Warm Up

Isotope	protons	neutrons	electrons
copper - 64	29	35	29
Chromium -53	24	29	24
sulfur - 33	16	17	16
calcium- 41	20	21	20
gold - 108	79	29	79

$^{12}_{6}C$

$6p^+$

$6n$

$6e^-$

mass# 12

Carbon-12

$^{13}_{6}C$

$6p^+$

$7n$

$6e^-$

mass# 13

Carbon-13

l

Homework #15-20

20 chromium-50 chromium-52 chromium-53

$24p^+$ $24p^+$ $24p^+$

26n 28n 29n

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.

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?

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

$$= \boxed{12.01}$$

Mass Number	Atomic Mass
• Sum of p^+ + n ←	• decimal
• specific for each isotope	• weighted average
• whole number	

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.

$$10.012(0.1991) + 11.009(0.8009)$$

$$= 10.81$$

Homework

Practice Problems #21-24 p.110-118

Worksheet