4.1 Structure of the Atom

- How do atoms differ from each other?
- What are atoms composed of?
- What are the subatomic particles?

Structure of the Atom

- Atoms actually are divisible. They are composed of subatomic particles.
- Subatomic particles include:
  - One kind of particle found outside the nucleus
    - Electrons - negatively charged subatomic particles
  - Two kinds of particles found in the nucleus (center of the atom)
    - Protons - positively charged subatomic particles
    - Neutrons - uncharged subatomic particles
- How did we find out about these particles?

The Discovery of Electrons

- J.J. Thomson in 1897 conducted a series of experiments with cathode ray tubes showed that matter contains negatively charged particles.
- The beam of particles is attracted to the positively charged plate.
- This streams of small negatively charged particles called *electrons*.

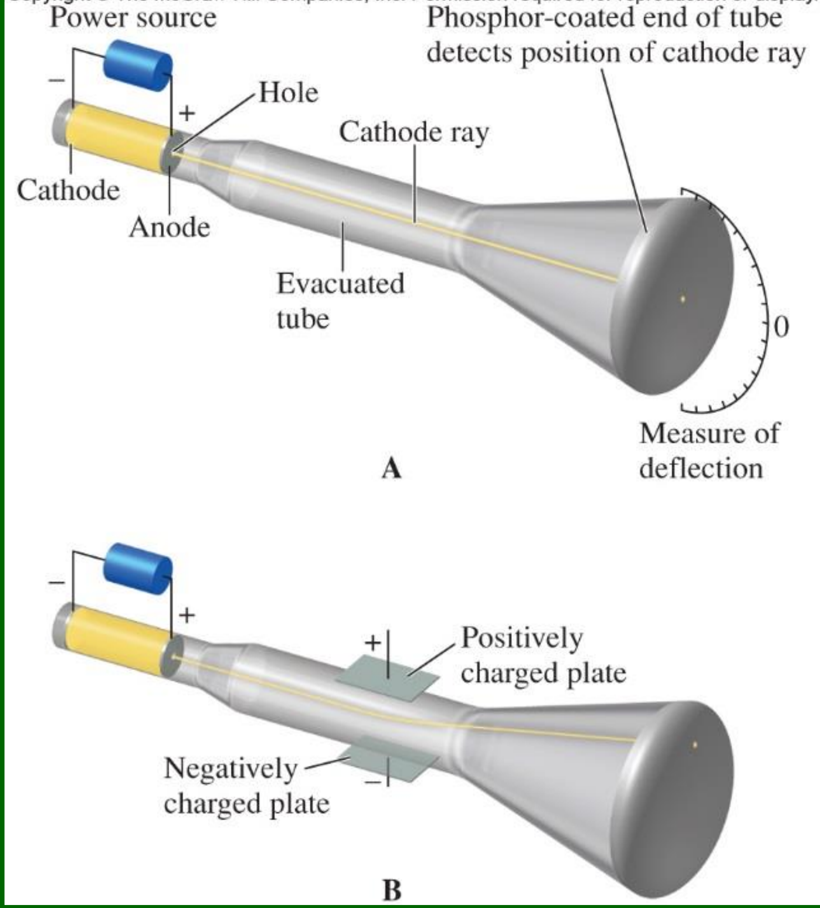


Figure 2.5

Determination of Mass and Charge of the Electron

- In 1909 Millikan's oil-drop experiment, the electric field strength required to suspend an oil droplet was dependent upon the number of extra electrons on it.
- This allowed the determination of the electron's charge (-1.6022×10^{-19} C) and mass (9.1094×10^{-28} g).

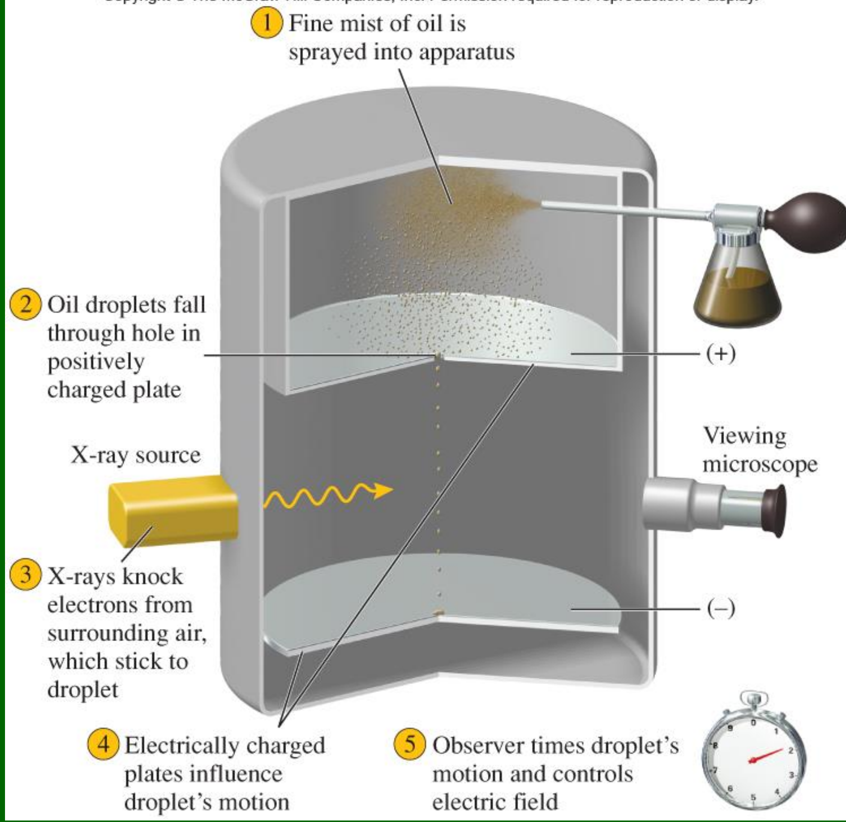


Figure 2.6

The Proton

- Scientists reasoned that if atoms have negatively charged particles, they must also have positively charged particles (protons).
- From his experiments with electrons, J.J. Thomson proposed that electrons might be embedded in a sphere of positive charge ("plum pudding" model of the atom).

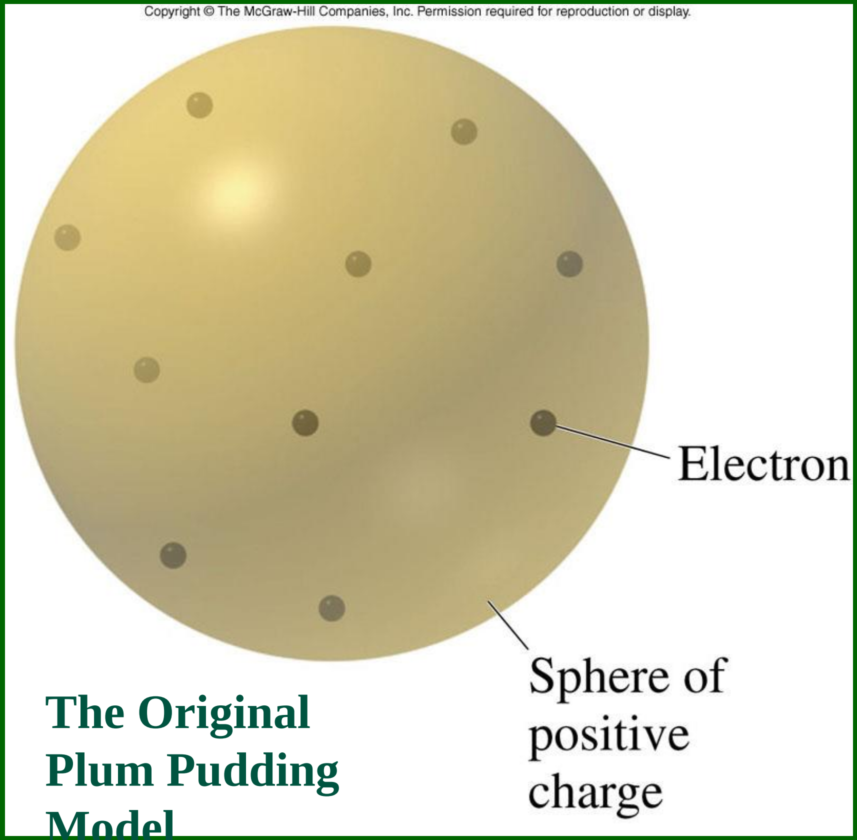
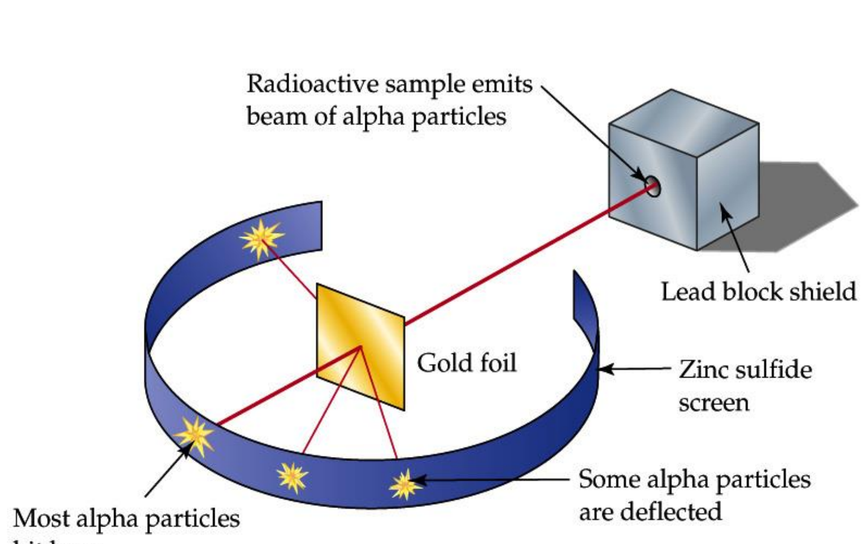


Figure 2.7

Rutherford's Scattering Experiment

1907 Ernest Rutherford designed an experiment to test J.J. Thomson's "plum pudding" model of the atom
~ use α particle to study the inner structure of atoms. When he directed a beam of α-particles at a thin gold foil, he found that

- The majority of α-particles penetrated the foil undeflected.
- Some α particles experienced slightly deflections.
- A few (about one in every 20,000) suffered rather serious deflections as they penetrated the foil.
- A similar number did not pass through the foil at all, but bounced back in the direction from which they had come.



(a)

Rutherford's Gold Foil Experiment

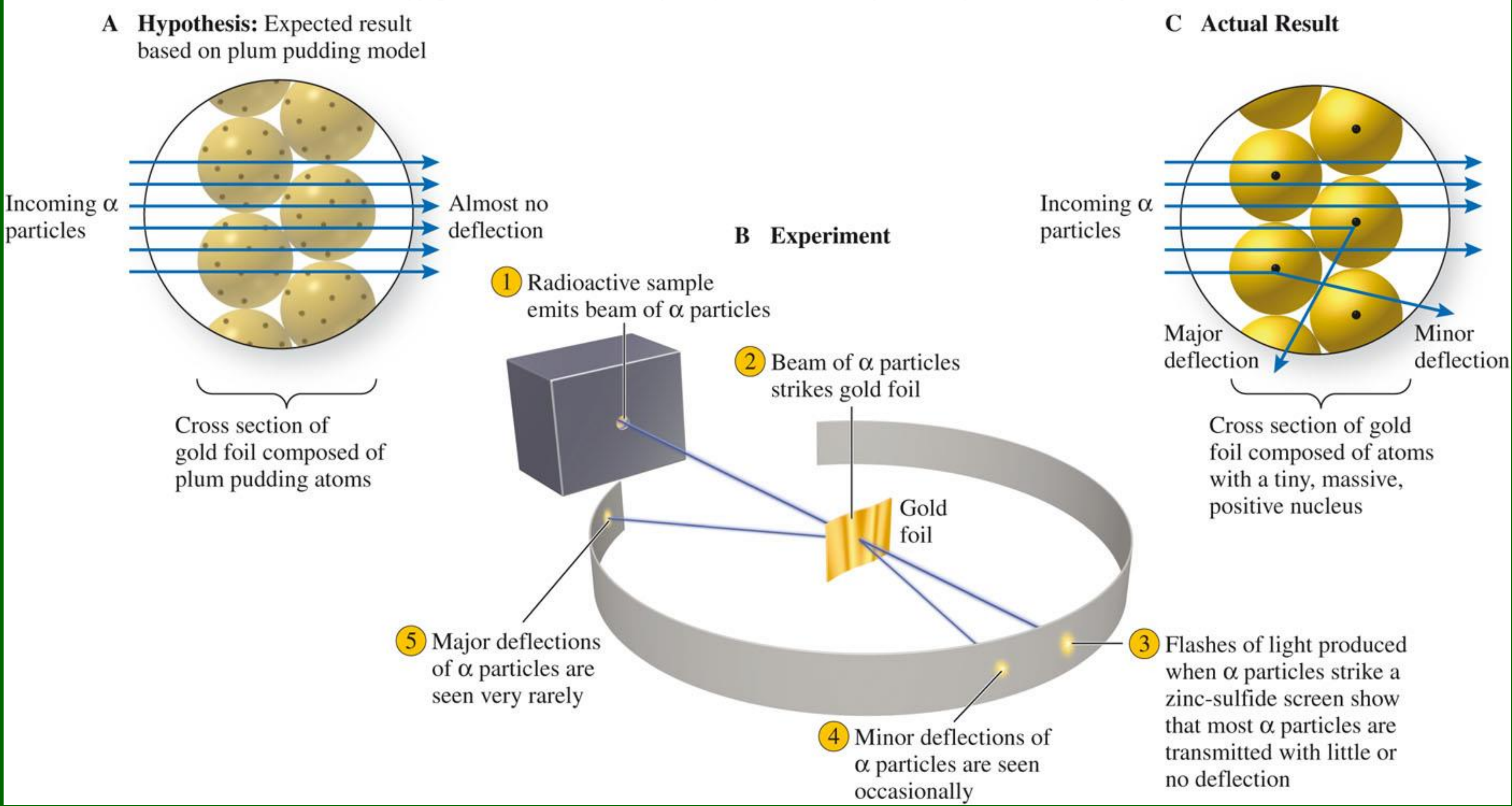


Figure 2.8

The Nuclear Atom: Protons and Neutrons

1911 Rutherford explained his results by proposing a model of the atom known as the nuclear atom and having these features.

1. Most of the mass and all of the positive charge of an atom are centered in a very small region called the nucleus. The atom is mostly empty space.
2. The magnitude of the positive charge is different for different atoms and is approximately one-half the atomic weight of the element.
3. There are as many electrons outside the nucleus as there are units of positive charge on the nucleus. The atom as a whole is electrically neutral.

Rutherford's nuclear atom suggested the existence of positively charged fundamental particles of matter in the nuclei of atoms- called **protons**.

The Neutron

- Because the protons in the atom could account for only about half the mass of most atoms, scientists knew there was another heavy particle in the nucleus.
- Neutrons were proposed by Ernest Rutherford in 1907 (to account for a mass discrepancy in the nucleus) and discovered in 1932 by James Chadwick.
- The neutron has about the same mass as a proton but with no charge.

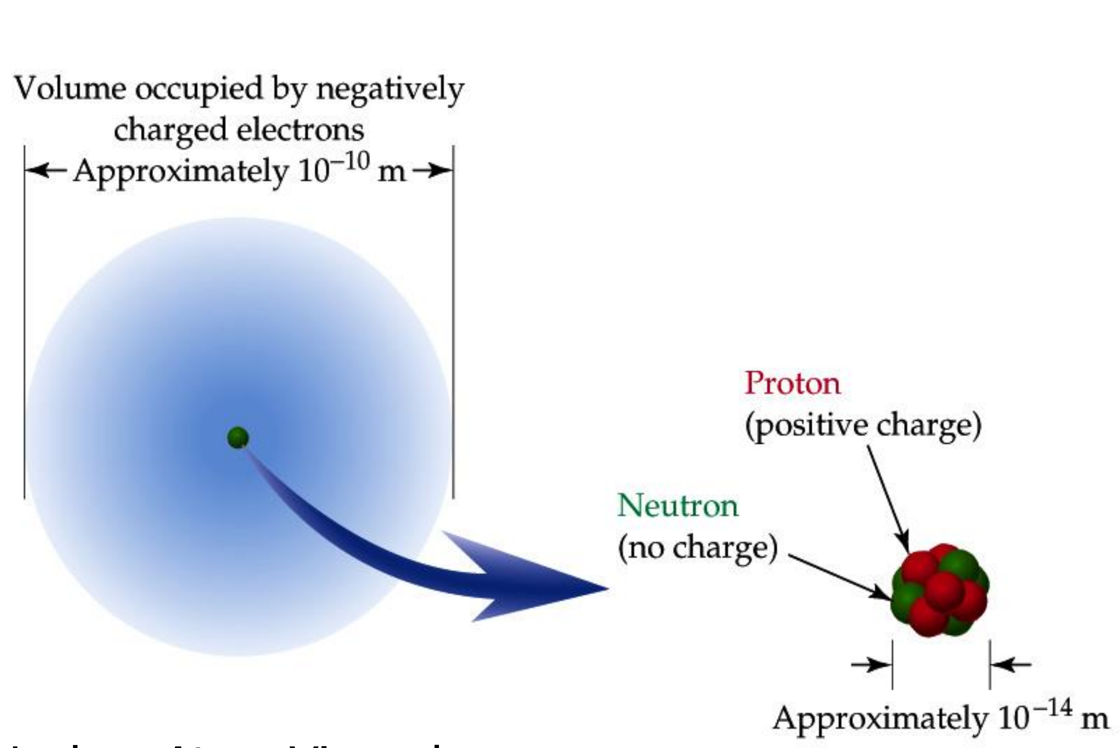
Summary of Subatomic Particles

TABLE 4.5 Subatomic Particles in the Atom

Particle	Symbol	Relative Charge	Mass (g)	Mass (amu)	Location in Atom
Proton	p or p^+	1+	1.673×10^{-24}	1.007	Nucleus
Neutron	n or n^0	0	1.675×10^{-24}	1.008	Nucleus
Electron	e^-	1−	9.110×10^{-28}	0.000 55	Outside nucleus

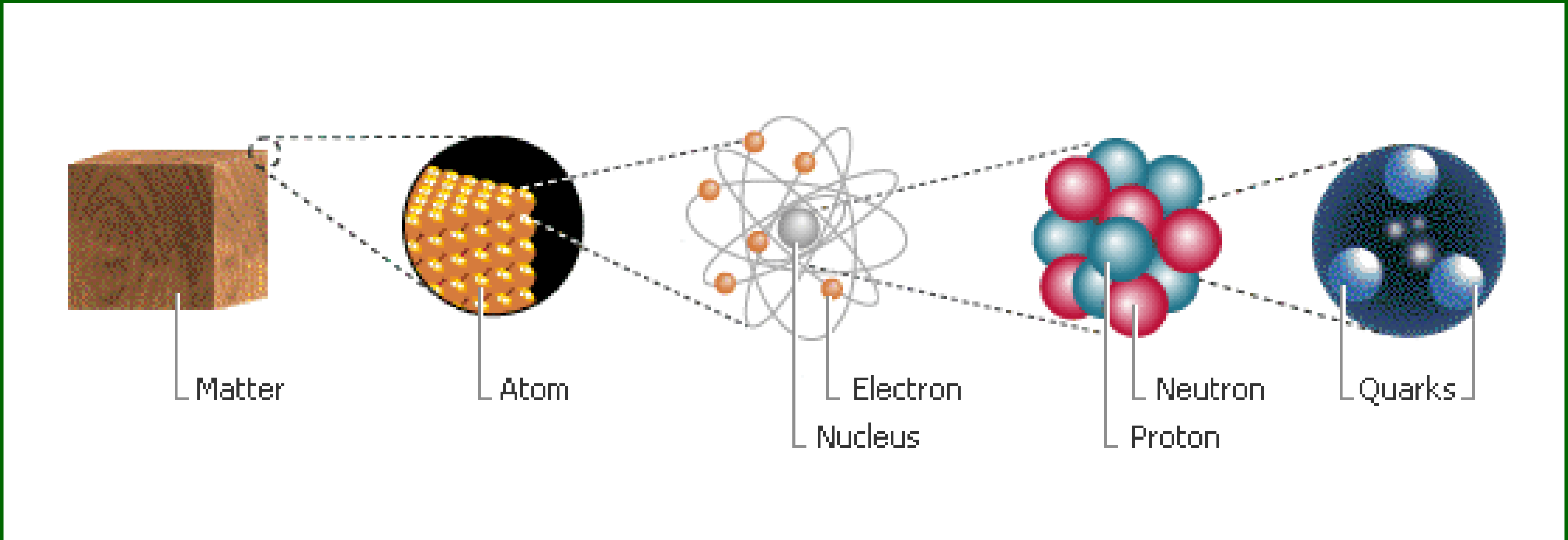
The Modern View of Atomic Structure

- Protons and neutrons form a compact central body- Nucleus which contains most of the mass of the atom
- Electrons are distributed in space like a cloud around the nucleus
- The electron moves about the nucleus in such a way that most of the volume of the atom is empty.
- Atoms have a zero charge.



Nuclear Atom Viewed in Cross Section

Conclusion



- Modern physics has revealed successively deeper layers of structure in ordinary matter. Matter is composed, on a tiny scale, of particles called atoms. Atoms are in turn made up of minuscule nuclei surrounded by a cloud of particles called electrons. Nuclei are composed of particles called protons and neutrons, which are themselves made up of even smaller particles called quarks. Quarks are believed to be fundamental, meaning that they cannot be broken up into smaller particles.

Chemical Elements

- If all atoms are composed of same subatomic particles (electrons, neutrons and protons), why do different atoms have different chemical properties?
- What is that makes one atom different from another?
- **Elements differ from one another according to the number of protons.** All atoms of an element have the same number of protons.
- The number of protons in an atom of an element is called its atomic number

Atomic Number and Mass Number

- Atomic Number (Z)
  - Unique to each element
  - Determine by the number of protons in the nucleus of an element’s atom
- Mass Number (A)
  - the number of protons and neutrons in the nucleus of an element’s atom

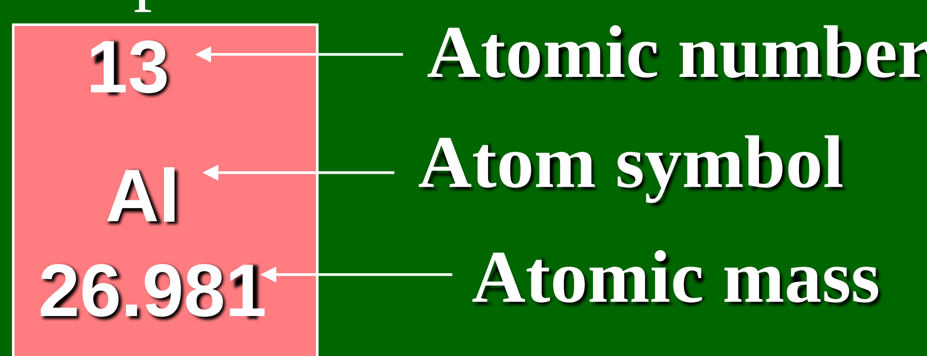
$$A = Z + N$$

Atomic Number, Z and Mass number, A

- Atomic Number (*Z*)
  - Unique to each element
  - Determine by the number of protons in the nucleus of an element's atom
- Mass Number (*A*)
  - the number of protons and neutrons in the nucleus of an element's atom

$$A = Z + N$$

- All atoms of particular element have the same atomic number, Z, on the other hand all atoms with the same number of protons are the same element.



Isotopes

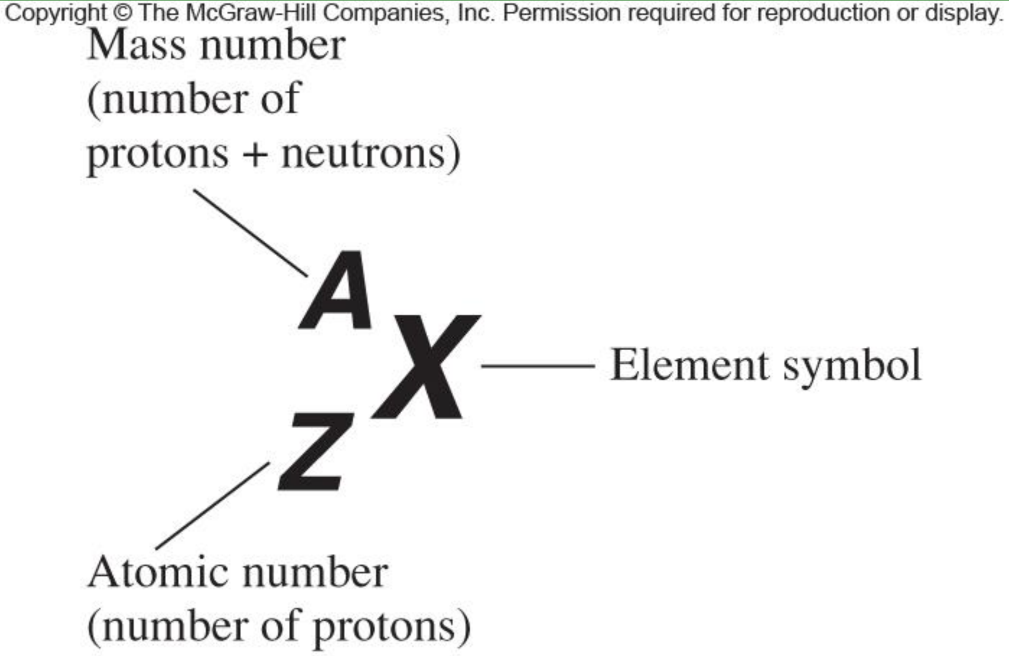
- Contrary to what Dalton thought, we know that atoms of an element do not necessarily all have the same mass.
- An isotope of an element
  – Atoms of the same element (same Z) that have different mass number (A).
  – is an atom that contains a specific number of neutrons.
  – Most elements occur naturally as a mixture of different isotopes.
  – The naturally occurring percentages of isotopes of a particular element are referred to as the percent natural abundance of that element.

Isotope Symbol

- An isotope symbol
  - is a common notation that represents the mass number, atomic number, and elemental symbol.
    - The subscript in the isotope symbol is the atomic number.
    - The superscript in the isotope symbol is the mass number.
- Representations for a boron atom with 6 neutrons:

$${}^{11}_{5}\mathrm{B}$$

- boron-11



Isotopes

- An isotope of an element is an atom that contains a specific number of neutrons.
- There are three isotopes of hydrogen:

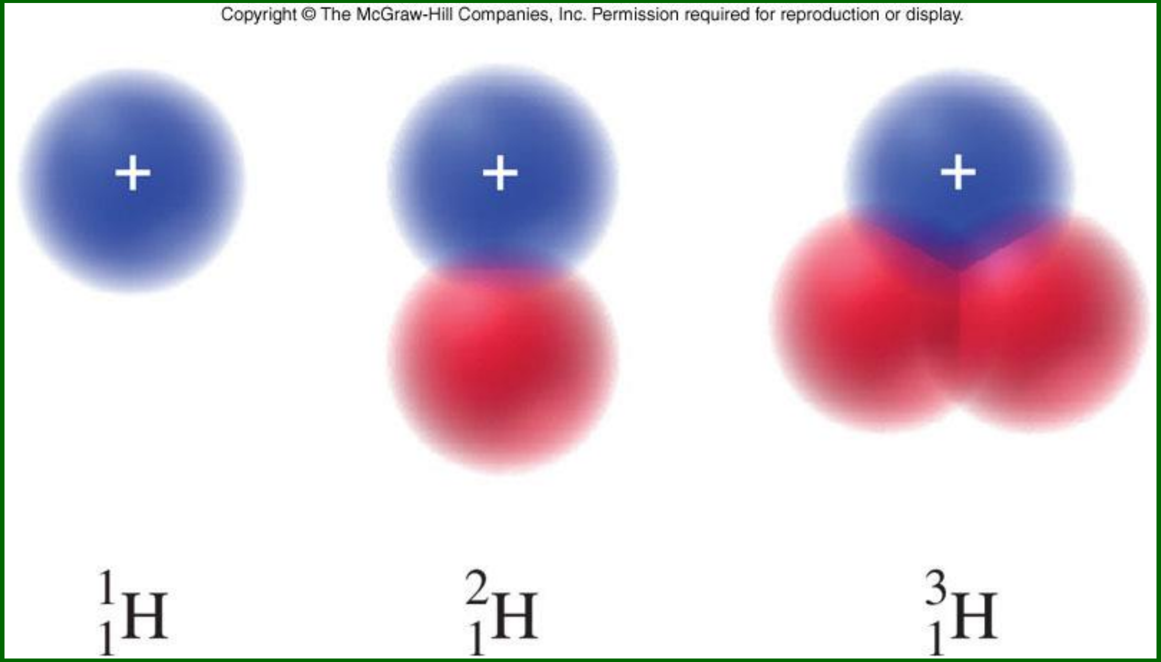
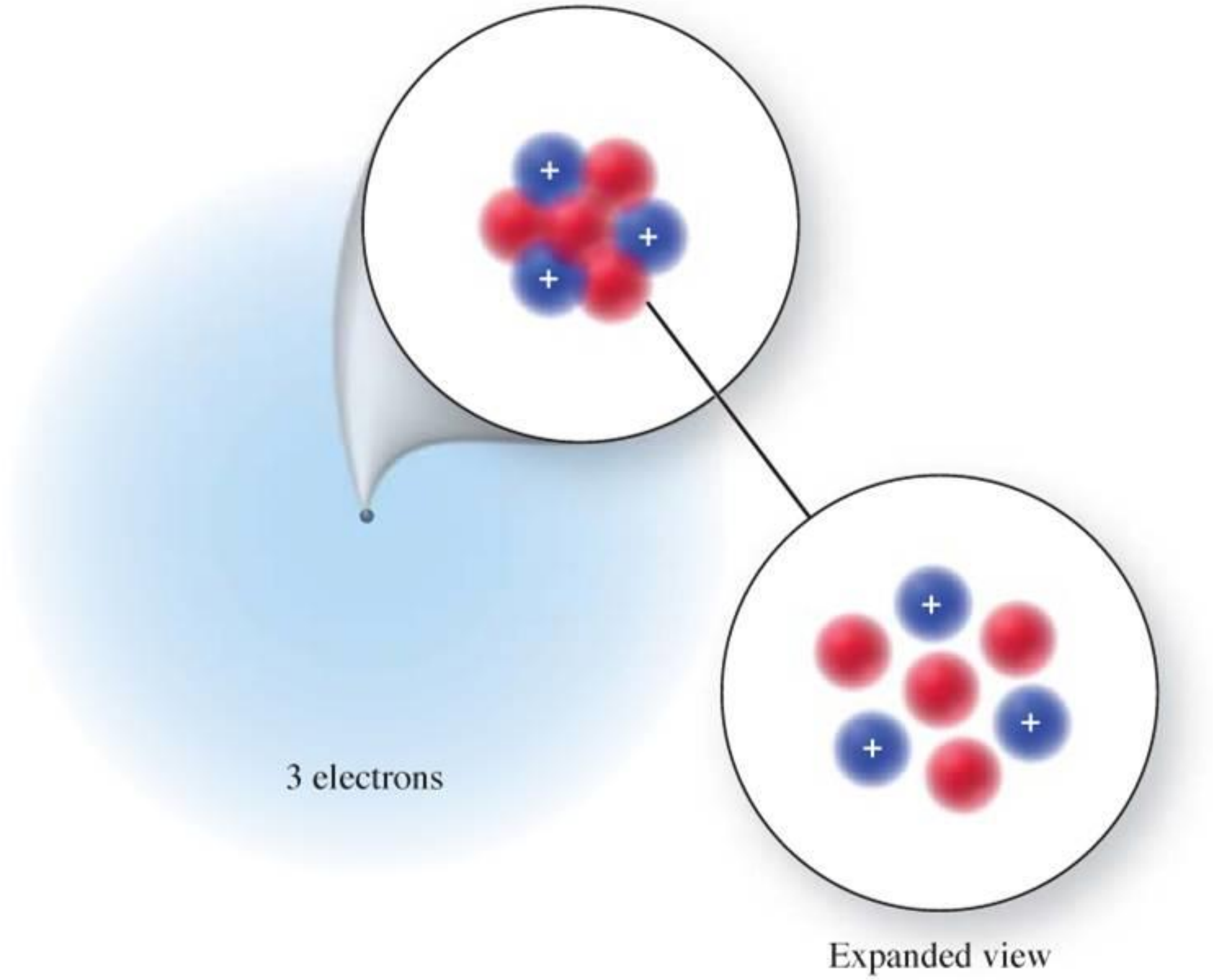


Figure 2.11

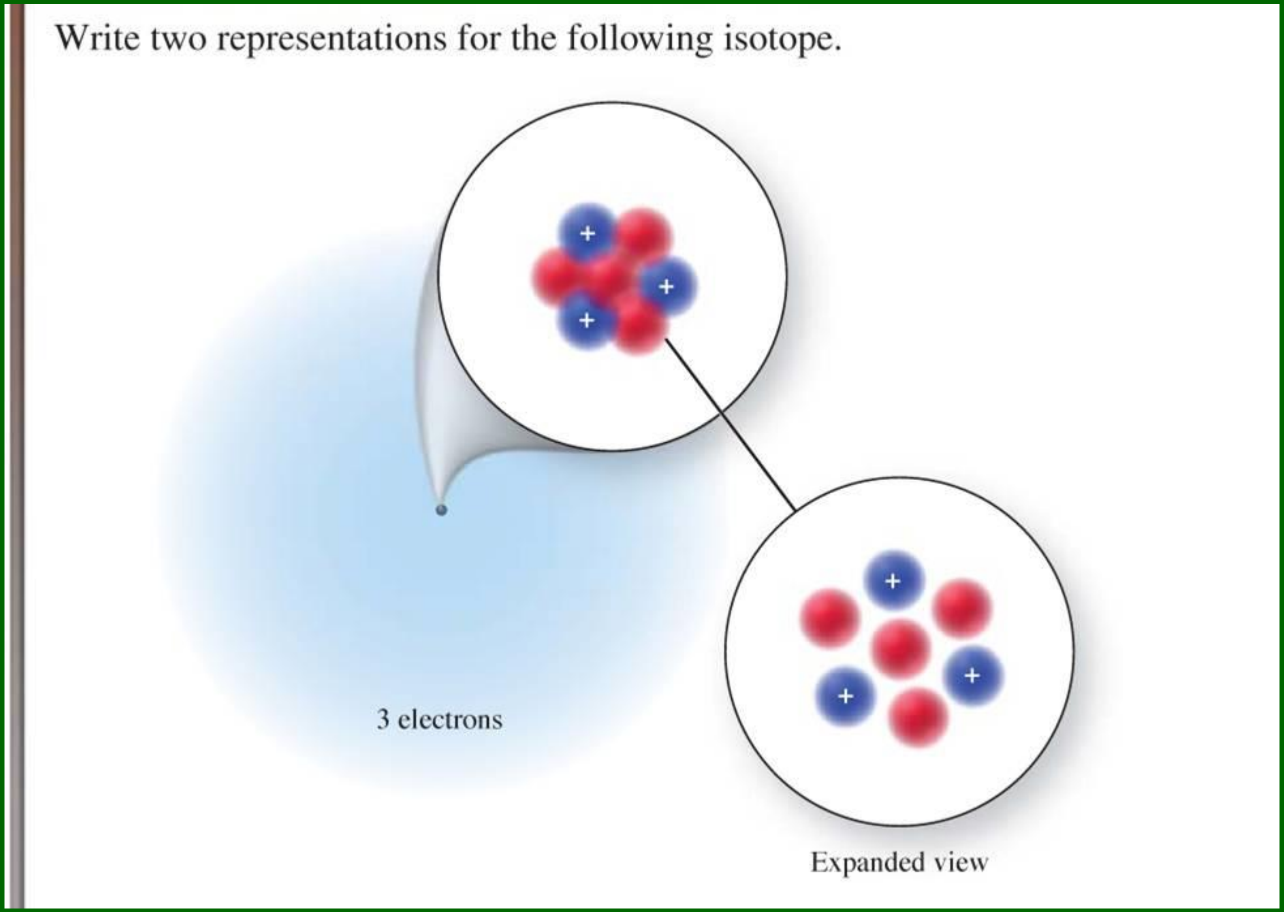
Activity: Writing Isotope Symbols

Write two representations for the following isotope.



Activity Solution:

- $Z = 3$
- $N = 4$
- Element = Li
- $A = 3 + 4 = 7$
- Symbol = ${}^{7}_{3}\mathrm{Li}$
  or lithium-7



Activity: Isotopes

- Determine the number of protons and neutrons in each of the following isotopes:

$^{35}_{17}\mathrm{Cl}$ **17 protons, 18 neutrons**

carbon-11 **6 protons, 5 neutrons**

$^{198}\mathrm{Au}$ **79 protons, 119 neutrons**

Activity: Isotope Symbols

- Practice writing the isotope symbols for the following isotope pairs.

1. carbon-13 and carbon-14
2. chlorine-35 and chlorine-37
3. uranium-235 and uranium-238
4. lithium-6 and lithium-7

Solution

1. carbon-13 and carbon-14

$${}^{13}_{6}\mathrm{C} \text{ and } {}^{14}_{6}\mathrm{C}$$

Activity Solutions: Isotope Symbols

2. chlorine-35 and chlorine-37

$$^{35}_{17}\mathrm{Cl} \text{ and } ^{37}_{17}\mathrm{Cl}$$

3. uranium-235 and uranium-238

$$^{235}_{92}\mathrm{U} \text{ and } ^{238}_{92}\mathrm{U}$$

4. lithium-6 and lithium-7

$$^{6}_{3}\mathrm{Li} \text{ and } ^{7}_{3}\mathrm{Li}$$

Heavy Water

- One ice cube is made with water that contains only the hydrogen-2 isotope. The other ice cube is composed of water with normal water which contains mostly hydrogen-1.
- Which is which?

Figure 2.13

2.3 Ions

- Dietary sodium and sodium found in the sea and in the earth's crust is in the form of sodium ions.
- A sodium ion has a net +1 charge: Na^+
  – It has one less electron than the number of protons.
- How many protons and electrons does a sodium ion, Na^+, have?

Ions

- an electrically charged particle obtained from an atom or chemically bonded group of atoms lose or gain electrons. The charge on an ion is equal to the numbers of protons minus the numbers of electrons.
- Cations are positively charged. They have fewer electrons than in the neutral atom.
- Anions are negatively charged. They have more electrons than in the neutral atom.

number p + number n

$$^{A}_{Z}E^{\pm ?}$$

number p - number e

number p

Formation of Mg^{2+} Cation

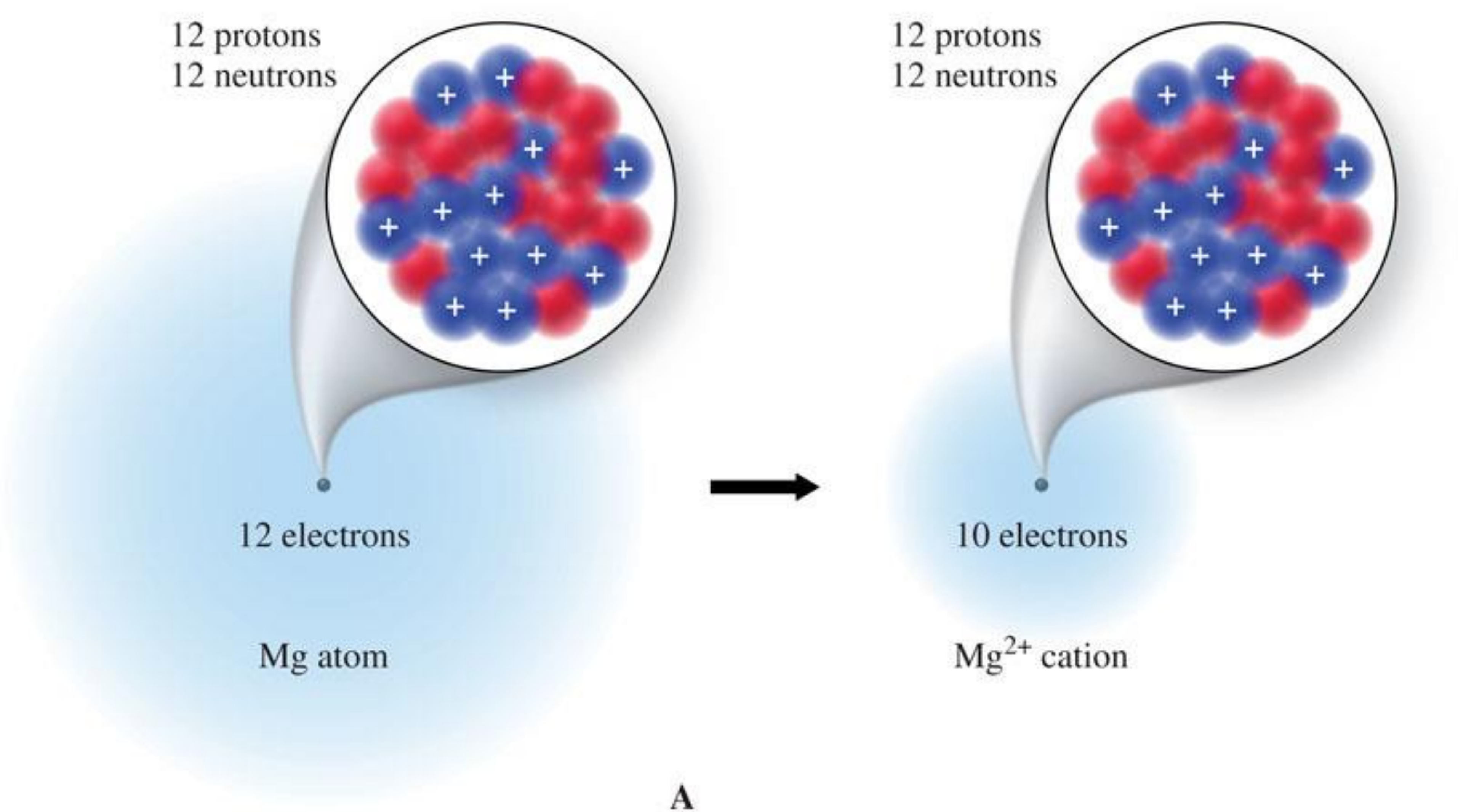


Figure 2.14

Formation of N^{3-} Anion

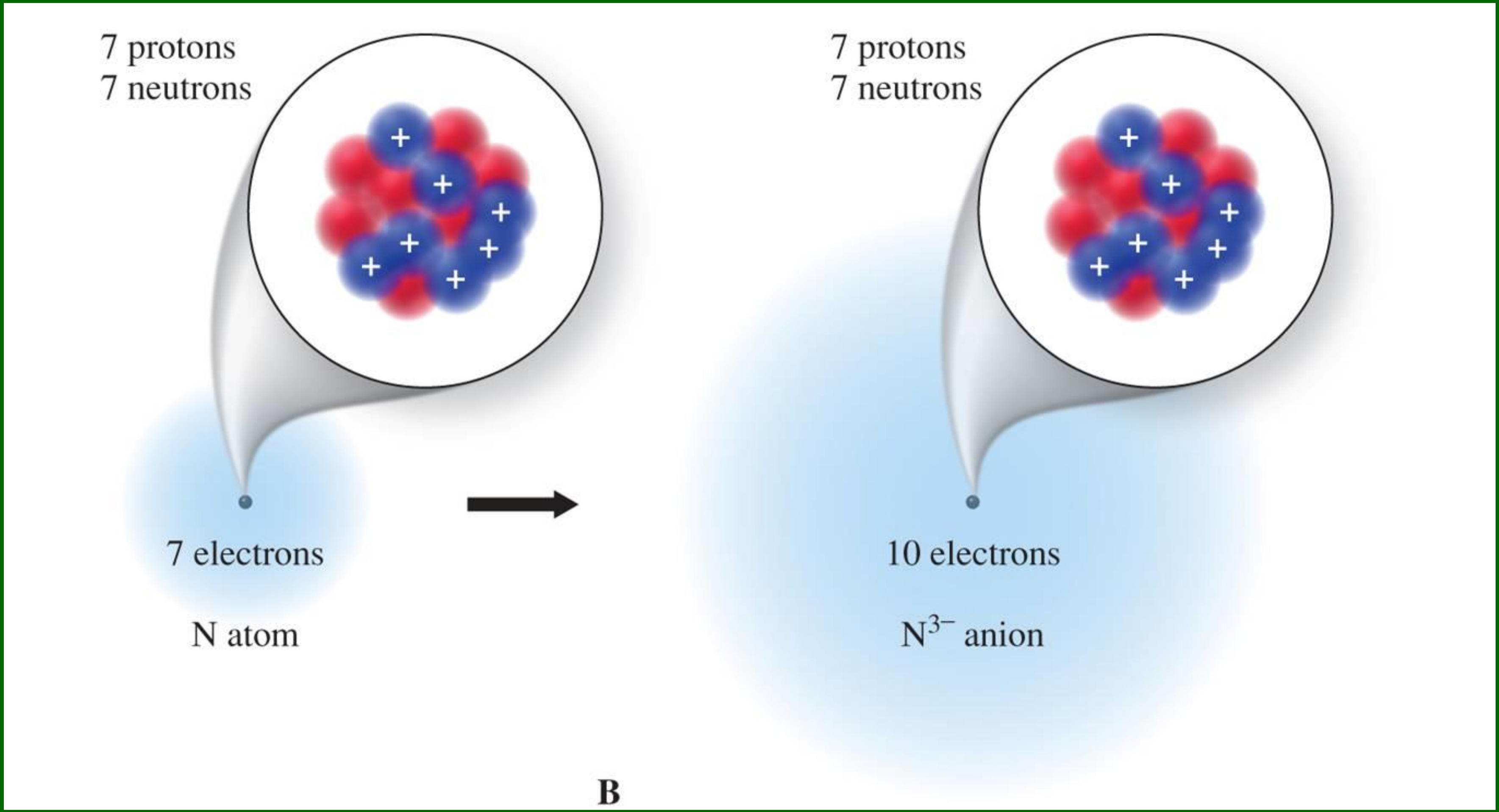
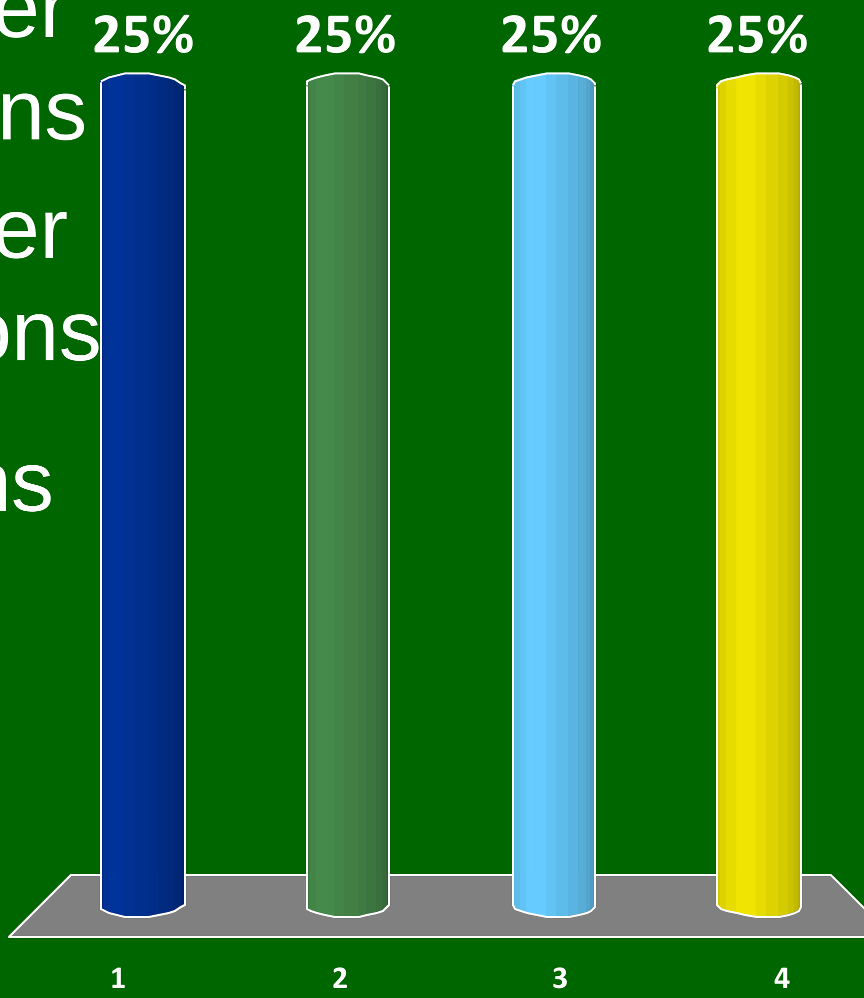


Figure 2.14

What does the mass number refer to in an atom?

A. The sum of the number of protons and neutrons
B. The sum of the number of protons and electrons
C. The number of protons
D. The number of electrons



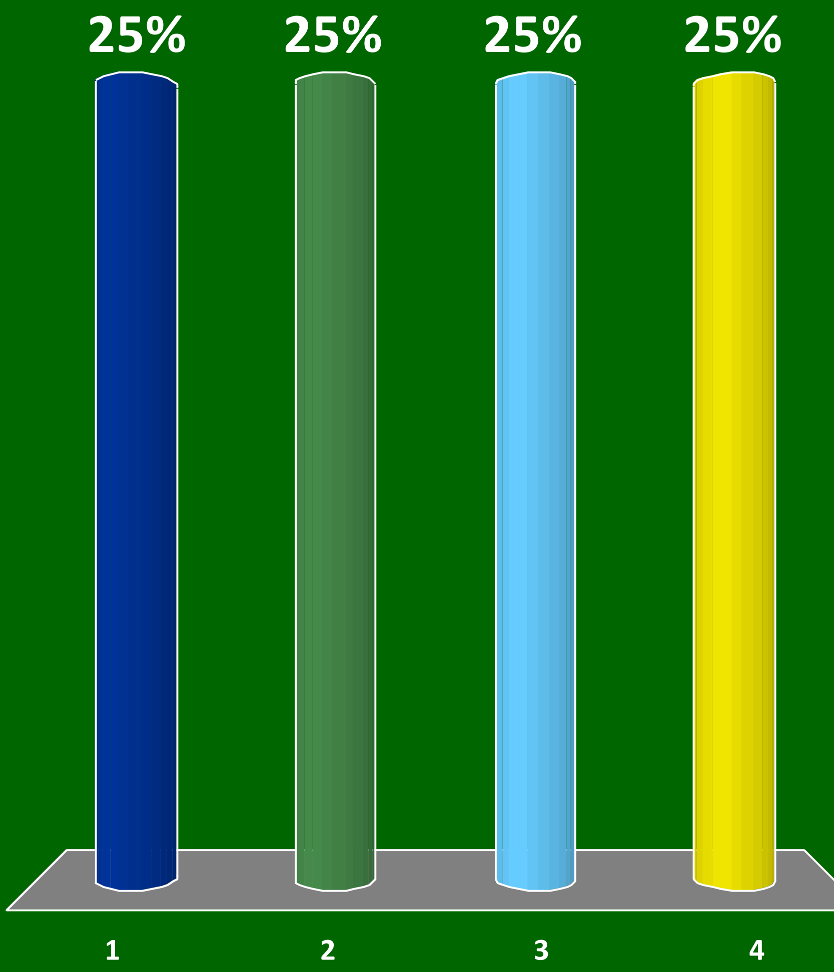
How many protons, neutrons, and electrons are in ^{70}Ge?

A) 32 protons, 32 neutrons, 38 electrons

B) 32 protons, 70 neutrons, 70 electrons

C) 32 protons, 38 neutrons, 32 electrons

D) 70 protons, 32 neutrons, 38 electrons



Solution:

protons = atomic # =32

neutrons = mass # - # protons = 70-32 =38

electrons for a neutral species = # protons

Activity: Counting Protons and Electrons

- Write the number of protons and electrons for the following ions:

1. Na^{+}
2. Cl^{-}
3. O^{2-}
4. Al^{3+}

Activity Solutions: Counting Protons and Electrons

- Write the number of protons and electrons for the following ions:
  1. Na^+

     Sodium has an atomic number equal to 11. Thus, it has 11 protons.

     11- number of e- = +1

     number of e- = 10
  2. Cl^-

     Chlorine has an atomic number equal to 17. It has 17 protons.

     17 – number e- = -1

     number e- = 18.

Activity Solutions: Counting Protons and Electrons

- Write the number of protons and electrons for the following ions:

3. O^{2-}

   Oxygen has an atomic number equal to 8. It has 8 protons.

   8 – number e- = -2

   Number e- = 10.

4. Al^{3+}

   Aluminum has an atomic number equal to 13. It has 13 protons.

   13 – number e- = +3

   Number e- = 10

Activity: Ions

- Write the symbol for the ion that has the following number of protons and electrons:

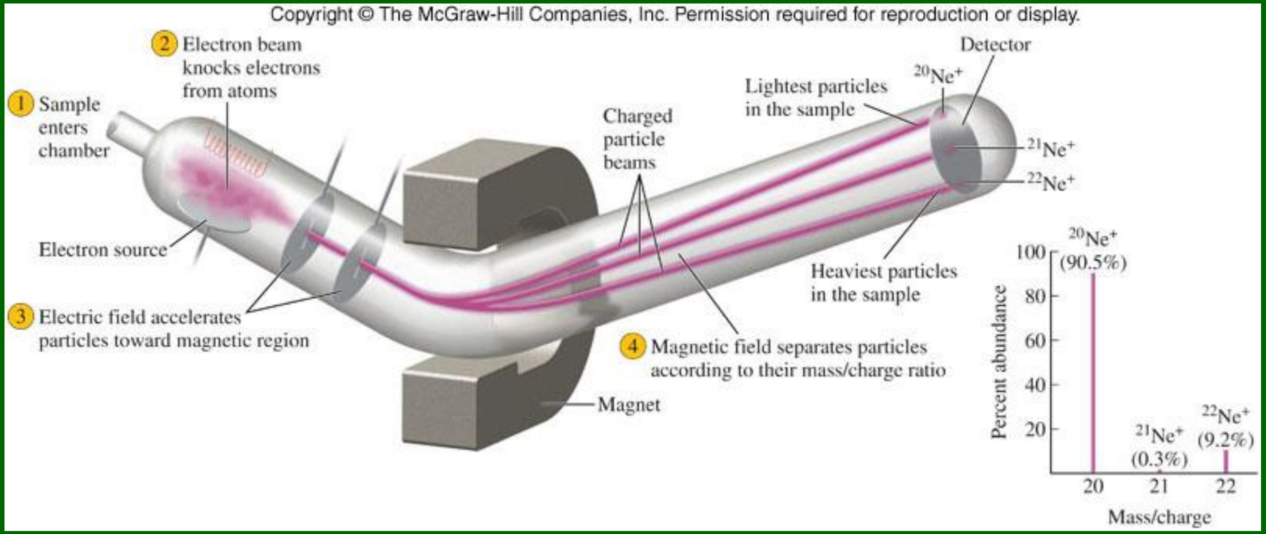
1. 20 protons and 18 electrons **Ca^{2+}**
2. 16 protons and 18 electrons **S^{2-}**
3. 26 protons and 23 electrons **Fe^{3+}**

2.4 Atomic Mass

- We can use mass number to compare the approximate relative masses of different isotopes.
  - A carbon-12 atom is about twelve times the mass of a hydrogen-1 atom.
  - An oxygen-16 atom is about four times the mass of a helium-4 atom.
- But keep in mind that the mass number is NOT an actual mass!

Determination of Atomic Mass

- The individual masses of the isotopes of an element can be determined by mass spectrometry.
- This technique is also used to determine the relative amounts of the isotopes in a sample of an element.



AMU Scale

- The mass of a carbon-12 atom is 1.99272×10^{-23} g. Because this is such a small number, the atomic mass unit (amu) scale was developed and based on the mass of carbon-12:

$$1 \text{ amu} = \frac{\text{Mass of one } ^{12}\text{C atom}}{12} = 1.660539 \times 10^{-24} \text{ g}$$

- This means that carbon-12 has a mass of exactly 12 amu.

Example

1) From the mass spectral data, the ratio of the mass of ^{16}O to ^{12}C is found to be 1.33291. What is the mass of an ^{16}O atom?

$^{16}O/^{12}C = 1.33291$

mass of ^{16}O = 1.33291 x 12amu = 15.9949amu

2) From precise measurement show that the mass of ^{10}B is 0.83442 times the mass of ^{12}C. What is the isotope mass of ^{10}B?

mass of ^{10}B = 0.83442 x 12 amu

= 10.013 amu

Atomic Masses

- Elements occur in nature as mixtures of isotopes Each naturally occurring isotope of an element contributes a certain portion to the atomic mass
- Weighted averages of the isotopic masses according to their natural abundance

Atomic mass of an element

$$= \left(\begin{array}{l}\textbf{natural}\\ \textbf{abundance}\\ \textbf{of isotope(1)}\end{array} \times \begin{array}{l}\textbf{mass of}\\ \textbf{isotope (1)}\end{array}\right) + \left(\begin{array}{l}\textbf{natural}\\ \textbf{abundance}\\ \textbf{of isotope(2)}\end{array} \times \begin{array}{l}\textbf{mass of}\\ \textbf{isotope (2)}\end{array}\right) + \left(\ldots\right.$$

- Abundance (Percent Abundance)
  - Relative proportions (%) in which the isotopes of an element are found in natural sources

Example

- An unknown element (X) discovered on a planet in another galaxy was found to exist as two isotopes. Their atomic masses and percent abundances are listed in the following table. What is the relative atomic mass of the element?

Isotope	Mass (amu)	Natural Abundance (%)
^{22}X	21.995	75.00
^{20}X	19.996	25.00

Solution

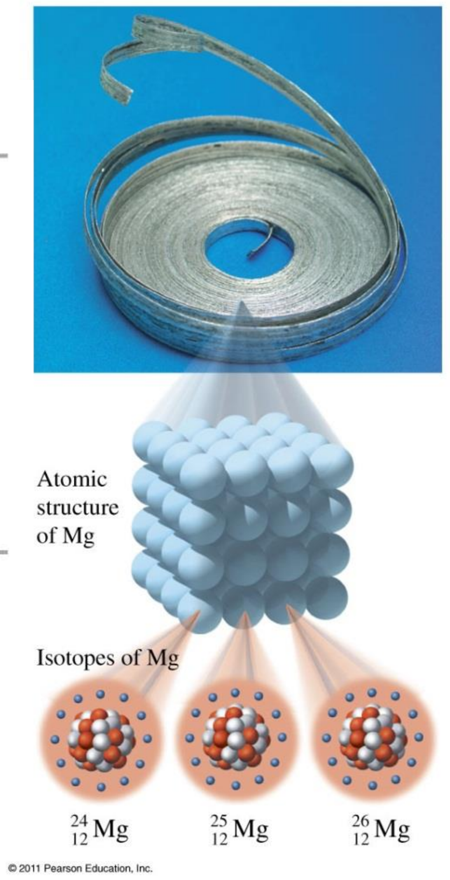
$$(21.995 \times 0.7500) + (19.996 \times 0.2500) = 21.50 \text{ amu}$$

What is the atomic mass of Magnesium ?

In naturally occurring magnesium, there are three isotopes.

TABLE 4.7 Isotopes of Magnesium

Atomic symbol	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{26}_{12}Mg$
Number of protons	12	12	12
Number of electrons	12	12	12
Mass number	**24**	**25**	**26**
Number of neutrons	**12**	**13**	**14**
Mass of isotope (amu)	**23.99**	**24.99**	**25.98**
% abundance	78.70%	10.13%	11.17%



Calculating Atomic Mass

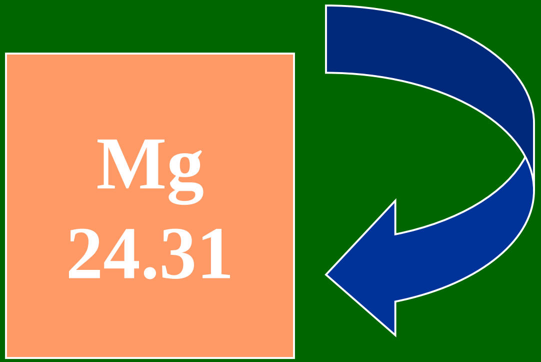
Isotope Mass Abundance

^{24}Mg = 23.99 amu x 78.70/100 = 18.88 amu

^{25}Mg = 24.99 amu x 10.13/100 = 2.531 amu

^{26}Mg = 25.98 amu x 11.17/100 = <u>2.902 amu</u>

Atomic mass (average mass) Mg = 24.31 amu



Radioactivity

Nuclear reactions can change one element into another element. In the late 1890s, scientists noticed some substances spontaneously emitted radiation, a process they called **radioactivity**. The rays and particles emitted are called **radiation**. A reaction that involves a change in an atom's nucleus is called a **nuclear reaction**.

Radioactive Decay

Unstable nuclei lose energy by emitting radiation in a spontaneous process called **radioactive decay**. Unstable radioactive elements undergo radioactive decay thus forming stable nonradioactive elements.

There are three types of radiation: alpha, beta, and gamma

Alpha Radiation

Alpha radiation is made up of positively charged particles called **alpha particles**. Each alpha particle contains two protons and two neutrons and has a 2+ charge. The figure shown below is a nuclear equation showing the radioactive decay of radium-226 to radon-222.

$$\begin{array}{ccccc} ^{226}_{88}\mathrm{Ra} & \rightarrow & ^{222}_{86}\mathrm{Rn} & + & \alpha \\ \text{radium-226} & & \text{radon-222} & & \text{alpha particle} \end{array}$$

An alpha particle is equivalent to a helium-4 nucleus and is represented by $^{4}_{2}\mathrm{He}$ or . Thus, showing mass is conserved in a nuclear equation.

α

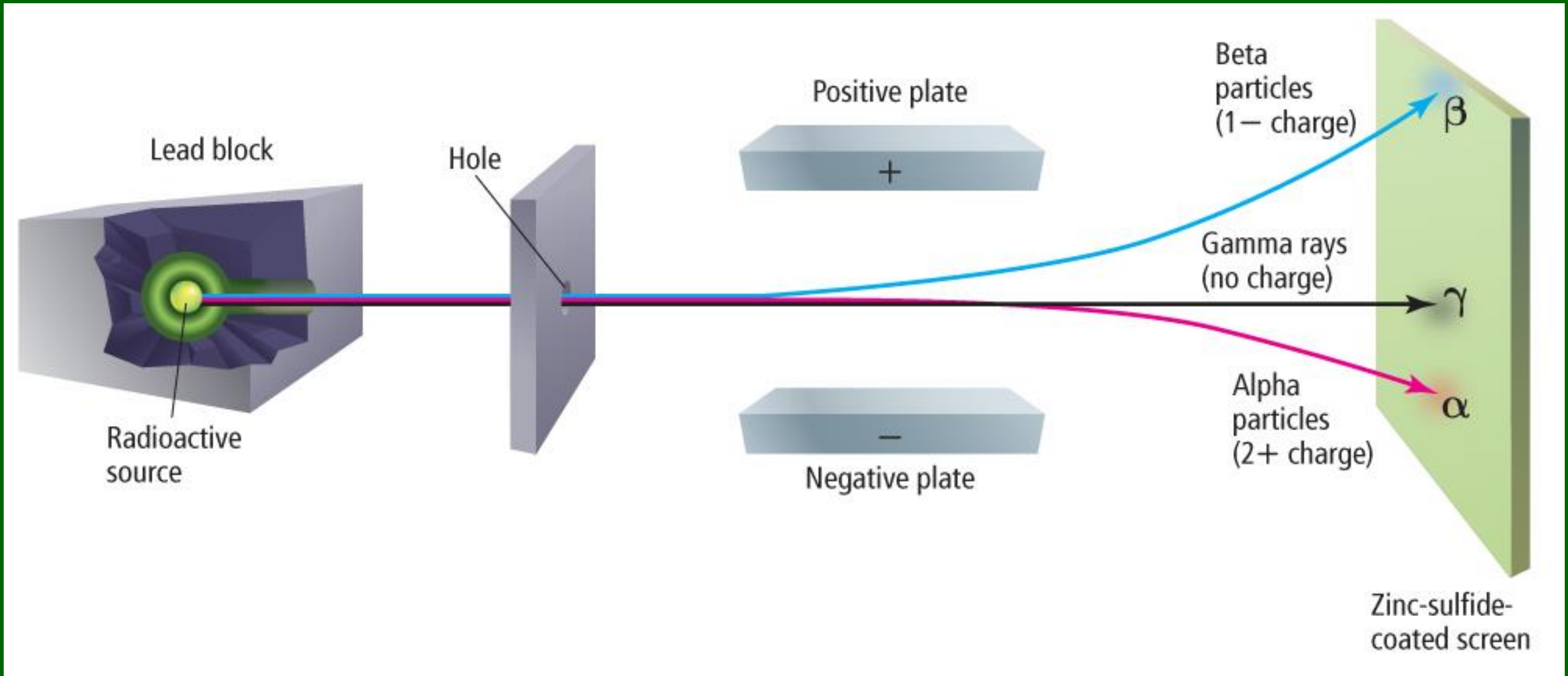
Beta Radiation

Beta radiation is radiation that has a negative charge and emits beta particles. Each **beta particle** is an electron with a 1– charge.

$^{14}_{6}C$	$\rightarrow$	$^{14}_{7}N$	$\rightarrow$	β
carbon-14		nitrogen-14		beta particle

During Beta decay, a neutron is converted to a proton and an electron. The electron is emitted and the proton stays in the nucleus.

Radiation Deflection



Gamma Radiation

Gamma rays are high-energy radiation with no mass and are neutral. They usually accompany alpha and beta radiation.

$^{238}_{92}U$	$\rightarrow$	$^{234}_{90}Th$	$+$	α	$+$	2γ
uranium-238		thorium-234		alpha particle		gamma rays

Gamma rays account for most of the energy lost during radioactive decay.

Nuclear Stability

Atoms that contain too many or two few neutrons are unstable and lose energy through radioactive decay to form a stable nucleus. Few exist in nature—most have already decayed to stable forms.

Table 5 Characteristics of Radiation

	Alpha	Beta	Gamma
Symbol	${}^{4}_{2}He$ or α	e^- or β	γ
Mass (amu)	4	$\frac{1}{1840}$	0
Mass (kg)	6.65×10^{-27}	9.11×10^{-31}	0
Charge	2+	1–	0