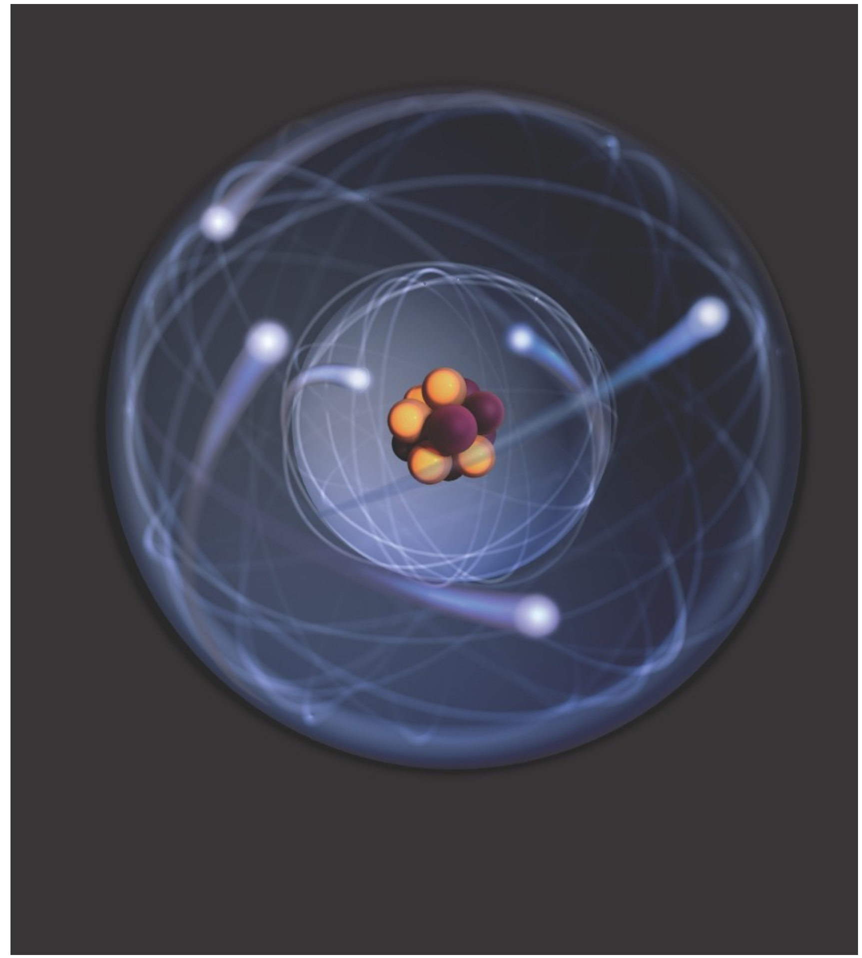
Isotopes

Chemistry 2020
Unit 2 Module 3

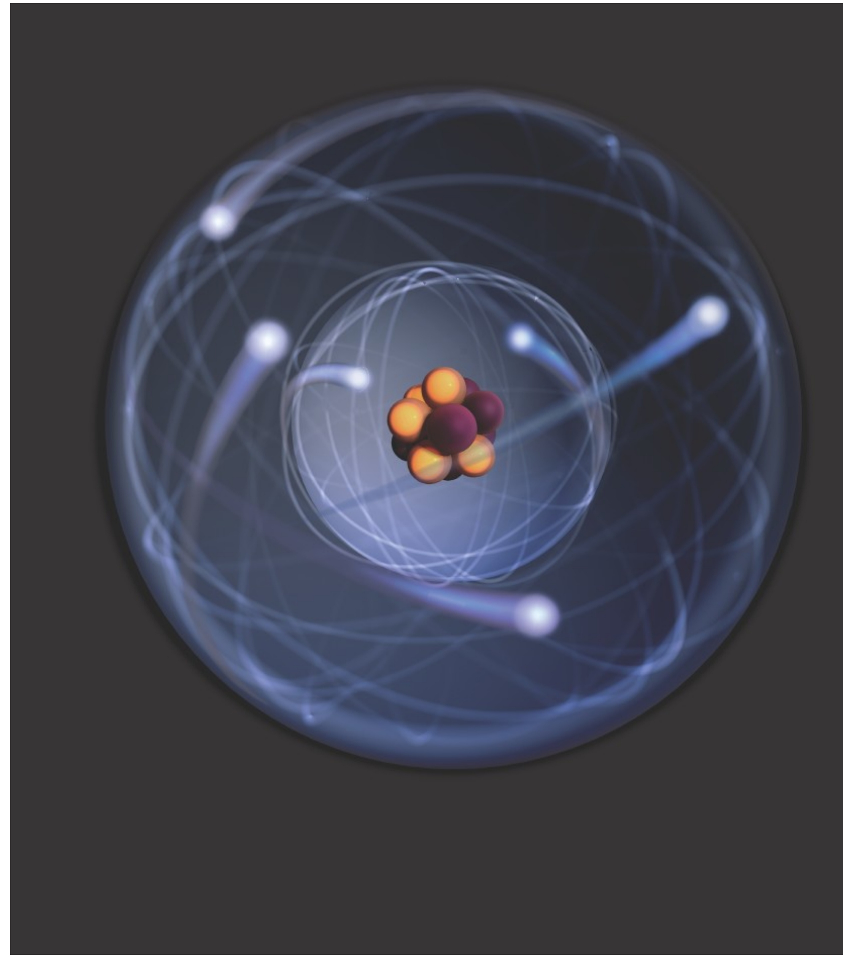
Review: Basics of Atomic Structure

- What are the three major subatomic particles?
- What charge does each one have?
- Where is each one located?

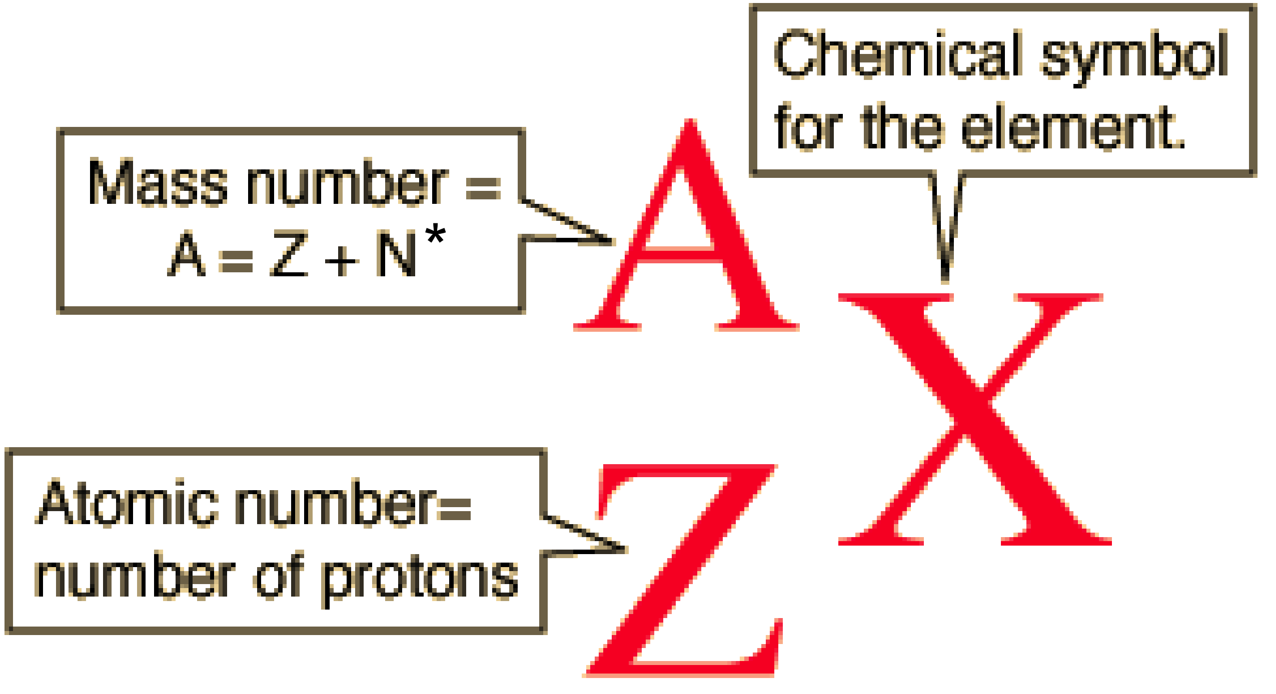


Review: Basics of Atomic Structure

- There are three major sub-atomic particles located within two regions of the atom
- Regions
  - Nucleus
  - Electron Cloud
- Particles
  - Proton (+)
  - Neutron
  - Electron (-)

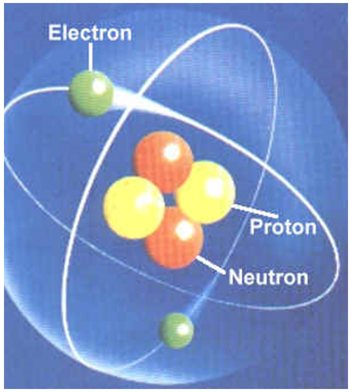


Nuclear Symbols for Neutral Atoms

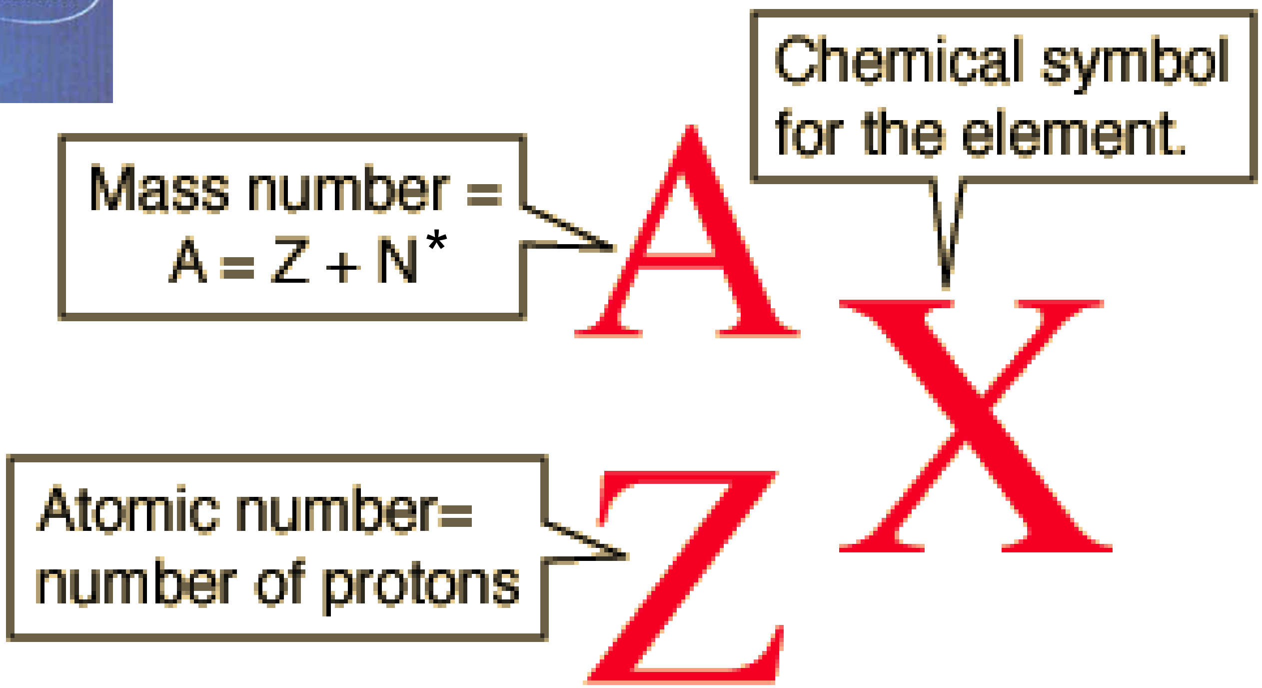


*N = number of neutrons

Nuclear Symbols for Neutral Atoms

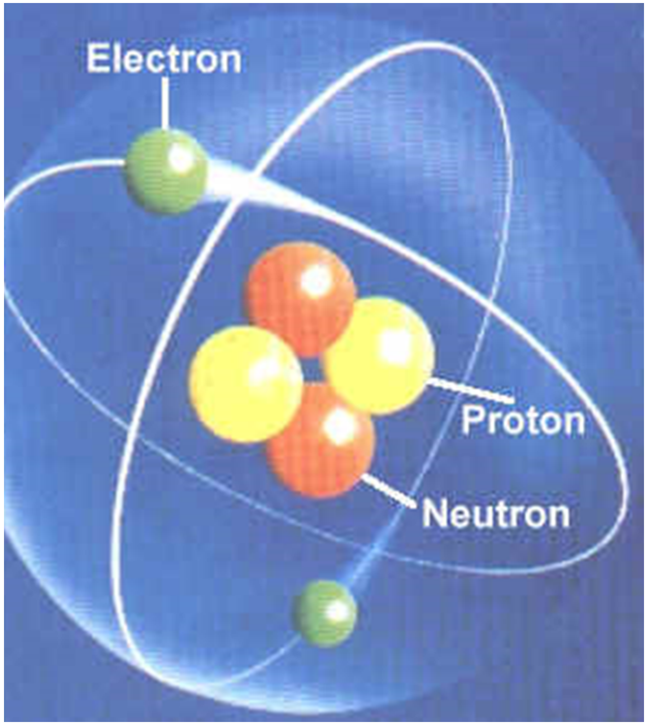


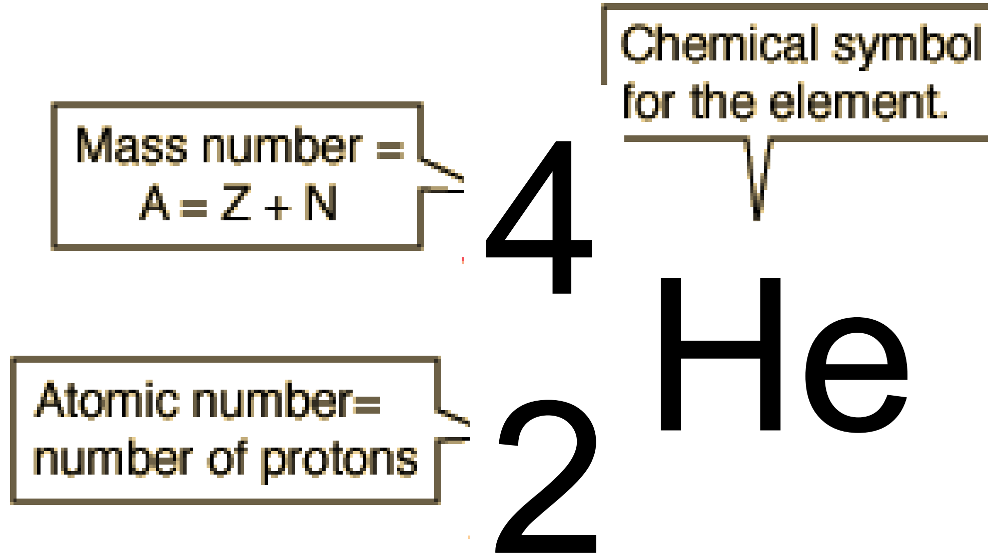
In this notation, the A and Z represent the number of particles in an atom or a population of atoms. It's not a mass.



*N = number of neutrons

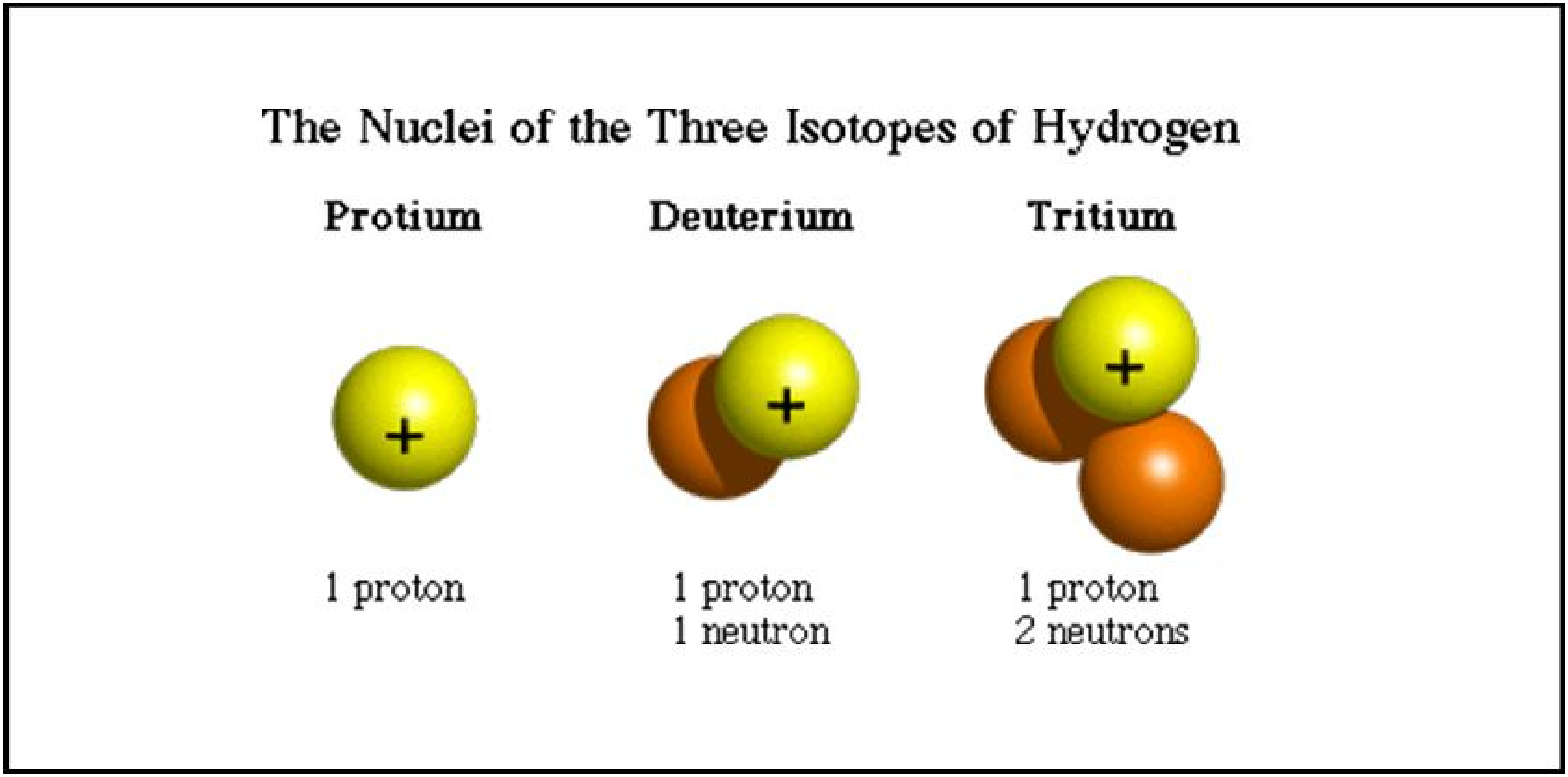
What Element Is This?



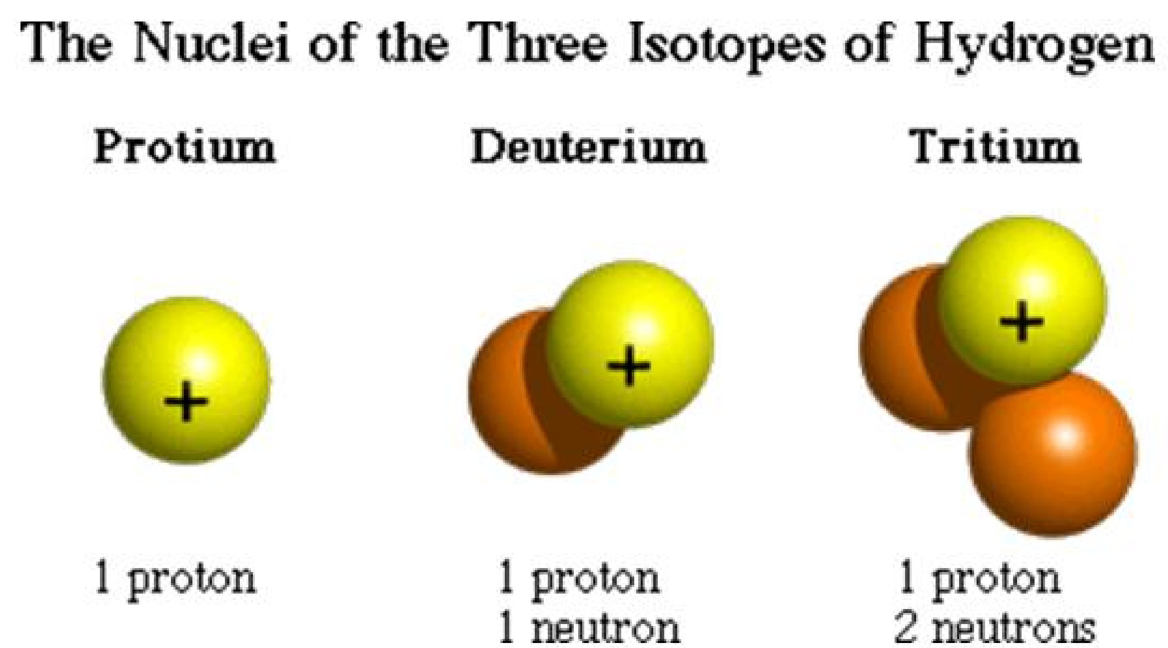


- There are 2 protons, so the *atomic number* is 2.
- The elemental *identity* of the atom depends <u>only</u> on the atomic number. It is helium, represented by He.
- There are 2 neutrons, so the *mass number* is 2+2=4.
- There are 2 electrons, so there is no charge.

The Isotopes of Hydrogen



The Isotopes of Hydrogen



$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$

If $m_{proton}=m_{neutron}=1amu$, what is the mass of each atom in amu?

The Isotopes of Hydrogen

Remind me...

1 mole H = 1.01g = 6.02×10^{23} atoms

There are a LOT of atoms in any lab-scale sample of a substance.

If I have 1 mole of H atoms, some of them will weigh 1 amu, some will weigh 2 amu, and some will weigh 3 amu.

So how do we come up with 1 mol H = 1.01g?

Average Atomic Mass

- The atomic mass given for an element on the periodic table is an average, which takes into account the various isotopes of an element and their relative abundance.
- Example: The mass of hydrogen found on the periodic table is 1.00794. Hydrogen has three isotopes, as shown in the previous module notes.

Average Atomic Mass

- So, the atomic mass of 1.00794 reflects a "weighted" average of all three of these isotopes. Which of the three isotopes of hydrogen must be the most abundant?

- Analogy - Your grades in many classes are calculated using weighted percentages. If the highest weighted category toward the calculation of your grade is tests, your overall average is most affected by your test performance.

- Apply this concept to average atomic mass. The "average" for Hydrogen is close to the number one. Which of the three isotopes of hydrogen must be the most abundant? Hydrogen – 1 (Protium), Hydrogen – 2 (Deuterium) or Hydrogen – 3 (Tritium)?

Calculating Average Atomic Mass

$$\begin{array}{c}\text{average}\\\text{atomic}\\\text{mass}\end{array} = \left(\begin{array}{c}\text{mass of}\\\text{isotope 1}\end{array}\right)\left(\begin{array}{c}\text{abundance}\\\text{of isotope 1}\end{array}\right) + \left(\begin{array}{c}\text{mass of}\\\text{isotope 2}\end{array}\right)\left(\begin{array}{c}\text{abundance}\\\text{of isotope 2}\end{array}\right) + \text{[and so on]}$$

Average Atomic Mass – Example #1

- Calculate the average atomic mass for magnesium which has three naturally occurring isotopes, Magnesium – 24 which is 78.70% abundant, Magnesium – 25 which is 10.13% abundant, and Magnesium – 26 which is 11.17% abundant.

Solution – Example #1

Average Atomic Mass =

[(24)(0.7870) + (25)(.1013) + (26)(0.1117)] =

18.888 + 2.5325 + 2.9042 = 24.3247 amu

= 24.32 amu

(customary to round final answer to 2 places after the decimal)

Average Atomic Mass – Example #2

- Element X has two isotopes. One is X-10 and has a mass of 10.012 amu and an abundance of 19.91%. The other is X-11, the average atomic mass for element X is 10.810 amu. What is the actual mass of X-11? What element is this?

Solution – Example #2

The sum of the isotope abundances must equal 100%. If the percent abundance of one isotope is 19.91%, the other must be 80.09%. Notice that when those percentages expressed as decimals are added together, it equals 1.0000!

$$10.810 = [(10.012)(0.1991) + X(0.8009)]$$

$$10.810 = 1.9934 + 0.8009X$$

$$8.8166 = 0.8009X$$

$$X = 11.008 \text{ amu}$$