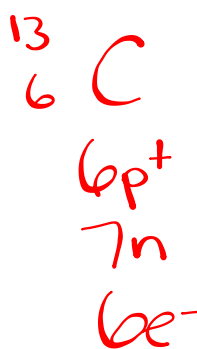


# Warm Up

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



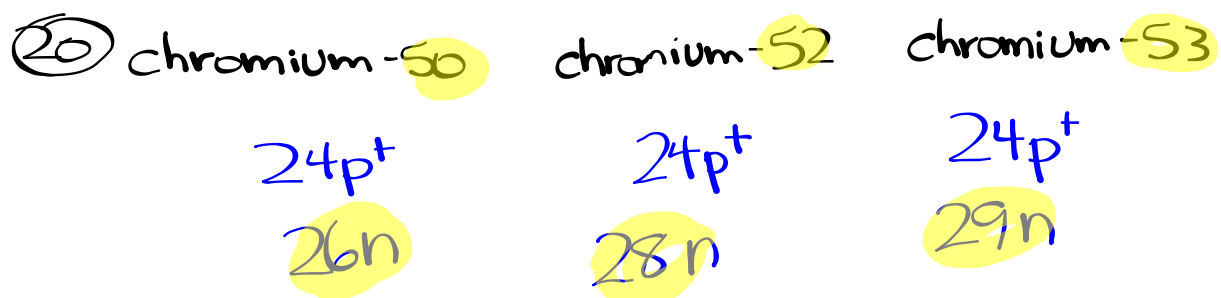
mass# 12  
Carbon-12



mass# 13  
Carbon-13

|

## Homework #15-20



# 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.

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

- Sum of  $p^+$  +  $n^0$
- specific for each isotope
- whole number

## Atomic Mass

- decimal
- weighted average

## Sample Problem

Element X has two natural isotopes. The isotope with a mass of 10.012 amu ( $^{10}\text{X}$ ) has a relative abundance of 19.91%. The isotope with a mass of 11.009 amu ( $^{11}\text{X}$ ) 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