Mass Relationships of Atoms

Atomic number and mass number

Atomic number (Z) = the number of protons in the nucleus.

Mass number (A) = the sum of the number of protons + neutrons in the nucleus.

$${}^{A}_{Z}X$$

Symbols for a few atoms

$^{1}_{1}H$ $^{4}_{2}He$ $^{12}_{6}C$ $^{16}_{8}O$ $^{65}_{30}Zn$

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Atomic number (Z)

Element

An element is a form of matter in which all of the atoms have the same atomic number.

However, two atoms of the same element can have different mass numbers.

Isotope

Atoms that have the same atomic number but different mass numbers are called isotopes.

- **same number of protons in nucleus**
- **differ in number of neutrons**

Some isotopes

$^{1}_{1}H$ $^{12}_{6}C$ $^{16}_{8}O$ $^{235}_{92}U$

$^{2}_{1}H$ $^{13}_{6}C$ $^{17}_{8}O$ $^{238}_{92}U$

$^{3}_{1}H$ $^{14}_{6}C$ $^{18}_{8}O$

Atomic masses

- synonymous with atomic weight
- is a relative scale
- mass-12 isotope of carbon (carbon-12) is the reference atom and assigned an atomic mass of exactly 12
- one atomic mass unit (amu) is defined as a mass exactly equal to $^{1}/_{12th}$ the mass of one carbon-12 atom

relative masses of carbon-12 and carbon-13 in a random sample carbon has a ratio of 1.0836129

$$\frac{{}^{13}_{6}\mathrm{C}}{{}^{12}_{6}\mathrm{C}} = 1.0836129$$

Since the atomic mass unit is defined such that the mass of ^{12}C is exactly 12 atomic mass units, then

$$\frac{{}^{13}_{6}C}{{}^{12}_{6}C} = 1.0836129$$

$$\text{Mass of } {}^{13}_{6}C = (1.0836129)(12 \text{ amu})$$

$$= 13.003355 \text{ amu}$$

The Mole

- the fundamental SI measure of "amount of substance"
- the amount of substance that contains as many elementary entities as there are atoms in exactly 12 g of carbon-12
- this number of atoms is 6.022045×10^{23}
  Avogadro's number

The Mole vs. The Dozen

The Dozen - the amount of substance that contains 12 entities.

The Mole - the amount of substance that contains Avogadro's number (6.022×10^{23}) of entities.

Dozen Apples = 10 Lbs.

Mole of Helium atoms = 4.0026g

Dozen Apples = 12 Apples

Mole of Helium atoms = 6.022×10^{23} atoms

Converting to Dozens

Example

How many dozens of apples are represented by 1.3 Lbs. of apples.

$$1.3 \text{ Lbs} \times \frac{1 \text{ dozen}}{10 \text{ Lbs}} = .13 \text{ dozen}$$

Converting to Moles Example

How many moles of He are in 6.46 g of He?

$$6.46\ \cancel{g}\ \text{He} \times \frac{1\ \text{mol}}{4.003\ \cancel{g}} = 1.61\ \text{mol}$$

The Mole

- since 6.022045×10^{23} atoms of carbon have a mass of 12 grams,

$$\frac{6.022045 \times \cancel{10^{23}\ \text{atoms}}}{12\ \text{g}} \times \frac{12\ \text{amu}}{1\ \cancel{\text{atom}}}$$

then

$$6.022045 \times 10^{23}\ \text{amu} = 1\ \text{g}$$

Molar mass of an element

The mass of 6.022 x 10^{23} atoms of an element is equal to its atomic mass in grams.

BUT: what does the periodic table tell us about the atomic mass of carbon?

Atomic mass is weighted average of mixture of isotopes

6
C
12.01

Atomic weight of carbon

= (atomic mass $^{12}_{6}C$) (fraction $^{12}_{6}C$)

+ (atomic mass $^{13}_{6}C$) (fraction $^{13}_{6}C$)

= (12.0000 amu)(0.9889)+(13.0035 amu)(0.0111)

= 11.8670 amu + 0.1441 amu

= 12.0111 amu

Example

Copper, a metal known since ancient times, is used in Electrical cables and pennies, among other things. The atomic masses of its two stable isotopes, $^{63}_{29}Cu$ (69.09%) and $^{65}_{29}Cu$ (30.91%), are 62.93 amu and 64.9278 amu, respectively. Calculate the average atomic mass of copper. The percentages in parentheses denote the relative abundances.

Answer

= (atomic mass $^{63}_{29}Cu$) (fraction $^{63}_{29}Cu$)

+ (atomic mass $^{65}_{29}Cu$) (fraction $^{65}_{29}Cu$)

= (62.93 amu)(0.6909)+(64.9278 amu)(0.3091)

= 43.47 amu + 20.07 amu

= 63.54 amu

Mass Relationships of Atoms - Extended

Example

Calculate the number of grams of lead (Pb) In 12.4 moles of lead.

$$12.4\ \cancel{mol} \times \frac{207.2\ g}{1\ \cancel{mol}} = 2569\ g$$

Example

What is the mass in grams of one silver atom?

$$\frac{107.9\ \text{g}}{1\ \text{mol}} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}$$

$$= \frac{17.91 \times 10^{-23}\ \text{g}}{1\ \text{atom}}$$

Example

Calculate the number of atoms in 0.551 g of Potassium (K).

$$0.551\ \text{g} \times \frac{1\ \text{mol}}{39.10\ \text{g}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}$$

$$= 8.48 \times 10^{21}\ \text{atoms K}$$

Example

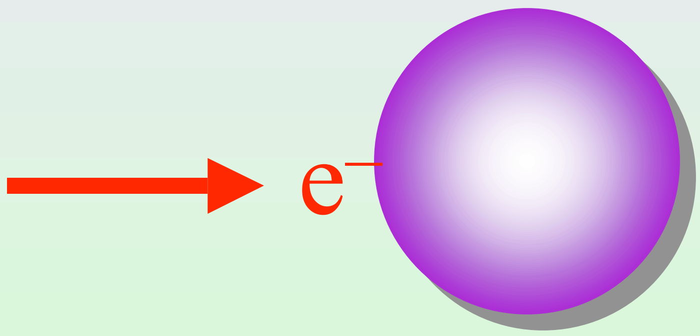
Calculate the number of molecules in a sample of oxygen gas (O_2) with a mass of 64.0g.

$$64.0\ \cancel{g} \times \frac{1\ \cancel{mol}}{32.0\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molec.}}{1\ \cancel{mol}}$$

$$= 1.20 \times 10^{24}\ \text{molec. } O_2$$

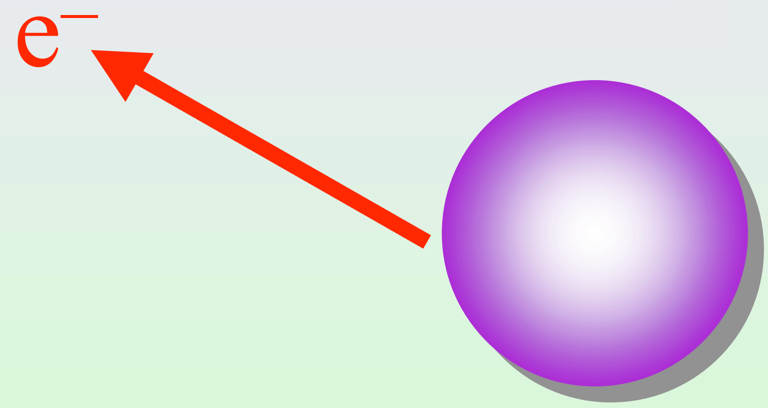
Experimental Determination Of Atomic & Molecular Masses

Atomic mass is measured by mass spectrometry



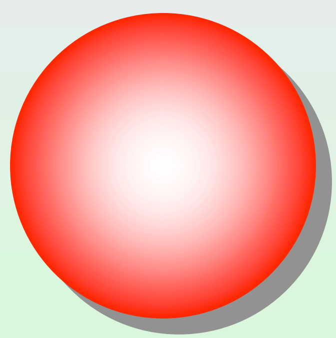
Atom is bombarded by stream of high Energy electrons.

Atomic mass is measured by mass spectrometry



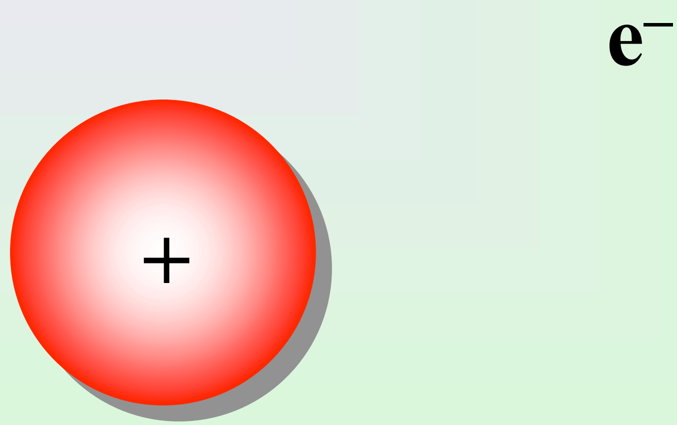
Electron collides with atom, “bounces” off

Atomic mass is measured by mass spectrometry



and transfers some of its energy to it.

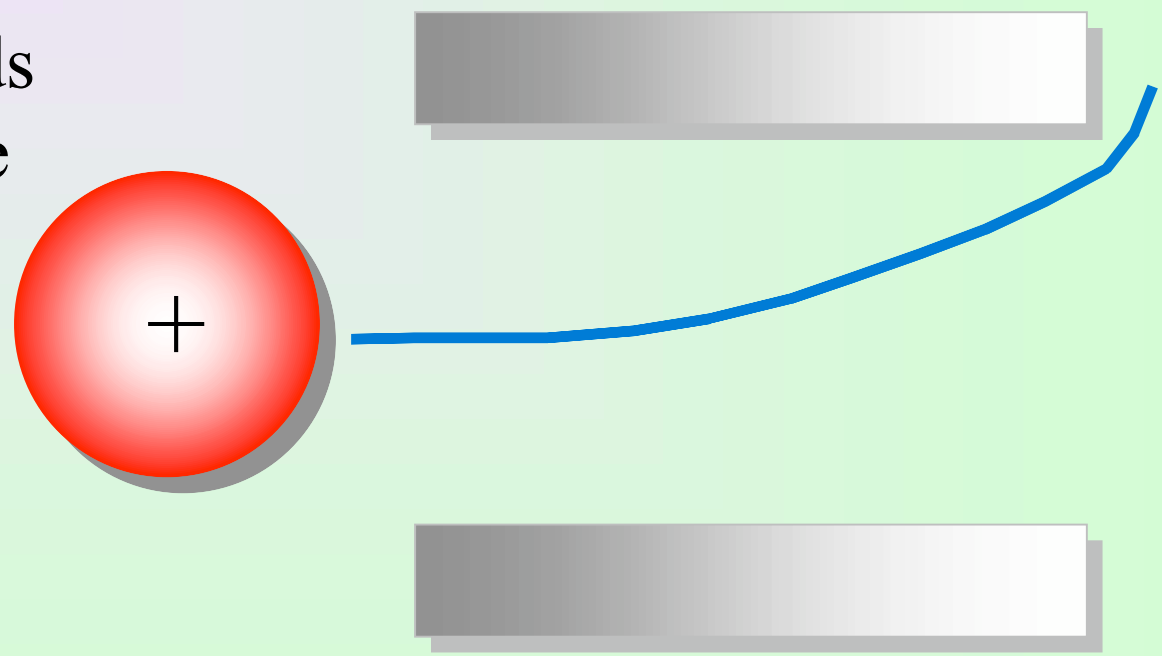
Atomic mass is measured by mass spectrometry



Atom dissipates its excess energy by expelling one of its electrons.

Ion is deflected by magnetic field

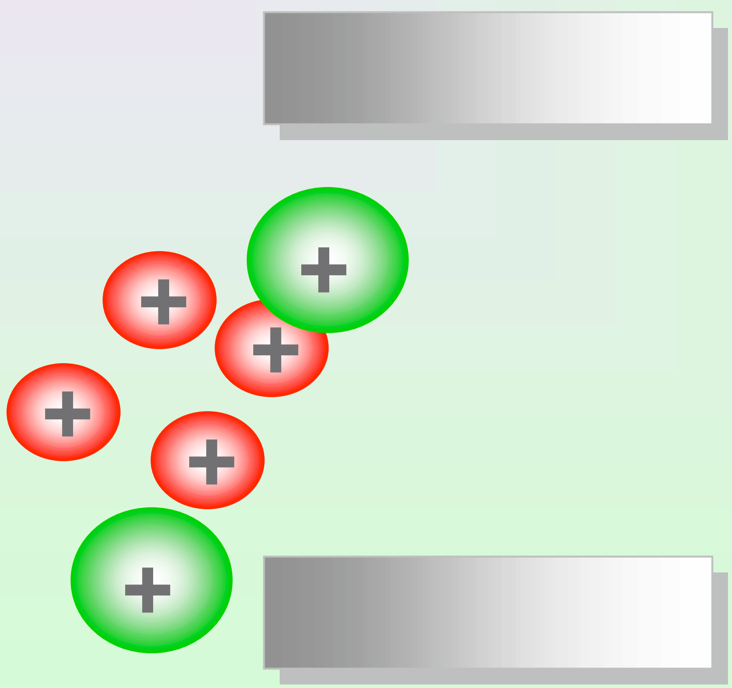
amount of deflection depends on mass to charge ratio



highest m/z deflected least

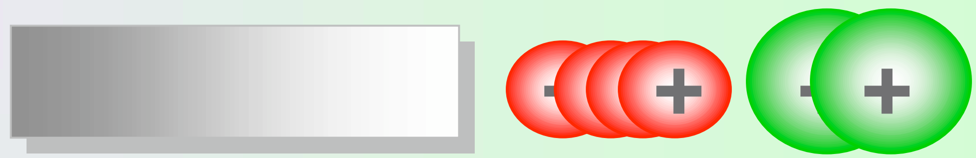
lowest m/z deflected most

Ions are detected after passage through magnetic field



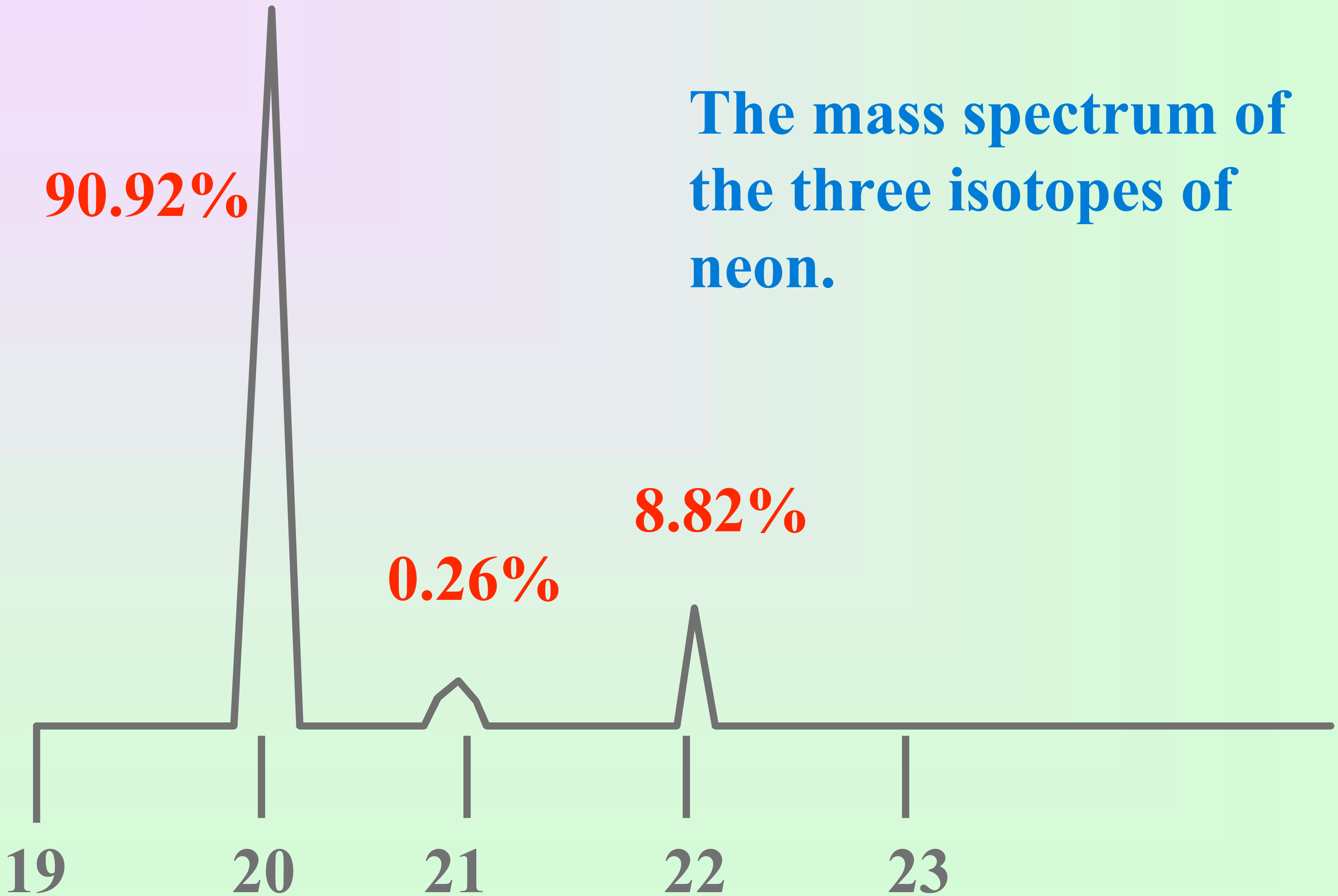
Ions are detected after passage through magnetic field

mixture of ions of different mass gives separate peak for each m/z



intensity of peak proportional to percentage of each atom of different mass in mixture

separation of peaks depends on relative mass



The mass spectrum of the three isotopes of neon.