Atomic Mass

Isotopes

- Nuclei that have the same number of protons but a different number of neutrons
- Have identical chemical properties (all have same number of protons and thus electrons) but different physical properties
- The existence of isotopes is evidence for the existence of neutrons inside the nucleus

Mass number

$${}^{A}_{Z}\mathrm{X}$$

Element symbol

Atomic or proton number

Examples

${}^{1}_{1}\mathrm{H}$, ${}^{2}_{1}\mathrm{H}$, ${}^{3}_{1}\mathrm{H}$ (hydrogen)

${}^{238}_{92}\mathrm{U}$, ${}^{235}_{92}\mathrm{U}$ (uranium)

Sometimes we drop the atomic number

${}^{24}\mathrm{Mg}$, ${}^{25}\mathrm{Mg}$, ${}^{26}\mathrm{Mg}$ (magnesium)

We can also represent the isotope like this:

Sodium-23, Sodium-24 (sodium)

Carbon-12, Carbon-14 (carbon)

Atomic Mass Unit

- Developed as a convenient method to show the mass of an element
  - u or μ or amu
  - Defined as 1/12 of the mass of a carbon-12 atom
  - Based on carbon because
    - Carbon is a very common element
    - It results in a whole number for the atomic mass of most elements
    - It gave hydrogen (the lightest element) a mass of 1

Average Atomic Mass

- If the atomic mass unit is based on carbon, why doesn't carbon have a mass of exactly 12?
  - Carbon has more than one isotope
  - The mass quoted on the periodic table is really the average atomic mass of all atoms of carbon
  - The amount of each isotope of an element is known as the percent abundance

Calculating Average Atomic Mass

- The average atomic mass is the weighted average of all of the isotopes of the element

$$
\begin{aligned}
&(\text{mass of isotope 1})(\text{percent abundance of isotope 1}) + \\
&(\text{mass of isotope 2})(\text{percent abundance of isotope 2}) + \\
&(\text{mass of isotope 3})(\text{percent abundance of isotope 3}) + \ldots \\
&= \text{average atomic mass}
\end{aligned}
$$

Example

- Potassium (K)

Isotope	Percent Abundance	Atomic Mass (u)
K-39	93.2581	38.963707
K-40	0.0117	39.963998
K-41	6.7302	40.961826

$$
\begin{array}{r}
(38.963707)(0.932581) \\
(39.963998)(0.000117) \\
+\,\underline{(40.961826)(0.067302)} \\
39.098301\ \text{u}
\end{array}
$$

- We can also calculate the percentage abundance of an isotope if we know the mass of the isotopes and the average atomic mass

Example 1

- Magnesium (Mg)

Isotope	Percentage Abundance (%)	Atomic Mass (u)
Mg-24	78.70	23.98504
Mg-25	10.13	24.98584
Mg-26	?	25.98259

The average atomic mass of Mg is 24.30955 u. Determine the percentage abundance of Mg-26.

$$(23.98504)(.7870)+(24.98584)(.1013)+(25.98259)(x)=24.30955$$

Solve for x

$$\begin{aligned}18.87623+2.53107+25.98259x&=24.30955\\21.4073+25.98259x&=24.30955\\25.98259x&=2.90225\\x&=0.1117\end{aligned}$$

The percentage abundance of Mg-26 is 11.17%

Example 2

- Boron (B)

Isotope	Percentage Abundance (%)	Atomic Mass (u)
^{10}B	?	10.012937
^{11}B	?	11.009306

The average atomic mass of boron is 10.811028 u. Determine the percentage abundance of each isotope of boron.

$$(10.012937)(x)+(11.009306)(y)=10.811028$$
$$x+y=1$$

We need to solve a system of equations with 2 unknowns.

We need to rearrange the second equation and substitute it into the first equation.

$$y=1-x$$
$$(10.012937)(x)+(11.009306)(1-x)=10.811028$$

Now, solve for x

$$10.012937x+11.009306-11.009306x=10.811028$$
$$-0.996369x=-0.198278$$
$$x=0.199$$

Now that we know x, we can solve for y

$$(10.012937)(0.199)+(11.009306)(y)=10.811028$$
$$1.992574+11.009306y=10.811028$$
$$11.009306y=8.818454$$
$$y=0.801$$

Therefore, the percentage abundances of each isotope are:

^{10}B = 19.9%

^{11}B = 80.1%