Atomic Mass

What is an appropriate unit for measuring the mass of a single atom?

- A Gram is much too large
  - 1 Oxygen atom = 2.65×10^{-23} g !!
- A new standard was chosen
  - Atomic mass units (amu)
  - 1.00amu = 1.66×10^{-24} g

Atomic mass units

- ^{12}C was chosen to be the mass standard for measuring the mass of an atom
  - Abundant, stable, safe isotope
- 1 atom of ^{12}C was defined as having a mass of exactly 12 amu’s
- Therefore, 1 amu = 1/12 the mass of a ^{12}C atom

Atomic mass

- Why, then, does the periodic table list the atomic mass of carbon as 12.011, and not just 12?
- Carbon shows up in nature as more than one isotope
  - ^{12}C, ^{13}C, ^{14}C
- The higher the mass number, the more mass the atom has

Atomic mass

- So then, an average of the masses of the isotopes must be used
- But, then, why isn't the atomic mass of carbon 13 amu?

$$(12+13+14)/3 = 13$$

Atomic mass

- Not all isotopes show up to an equal extent, or as frequently as the others
- A "weighted" average must be used
- The atomic mass of an element is the weighted average of the masses of all the isotopes of that element.

Calculating average Atomic mass

Process:

1. multiply each isotope's mass by its percent abundance
   - Use the mass number of the isotope as the mass of the isotope if no other data is given
2. Add the numbers together

Atomic mass sample problem

- Example problem: potassium exists as three naturally occurring isotopes. ^{39}K has an abundance of 93.26%, ^{40}K is 0.01% abundant, and ^{41}K is 6.73% abundant. What is the atomic mass of potassium?

Atomic mass sample problem

Process:

1. multiply each isotope's mass by its percent abundance

   - Use the mass number of the isotope as the mass of the isotope if no other data is given

2. Add the numbers together

$$39 \text{ amu} \times 0.9326 = 36.371 \text{ amu}$$

$$40 \text{ amu} \times 0.0001 = 0.004 \text{ amu}$$

$$41\text{amu} \times 0.0673 = \underline{2.759 \text{ amu}}$$

$$39.134 \text{ amu}$$

Calculate the atomic mass of Silicon

Isotope name	**Isotope mass** (amu)	**Relative Abundance**
Silicon-28	27.98	92.21
Silicon-29	28.98	4.70
Silicon-30	29.97	3.09

$$27.98 \times 0.9221 = 25.80$$

$$28.98 \times 0.0470 = 1.36$$

$$29.97 \times 0.0309 = \underline{0.93}$$

$$28.09 \text{ amu}$$

Follow up questions

Q: How many atoms of silicon are expected to have a mass of exactly 28.09 amu?

A: None of them! It is the **average** mass of a silicon atom.

Q: How can we use this number as the mass of potassium atoms if none of them actually have this mass?

A: Because individual atoms are so small, we always use extremely large samples ($\uparrow 10^6$) of silicon atoms