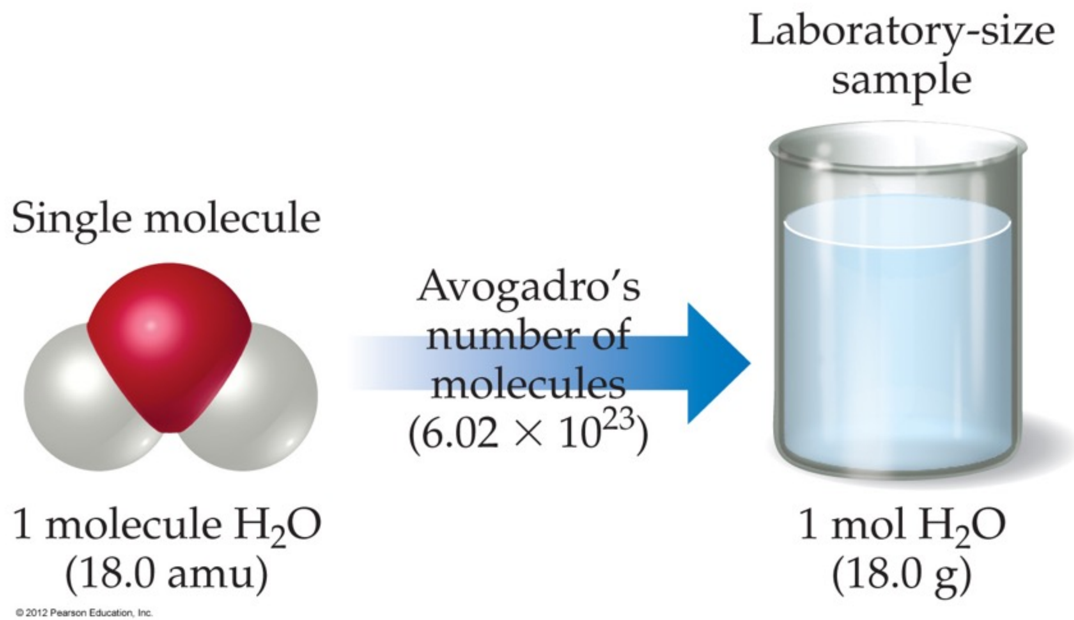
# Mass Relationships of Atoms

# 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 $6.02 \times 10^{23}$

Avogadro's number

# Avogadro's Number




- $6.02 \times 10^{23}$
- 1 mole of $^{12}C$ has a mass of 12.000 g.

# 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 ($6.02 \times 10^{23}$) of entities.

**Dozen tootsie rolls = 120.45g.**

**Mole of Helium atoms = 4.0026g**

**Dozen tootsie rolls = 12 tootsie rolls**

**Mole of Helium atoms = $6.02 \times 10^{23}$ atoms**

# Converting to Dozens

## Example

**How many dozens of tootsie rolls are represented by 13 g of tootsie rolls.**

$$13 \text{ g tootsie} \times \frac{1 \text{ dozen tootsie}}{120.45 \text{ g}} = 0.11 \text{ dozen}$$

# Converting to Moles Example

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

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

# The Mole

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

$$\frac{6.02 \times 10^{23}\ \text{atoms}}{12\ \text{g}} \times \frac{12\ \text{amu}}{1\ \text{atom}}$$

**6.02 x $10^{23}$ amu = 1 g**

# Mole Relationships

**TABLE 3.2 • Mole Relationships**

| Name of Substance | Formula | Formula Weight (amu) | Molar Mass (g/mol) | Number and Kind of Particles in One Mole |
|---|---|---|---|---|
| Atomic nitrogen | $N$ | 14.0 | 14.0 | $6.02 \times 10^{23}$ N atoms |
| Molecular nitrogen | $N_2$ | 28.0 | 28.0 | $6.02 \times 10^{23}$ $N_2$ molecules<br>$2(6.02 \times 10^{23})$ N atoms |
| Silver | $Ag$ | 107.9 | 107.9 | $6.02 \times 10^{23}$ Ag atoms |
| Silver ions | $Ag^+$ | 107.9* | 107.9 | $6.02 \times 10^{23}$ $Ag^+$ ions |
| Barium chloride | $BaCl_2$ | 208.2 | 208.2 | $6.02 \times 10^{23}$ $BaCl_2$ formula units<br>$6.02 \times 10^{23}$ $Ba^{2+}$ ions<br>$2(6.02 \times 10^{23})$ $Cl^-$ ions |

*Recall that the electron has negligible mass; thus, ions and atoms have essentially the same mass.



- One mole of atoms, ions, or molecules contains Avogadro's number of those particles.

- One mole of molecules or formula units contains Avogadro's number times the number of atoms or ions of each element in the compound.

# Mass Relationships of Atoms - Extended

# Example

**Determine the number of atoms in 2.5 mol Zn.**

$$2.5 \text{ mol Zn} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol Zn}} = 1.50 \times 10^{24} \text{ atoms Zn}$$

# Example

How many moles in $5.75 \times 10^{24}$ atoms Al ?

$$5.75 \times 10^{24} \text{ atoms Al} \times \frac{1 \text{mol Al}}{6.02 \times 10^{23} \text{ atoms Al}} = 9.55 \text{ mol Al}$$

# Example

**Determine the number of moles in 25.5g Ag.**

$$25.5\ \text{g Ag} \times \frac{1\text{mol Ag}}{107.86\ \text{g Ag}} = 0.24\ \text{mol Ag}$$

# Example

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

$$12.4 \text{ mol Pb} \times \frac{207.2 \text{ g}}{1 \text{ mol Pb}} = 2.57 \times 10^3 \text{ g}$$

# Example

**How many atoms in 55.2 g Li ?**

$$55.2 \text{ g Li} \times \frac{1 \text{mol Li}}{6.94 \text{ g Li}} \times \frac{6.02 \times 10^{23} \text{ atoms Li}}{1 \text{mol Li}}$$

$$= 4.79 \times 10^{24} \text{ atoms Li}$$

# Example

**What is the mass in grams of 1.00 x $10^{24}$ atoms Mn ?**

$$1.00 \times 10^{24} \text{ atoms Mn} \times \frac{1 \text{ mol Mn}}{6.02 \times 10^{23} \text{ atoms Mn}} \times \frac{54.9 \text{ g Mn}}{1 \text{ mol Mn}}$$

= **91.2 g Mn**

# Example

**Determine the number of moles of chloride ions in 2.50 mol $ZnCl_2$.**

$$2.50\ \text{mol}\ ZnCl_2 \times \frac{2\text{mol}\ Cl^-}{1\text{mol}\ ZnCl_2} = 5.00\ \text{mol}\ ZnCl_2$$

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**Determine the number of moles of each element in 1.25 mol glucose $C_6H_{12}O_6$.**

$$1.25 \text{ mol } C_6H_{12}O_6 \times \frac{6 \text{ mol C}}{1 \text{ mol } C_6H_{12}O_6} = 7.50 \text{ mol C}$$

$$1.25 \text{ mol } C_6H_{12}O_6 \times \frac{12 \text{ mol H}}{1 \text{ mol } C_6H_{12}O_6} = 15.0 \text{ mol H}$$

$$1.25 \text{ mol } C_6H_{12}O_6 \times \frac{6 \text{ mol O}}{1 \text{ mol } C_6H_{12}O_6} = 7.50 \text{ mol O}$$

# Example

**What is the mass in grams of one silver atom?**

$$\frac{107.9 \text{ g}}{1 \text{ mol Ag}} \times \frac{1 \text{ mol Ag}}{6.02 \times 10^{23} \text{ atoms}}$$

$$= \frac{17.9 \times 10^{-23} \text{ g}}{1 \text{ atom Ag}}$$

# Example

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

$$0.551\ \cancel{g}\ K \times \frac{1\ \cancel{mol}\ K}{39.10\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{atoms K}}{1\ \cancel{mol}\ K}$$

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

# Molecular Mass

# Molecular Mass

**synonymous with molar mass and molecular weight**

**is the sum of the atomic masses of all the atoms in a molecule**

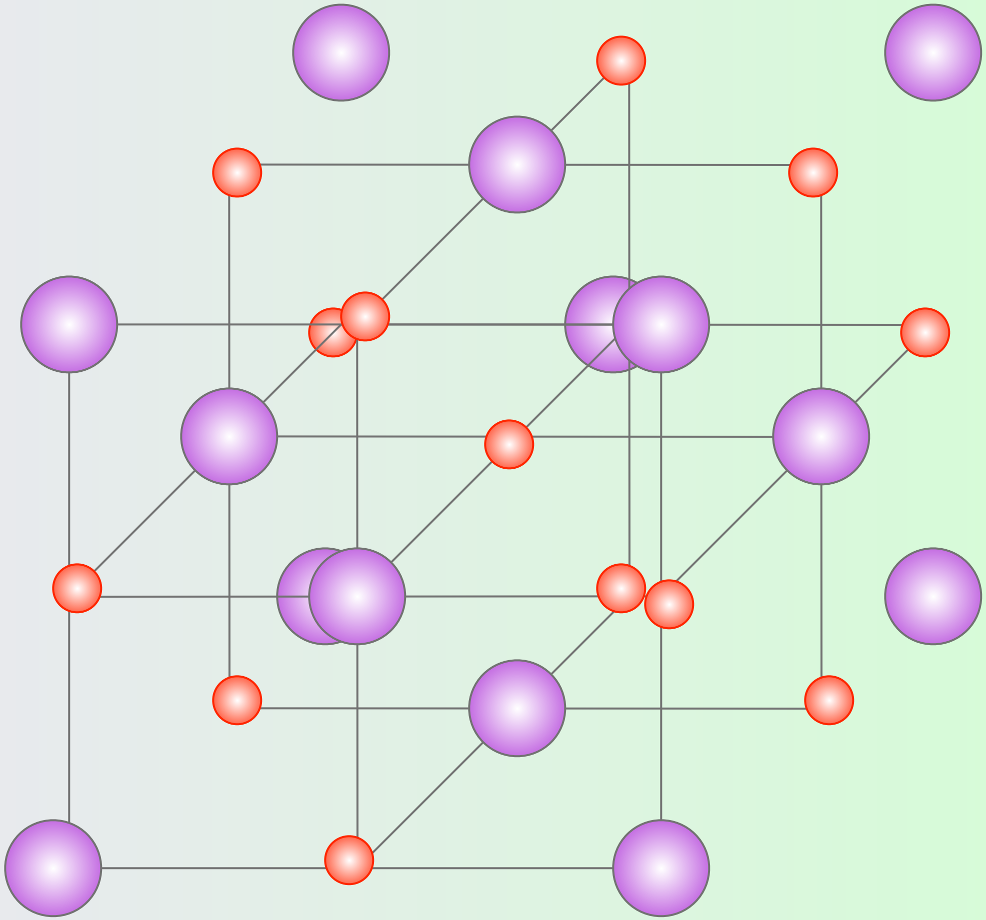
**the mass in grams of one mole of a compound**

# Formula Mass

**not all compounds are molecular**

**calculated exactly the same way as molecular mass**

**Solid structure of NaCl**



## Example

Calculate the number of moles of chloroform ($CHCl_3$) in 198 g of chloroform.

Molecular mass of chloroform:

1 mol C = 12.01 g

1 mol H = 1.008 g

3 mol Cl = 3(35.46 g) = 106.38 g

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1 mol $CHCl_3$ = 119.4 g

$$198 \text{ g } CHCl_3 \times \frac{1 \text{ mol } CHCl_3}{119.4 \text{ g } CHCl_3} = 1.66 \text{ mol } CHCl_3$$

# Example

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

$$64.0\ \cancel{g}\ O_2 \times \frac{1\ \cancel{mol}\ O_2}{32.0\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molec.}\ O_2}{1\ \cancel{mol}\ O_2}$$

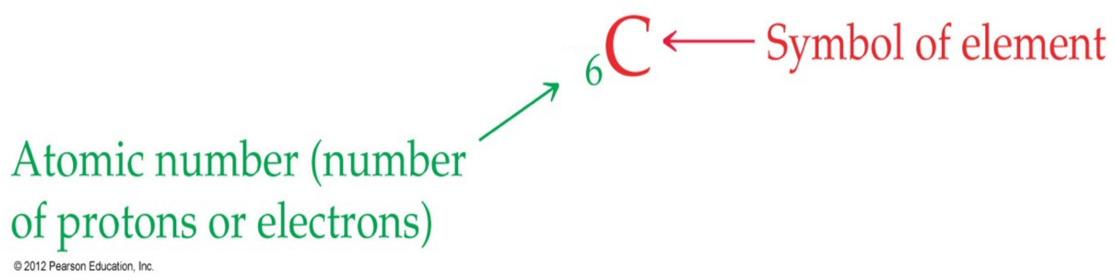
$$= 1.20 \times 10^{24}\ \text{molec.}\ O_2$$

# Symbols of Elements

C ⟵ Symbol of element

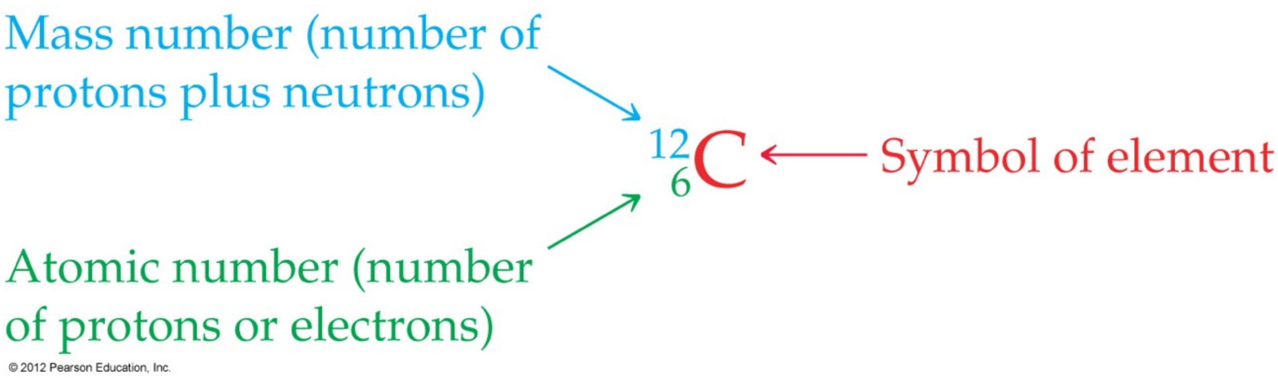
Elements are symbolized by one or two letters.

# Symbols of Elements




All atoms of the same element have the same number of protons, which is called the **atomic number**, *Z*.

# Symbols of Elements




The mass of an atom in atomic mass units (amu) is the total number of protons and neutrons in the atom.

# Isotopes

- **Isotopes** are atoms of the same element with different masses.
  - Isotopes have different numbers of neutrons.

**TABLE 2.2 • Some Isotopes of Carbon***

| Symbol | Number of Protons | Number of Electrons | Number of Neutrons |
|---|---|---|---|
| $^{11}C$ | 6 | 6 | 5 |
| $^{12}C$ | 6 | 6 | 6 |
| $^{13}C$ | 6 | 6 | 7 |
| $^{14}C$ | 6 | 6 | 8 |

*Almost 99% of the carbon found in nature is $^{12}C$.

# 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

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

# Symbols for a few atoms

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Mass number (A)

# Symbols for a few atoms

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

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**

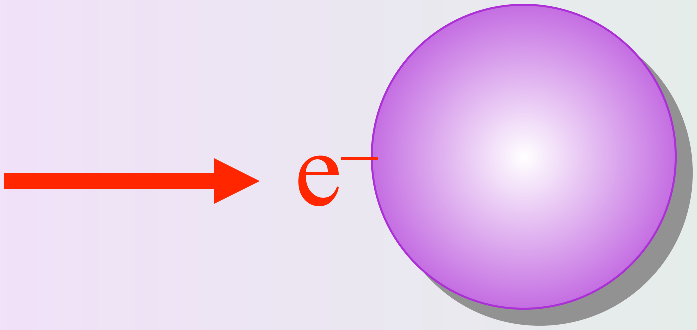
$$\frac{{}^{13}_{6}C}{{}^{12}_{6}C} = 1.0836129$$

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

$$= 13.003355\ \text{amu}$$

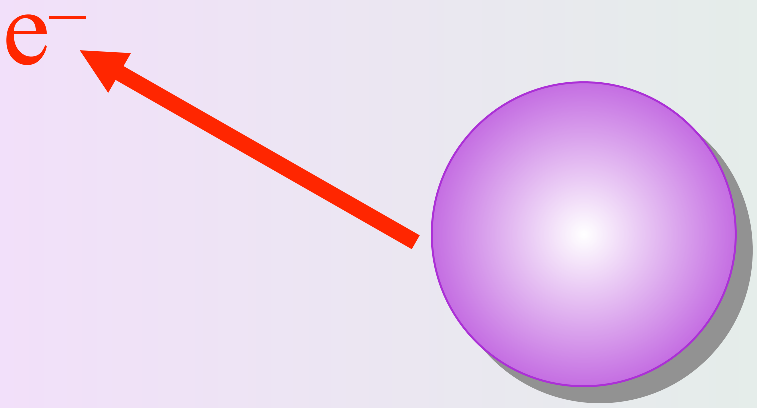
# Experimental Determination Of Atomic & Molecular Masses

# Atomic mass is measured by mass spectrometry



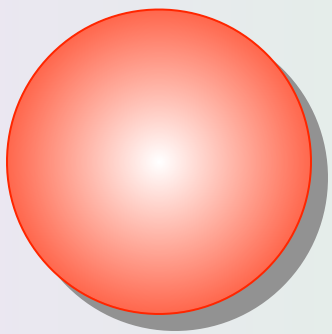

Atom is bombarded by stream of high Energy electrons.

# Atomic mass is measured by mass spectrometry



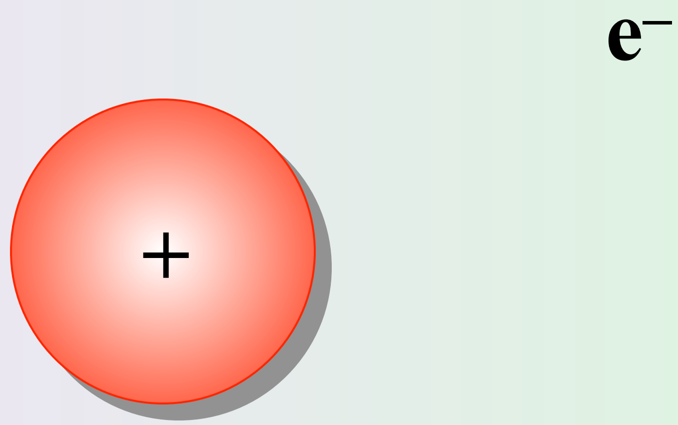

Electron collides with atom, “bounces” off

# Atomic mass is measured by mass spectrometry



and transfers some of its energy to it.

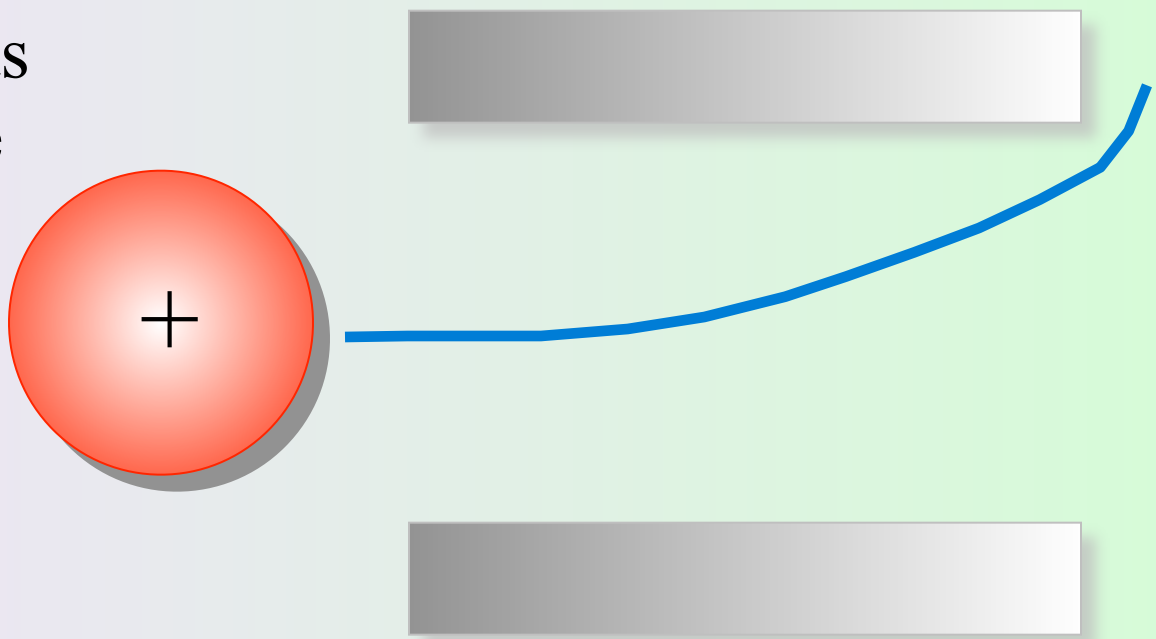
# Atomic mass is measured by mass spectrometry




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

# Ion is deflected by magnetic field

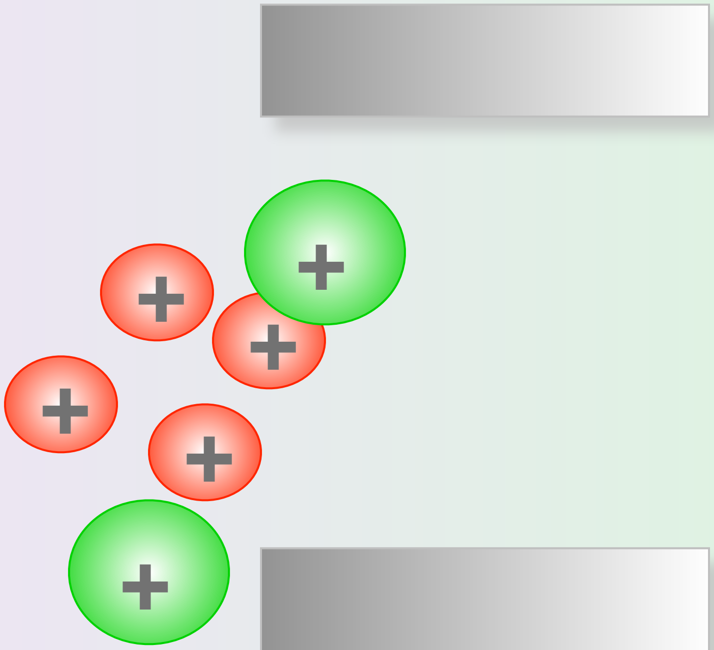
amount of deflection depends on mass to charge ratio




highest m/z deflected least

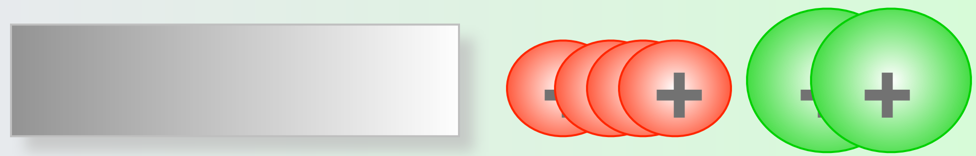
lowest m/z deflected most

# Ions are detected after passage through magnetic field



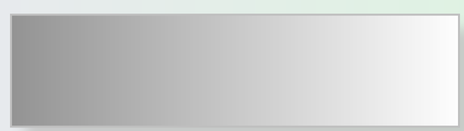
# Ions are detected after passage through magnetic field

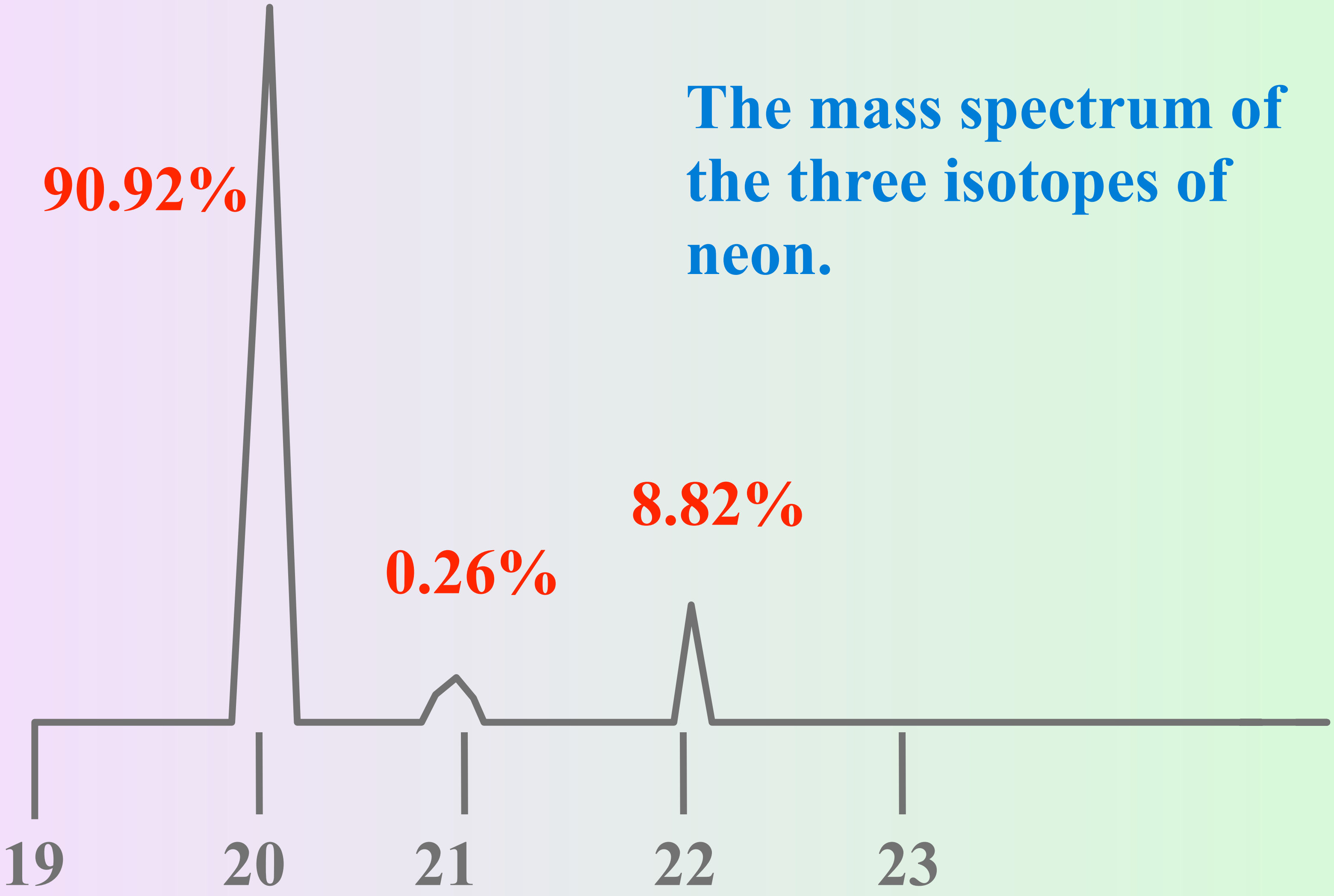
mixture of ions of different mass gives separate peak for each m/z



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

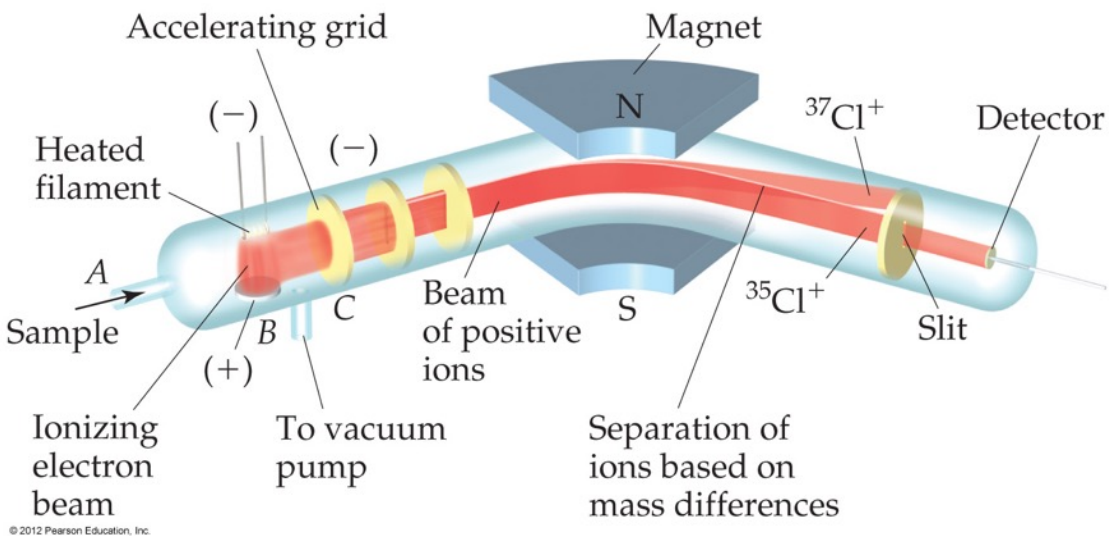
separation of peaks depends on relative mass






The mass spectrum of the three isotopes of neon.

# Atomic Mass




Atomic and molecular masses can be measured with great accuracy using a mass spectrometer.

# 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 |

# Average Mass

- Because in the real world we use large amounts of atoms and molecules, we use average masses in calculations.
- **Average mass** is calculated from the isotopes of an element weighted by their relative abundances.

# 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

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

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

$$= (62.93 \text{ amu})(0.6909)+(64.9278 \text{ amu})(0.3091)$$

$$= 43.47 \text{ amu} + 20.07 \text{ amu}$$

$$= 63.54 \text{ amu}$$