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Isotopes

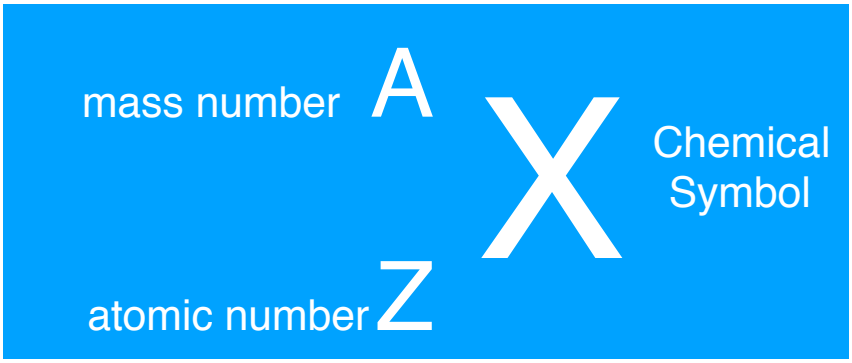
- Two atoms are isotopes if they have the same number of protons, but they have different numbers of neutrons.
- This means that:
  - Isotopes are atoms of the **same element**.
  - Isotopes have **different atomic masses**.
  - Isotopes have **different number of neutrons.**

Isotope Notation

- When using isotope notation we use:
  - Z to represent atomic number
  - A to represent mass number

To find the number of neutrons, N, subtract the atomic number from the mass number:

$$N = A - Z$$



Example Using Magnesium:

- Magnesium has 3 isotopes, here is how they compare. Using isotope notation, they are:

$^{24}_{12}Mg$	^{25}Mg	$^{26}_{12}Mg$
P=	P=	P=
E=	E=	E=
N=	N=	N=

- They have different number of neutrons and different atomic masses.

The Role of the Neutron

- As the atom grows, the number of neutrons increases more rapidly.
  - In larger atoms neutrons have a stabilizing effect (act as glue that hold atoms together).

Average Atomic Mass

- In the periodic table, the atomic mass of element is given in atomic mass units(u).
- To determine the atomic mass of an element, you must determine the *average atomic mass*.

Isotopic Abundance (% Abundance)

- ***Isotopic Abundance*** is the amount of a given isotope of an element that exists in nature, expressed as a percentage of the total amount of this element.

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Calculating Average Atomic Mass

Average Atomic Mass =

(Isotope 1 Abundance)(Mass Isotope 1)
+ (Isotope 2 Abundance)(Mass Isotope 2) . . .

*Remember isotope abundance is represented as a percentage, so it is expressed as a decimal.

Example: Using the information in the table below to calculate the average atomic mass of copper

Isotope	Mass(u)	Isotopic Abundance (%)
copper-63	62.93	69.2
copper-65	64.93	30.8

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Example: Using the information in the table below to calculate the average atomic mass of iron

Isotope	Mass(u)	Isotopic Abundance (%)
Iron-56	56.00	5.10
Iron-55	55.00	3.15
Iron-54	54.00	91.75