Ch 2. Atoms and Elements

Introduction to Atoms

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ATOMIC COMPOSITION

• Protons
  > positive electrical charge
  > mass = \(1.672623 \times 10^{-24}\) g
  > relative mass = 1.007 atomic mass units (amu)

• Electrons
  > negative electrical charge
  > relative mass = 0.0005 amu

• Neutrons
  > no electrical charge
  > mass = 1.009 amu

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The modern view of the atom was developed by Ernest Rutherford (1871-1937).

Two of his students did an experiment in which they shot alpha particles (2 neutrons plus 2 protons) at a thin gold film. Some alpha particles scattered over large angles, demonstrating that atoms had massive, small nuclei.

See the book CD 2.10.

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In relative terms, if the nucleus were the size of a quarter (about 1 cm), one atom would have a radius of about \(10^5\) cm = 1000 m \(\sim\) \(1/2\) mile, roughly the size of the UMass campus.

nuclear radius is about 1 fm = \(10^{-15}\) m

atomic radius is about 100 pm = \(10^{-10}\) m
**Atomic Number, Z**

All atoms of the same element have the same number of protons in the nucleus, \( Z \)

\[
\begin{array}{c|c|c}
\text{atomic number} & \text{symbol} & \text{atomic weight} \\
13 & Al & 26.9815
\end{array}
\]

**Mass Number, A**

- C atom with 6 protons and 6 neutrons is the mass standard
- \( = 12 \) atomic mass units
- **Mass Number**
  - \( = \# \text{ protons} + \# \text{ neutrons} \)
- A boron atom can have
  - \( A = 5 \text{ p} + 5 \text{ n} = 10 \text{ amu} \)

\[
\begin{array}{c|c|c}
A & 10 \\
Z & 5 \\
\end{array}
\]

**Isotopes**

- Atoms of the same element (same \( Z \)) but different mass number (\( A \)).
- Boron-10 (\(^{10}\text{B}\)) has 5 p and 5 n.
- Boron-11 (\(^{11}\text{B}\)) has 5 p and 6 n.

\[
\begin{array}{c|c|c}
\text{Isotope} & \text{mass} & \text{number} \\
\hline
^{10}\text{B} & 5 \text{ p} & 5 \text{ n} \\
^{11}\text{B} & 5 \text{ p} & 6 \text{n}
\end{array}
\]

**Mass spectrometer**

- Ions of same charge but different mass are deflected by different amounts in the magnetic field.
- Can get the exact mass of each isotope of an element.
Isotope abundance

- Because of the existence of isotopes, the mass of a collection of naturally occurring atoms has an average value.
- Average mass = ATOMIC WEIGHT for the mixture of isotopes.
- Boron is 19.9% \(^{10}\text{B}\) and 80.1% \(^{11}\text{B}\). That is, of all the boron on earth, 80.1 percent is \(^{11}\text{B}\).
- Mass of \(^{10}\text{B}\) is 10.01 amu and of \(^{11}\text{B}\) is 11.01 amu
- Therefore, for boron, atomic weight
  \[
  = 0.199 (10.01 \text{ amu}) + 0.801 (11.01 \text{ amu}) = 10.8 \text{ amu}
  \]

This leaves us with 2 equations in 2 unknowns \((p_1\) and \(p_2\)), so we can do the algebra to solve them.

Substitute \(p_1 = 100 - p_2\) into the first equation:

\[
m = m_1 \times \left(\frac{100 - p_2}{100}\right) + m_2 \times \left(\frac{p_2}{100}\right)
\]

and then plug in the known values for \(m\), \(m_1\) and \(m_2\):

\[
1.00794 \text{ amu} = 1.0078 \text{ amu} \times \left(\frac{100 - p_2}{100}\right) + 2.0141 \text{ amu} \times \left(\frac{p_2}{100}\right)
\]

Solve:

\[
1.00794 = 1.0078p_2/100 + 0.0141p_2/100 = 1.0078 + 0.0063p_2/100
\]

or

\[
0.0001 = 0.0063p_2/100 \quad \Rightarrow \quad p_2 = 0.01%\]

PROBLEM: What is the natural abundance of \(^2\text{H}\)?
DATA: \(^1\text{H}\) mass is 1.0078 amu
\(^2\text{H}\) mass is 2.0141 amu
average mass of hydrogen is 1.00794

First, write down an equation for the average mass of hydrogen.

\[
m = m_1 \times \left(\frac{p_1}{100}\right) + m_2 \times \left(\frac{p_2}{100}\right)
\]

where \(m\) is mass and \(p\) is percentage

Now write down a second equation with \(p_1\) and \(p_2\) that totals to 100%.

\[
100 = p_1 + p_2
\]

Try to think about the chemistry before you think about the math:

Since the average mass of hydrogen is very close to that of \(^1\text{H}\), the dominant isotope of hydrogen will be \(^1\text{H}\), just as we saw in our solution (there is 0.01% \(^2\text{H}\) compared to 99.99% \(^1\text{H}\)).