Isotopes, Relative Abundance
Average Atomic Mass Calculations

Isotopes: atoms of the same element (i.e. same number of protons), but with a different number of neutrons and hence a different mass

- eg:

$^{37}_{17}Cl$	# of p^+ = 17 # of n = 20 # of e^- = 17
$^{35}_{17}Cl$	# of p^+ = 17 # of n = 18 # of e^- = 17

- same protons, different neutrons
- identical chemical properties (chemical properties are determined by the number of p^+ and hence the number of e^-)
- near identical physical properties*
- both types of Cl are present wherever chlorine is found!!

Natural Relative Abundance: the percent abundance of each isotope as found in nature

- extremely consistent for most elements found on earth

- eg: silver is a combination of Ag-107 and Ag-109 such that 51.84 % of silver is Ag-107 and the remaining 48.16 % of silver is Ag-109 (this percentage is the same for all silver no matter where it is mined on earth)

Precise Isotopic Mass: the exact mass of a given isotope based on the C-12 standard

- C-12 standard is that C-12 has a mass of 12.000000 u

- precise isotopic mass is affected by many factors, including Einstein's famous law of mass energy equivalency ($E = mc^2$) as it applies to the strong nuclear force in the nucleus
- mass must be determined experimentally (using a mass spectrometer - see handout)

- eg: Ag-107 has a mass of 106.905095 u, Ag-109 has a mass of 108.904754 u

Average Atomic Mass Calculation: a weighted averaging of isotopic masses gives the average atomic mass (appears to be the actual mass of the element as found in nature)

- eg: using the above information for silver

  Ag-107: 106.905095 u x 0.5184 = 55.419601 u
  <u>Ag-109: 108.904754 u x 0.4816 = 52.448529 u</u>
  107.868130 u

- note that the average atomic mass listed on the periodic table is 107.868 u (this is why!!!)