0.9889) + (13.003)(0.0111)$$

$$= \boxed{12.01}$$

Sample Problem

Element X has two natural isotopes. The isotope with a mass of 10.012 amu (^{10}X) has a relative abundance of 19.91%. The isotope with a mass of 11.009 amu (^{11}X) has a relative abundance of 80.09%. Calculate the atomic mass of this element.

Homework

Section 4.3 p. 110-118

Practice Problems #17-22

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