GENERAL CHEMISTRY STANDARD 2.9

2.9: Calculate the atomic mass of a given sample with a given percent abundance, and vice versa

ISOTOPES AND ATOMIC MASS

- Atoms of the same element that have different numbers of neutrons
  - Isotopes have the same number of protons and electrons
  - Some elements have few isotopes, others have dozens

- The atomic mass for an element on the Periodic Table is the *average atomic mass of all natural isotopes of the element*.
  - Multiply the atomic mass of each isotope by its percent abundance
  - Make sure the percent abundance is in decimal form
  - All of the percent abundances should add up to 1.0 (within rounding error)
  - Add all of the products together to get the average atomic mass of the element
  - Be sure to follow significant figure rules for both the multiplication and addition parts!

ATOMIC MASS CALCULATION EXAMPLE

- Calculate the atomic mass of silicon. The three silicon isotopes have atomic masses and relative abundances of 27.9769 amu (92.2297%), 28.9765 amu (4.6832%), and 29.9738 amu (3.0872%).

27.9769 x 0.92297 = 25.822 amu
28.9765 x 0.046832 = 1.3570 amu
29.9738 x 0.030872 = 0.92341 amu

$$\begin{array}{r} 25.822 \\ 1.3570 \\ +\ 0.92341 \\ \hline 28.102\ \text{amu} \end{array}$$

ATOMIC MASS CALCULATION EXAMPLE

- Calculate the atomic mass of lead. The four lead isotopes have atomic masses and relative abundances of 203.973 amu (1.4%), 205.974 amu (24.1%), 206.976 amu (22.1%), and 207.977 amu (52.4%).

203.973 x 0.014 = 2.9 amu
205.974 x 0.241 = 49.6 amu
206.976 x 0.221 = 45.7 amu
207.977 x 0.524 = 109 amu

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   2.9
  49.6
  45.7
+ 109
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 208 amu
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