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 nu protons as all other forms of that element, but it has a **different numbe 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.

Isotope Notation:

$\boxed{^{12}_{6}C}$ mass # atomic # $\boxed{^{13}_{6}C}$

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	-

C – $6p^+$	C – $6p^+$
– $6e^-$	– $6e^-$
– 6n	– 7n
12 a.m.u.	13 a.m.u.

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

most common

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

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 → 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 a.m.u. – 98.89%

13.003 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