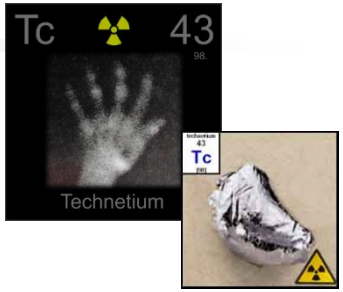
# CHEMISTRY 1000

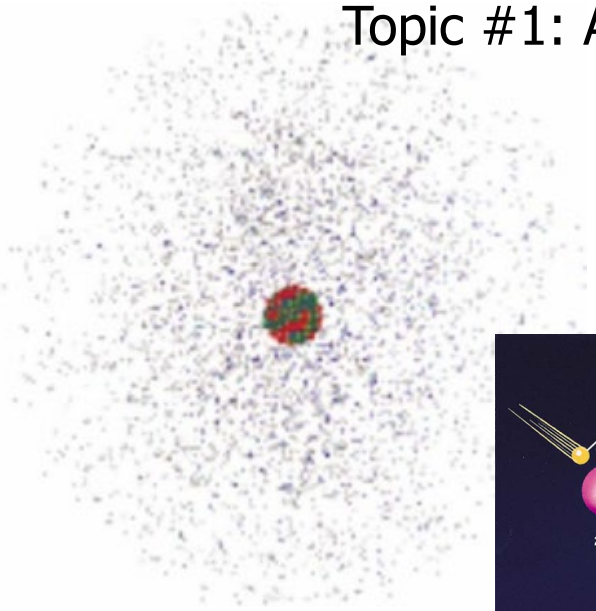


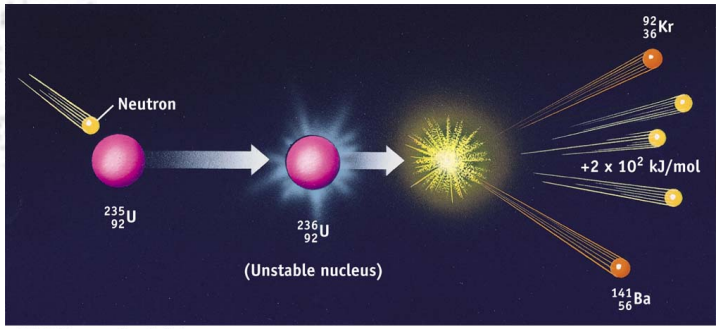

Topic #1: Atomic Structure and Nuclear Chemistry

Fall 2020

Dr. Susan Findlay

See Exercises 2.1 and 2.2



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# Who Thought of the Atom?

- Ancient Greek philosophers proposed that all matter consisted of some combination of four elements: air, earth, fire, water. Democritus (~460-370 B.C.) disagreed, proposing that all matter could be repeatedly subdivided until an indivisible particle was reached. He called this the atom (Greek: a = not; tomos = cut).
- In 1785, Antoine Lavoisier (1743-1794) is credited with discovery of the **law of conservation of mass**:

- In 1794, Joseph Proust (1754-1826) demonstrated the **law of constant composition** (*aka* **law of definite proportions**):

# Who Thought of the Atom?

- In 1808, John Dalton (1766-1844) based his **atomic theory of matter** on the ideas of Democritus, Lavoisier and Proust.
  - All matter consists of solid and indivisible atoms.
  - Atoms are indestructible and retain their identity in all chemical reactions.
  - All of the atoms of a given chemical element are identical in mass and in all other properties.
  - Different elements have different kinds of atoms; these atoms differ in mass from element to element.
  - Compounds consist of elements combined in small whole-number ratios.

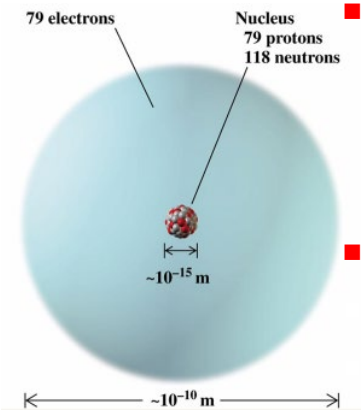
- While the essence of this theory has withstood the test of time, most of the postulates have since been modified. What notes would you add to the atomic theory listed here?

# What's in an Atom?

- Atoms consist of three types of subatomic particles:

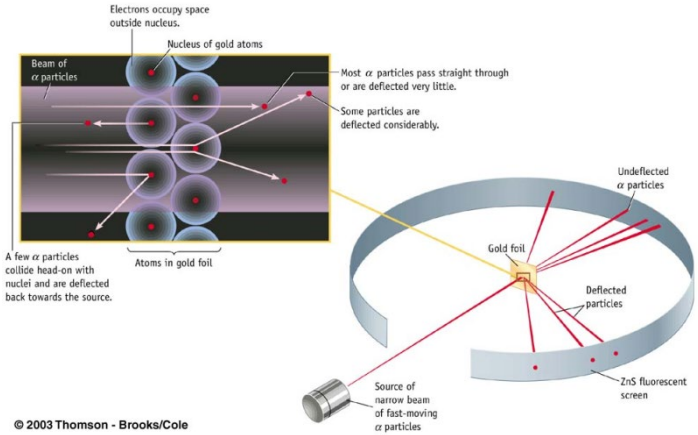
| | mass | | charge | | location |
|---|---|---|---|---|---|
| electron | $9.109382 \times 10^{-31}$ kg | 0.00054858 u | $-1.602176 \times 10^{-19}$ C | -1 | outer region |
| proton | $1.672622 \times 10^{-27}$ kg | 1.007276 u | $+1.602176 \times 10^{-19}$ C | +1 | nucleus |
| neutron | $1.674927 \times 10^{-27}$ kg | 1.008665 u | 0 C | 0 | nucleus |

- Most of an atom is empty space! (nucleus is only 5 x $10^{-46}$ $m^3$)



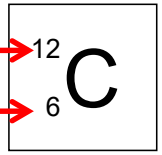

- A tiny nucleus (about 1/1,000,000,000,000,000$^{th}$ of the atom's volume) contains most of the atom's mass: protons and neutrons bound together in a region of positive charge (recall Rutherford's gold foil experiment).
- Electrons travel around the nucleus, balancing the overall charge of the atom:

Charge = #protons - #electrons



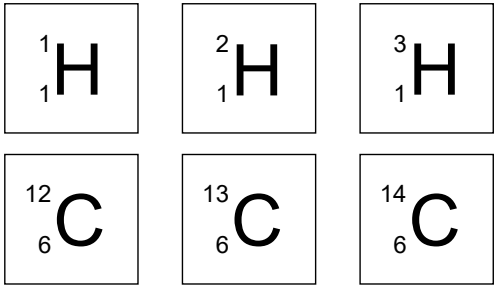
# Defining an Element

- Every atom has an atomic number and a mass number:
  - Mass number ($A$) = # nucleons*
  - Atomic number ($Z$) = # protons




- An element is defined by its atomic number. Changing the number of protons in an atom (as in a nuclear reaction) changes the element.

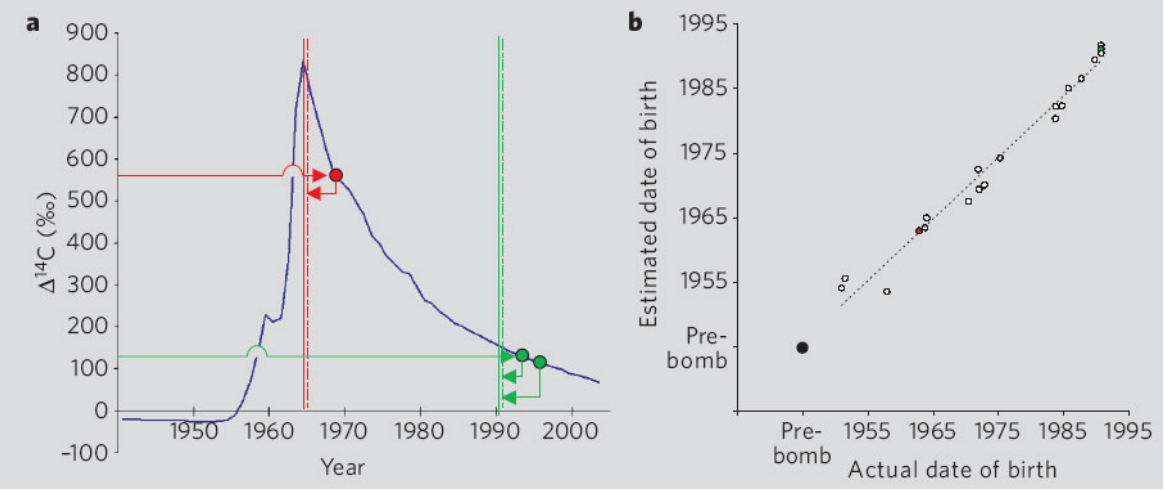
- While atoms of the same element must have the same atomic number, they may have different mass numbers. If so, they are referred to as **isotopes**. Most elements have more than one naturally occurring isotope:




*__nucleon__ is a general term for a proton or neutron (i.e. a particle in the nucleus)

# How Useful are Isotopes?

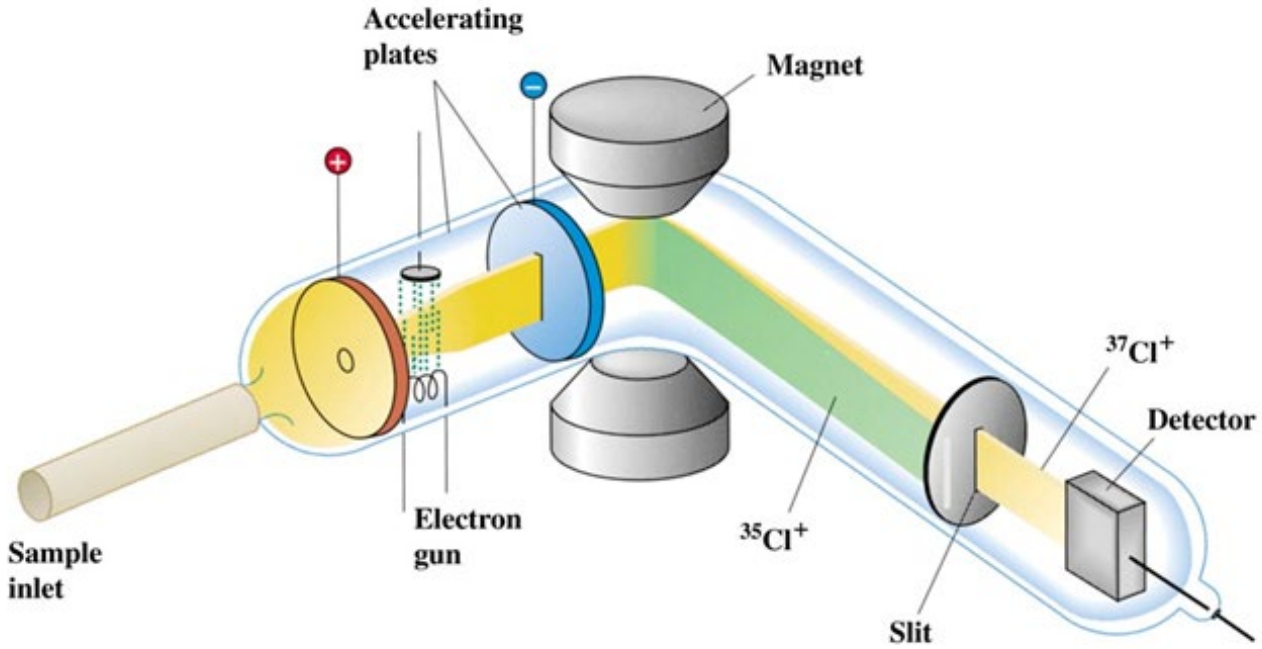
- Isotope ratios can be used to trace and/or date samples (in biology, geology, paleontology, archaeology, etc.)
- An interesting forensic application[1] was developed in Sweden
  - Atmospheric $^{14}C$ levels increased sharply during testing of nuclear bombs (1955-1963). Since banning the tests, $^{14}C$ levels decreased.
  - The amount of $^{14}C$ incorporated in tooth enamel can be used to determine the year in which the enamel was formed and, therefore, the year of birth (to within 1.6 years).




[1] K.L. Spalding, B.A. Buchholz, L.-E. Bergman, H. Druid and J. Frisén "Age Written in Teeth by Nuclear Tests" *Nature* **437**, 333 - 334 (14 Sep 2005).

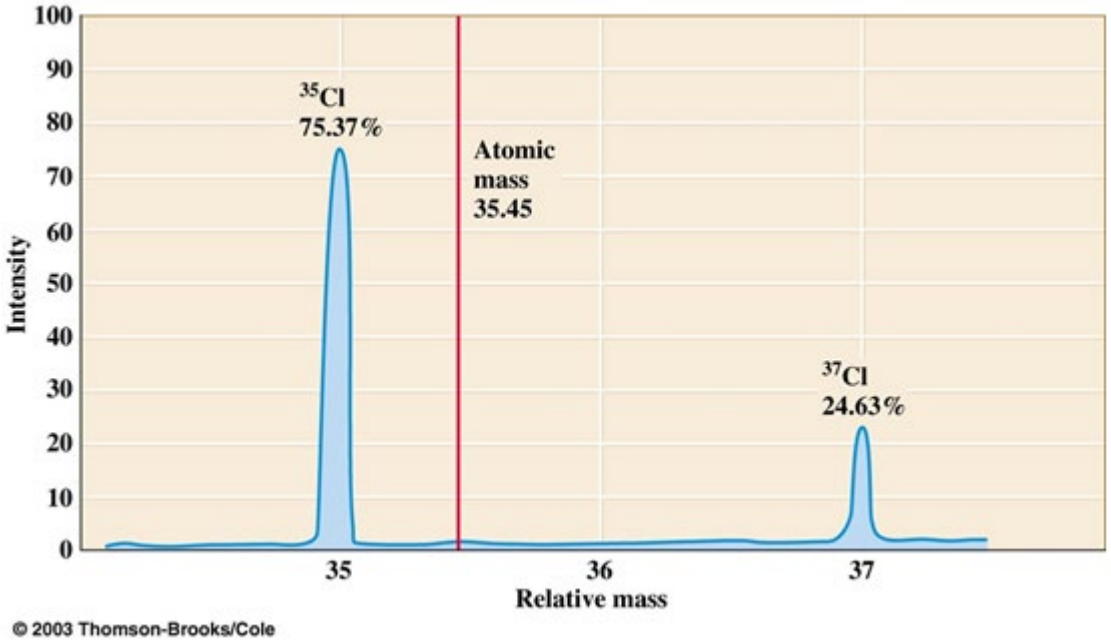
# How Can We Measure Isotopes?

- The average atomic mass of a chlorine atom is 35.4527 u. How can we determine which isotopes are present in a sample?
  - What is the main difference between isotopes of an element?
  - How can we take advantage of that difference?



# How Can We Measure Isotopes?

- The end result is a spectrum showing the proportion of atoms in the sample belonging to each isotope:




This is **mass spectrometry**! (the "MS" in GC/MS)

# "Average" Samples

- Only a few elements have just one naturally occurring isotope (e.g. $^{19}F$, $^{23}Na$, $^{31}P$). Most elements occur as mixtures of several isotopes. Chemists normally treat these elements as consisting of "averaged" atoms with "averaged" masses.

- **Atomic mass** (as shown on the periodic table) is the weighted average of all naturally occurring isotopes of an element. It factors in the mass and abundance of each isotope:

$$M_{av} = \sum\left(\frac{\%\text{ abundance of isotope}}{100\%}\right) \times M_{isotope}$$

*where*

$$\%\text{ abundance} = \frac{\#\text{ atoms of isotope}}{\text{total }\#\text{ atoms of element}} \times 100\%$$

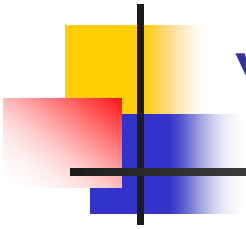
# "Average" Samples

- Silicon has three naturally occurring isotopes:
  - 92.23% $^{28}Si$ (27.9769 u)
  - 4.67% $^{29}Si$ (28.9765 u)
  - 3.10% $^{30}Si$ (29.9738 u).

  First estimate then calculate the average atomic mass of silicon.

# “Average” Samples

- Gallium has two naturally occurring isotopes and an average atomic mass of 69.723 u. $^{69}Ga$ has an atomic mass of 68.926 u, and $^{71}Ga$ has an atomic mass of 70.925 u. First predict which isotope is more abundant then calculate the natural abundance of each isotope of gallium.



# “Average” Samples