Isotopes

Review: Creating Ions: # of electrons changed, NOT protons

- gain e^- to create negative ion
- lose e^- to create positive ion

? what if keep protons constant But change neutrons?

Changing Neutrons $\rightarrow$ Mass changes

Isotopes: element with same number of protons and different number of neutrons
↳ makes them "heavier" or "lighter"

2 Different ways to express Isotopes:

1.) Hyphen notation

Cl - 35

Cl - 37

element name or abbreviation

Mass #

$\boxed{\text{MASS \# } = p + n}$

will always be a whole number AND may be different from mass on periodic table

2.) Nuclear notation

$^{35}_{17}Cl$

Mass # $\rightarrow$ 37

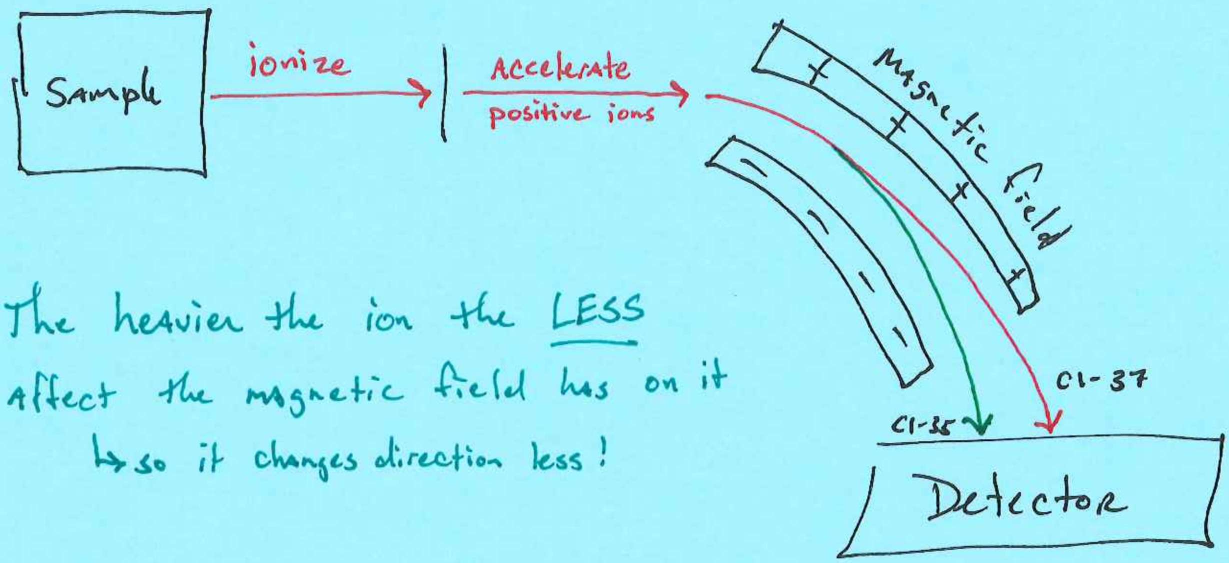
Atomic # $\rightarrow$ 17

$^{37}_{17}Cl$

Notice: in nuclear notation we are given atomic #

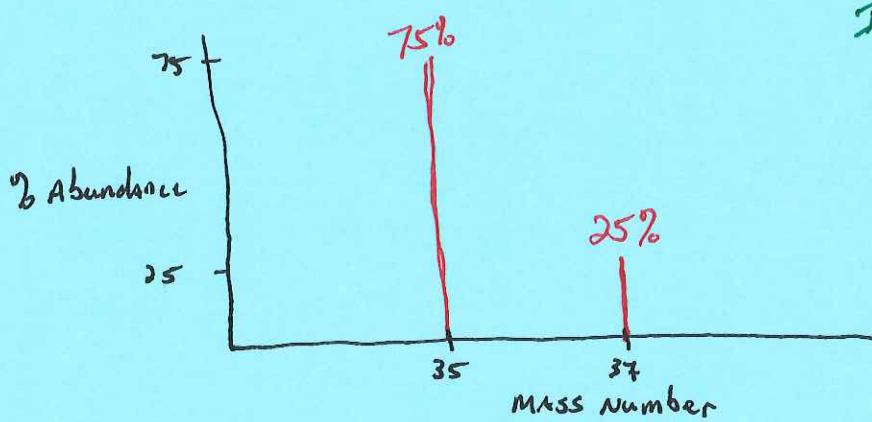
Atomic # = # of protons

Mass Spectrometer - uses a mass-vs-charge ratio to separate based on mass



✱ The heavier the ion the <u>LESS</u> affect the magnetic field has on it
↳ so it changes direction less!

<u>Data Output:</u>



Isotopes are naturally occurring <u>BUT</u> have differing abundances

Mass Spec data gives us the relative abundance of isotopes

Use Mass Spec Data

to calculate Average Atomic mass

this is a weighted average

Note!

A weighted average is NOT a regular average!

weighted average = the more abundant the more it weighs on the average

$35 \times 0.75 = 26.25$

$37 \times 0.25 = 9.25$

35.50 amu

Steps for calculating Average atomic mass

1.) Mass times relative abundance (Be sure to convert % back to decimal)

2.) add all products together

Br - Avg atomic Mass = 79.9

2 isotopes

Br - 79

Br - 81

Cu - Avg atomic Mass = 63.5

2 isotopes

Cu - 63

Cu - 65