GENERAL CHEMISTRY

Principles and Modern Applications

TENTH EDITION

PETRUCCI HERRING MADURA BISSONNETTE

Chapter 2:

Atoms and the Atomic Theory

Atoms and the Atomic Theory

CONTENTS

2-1 Early Discoveries and the Atomic Theory

Chemistry has been practiced for a very long time.

- The blast furnace for extracting iron from iron ore, A.D. 1300
- Chemicals: sulfuric acid (oil of vitriol), nitric acid (aqua fortis), and sodium sulfate (Glauber s salt) were all well known.
- Before the end of the eighteenth century, the principal gases of the atmosphere nitrogen and oxygen had been isolated, and natural laws had been proposed describing the physical behavior of gases.

Yet chemistry cannot be said to have entered the modern age until the process of combustion was explained.

- In this section, we explore the direct link between the explanation of combustion and Dalton's atomic theory.

Law of conservation of mass:

Lavoisier 1774

- a sealed glass vessel containing a sample of tin and some air
- the mass before heating and after heating were the same
- the product of the reaction, tin calx (tin oxide), consisted of the original tin together with a portion of the air.
- oxygen from air is essential to combustion, and also led him to formulate the **law of conservation of mass:**

The total mass of substances present after a chemical reaction is the same as the total mass of substances before the reaction.

Mass is conserved during a chemical reaction

Silver Nitrate + potassium chromate $\rightarrow$ silver chromate

FIGURE 2-2

(a) (b)

- law of conservation of mass says that matter is neither created nor destroyed in a chemical reaction.

EXAMPLE 2-1 Applying the Law of Conservation of Mass

A 0.455 g sample of magnesium is allowed to burn in 2.315 g of oxygen gas. The sole product is magnesium oxide. After the reaction, no magnesium remains and the mass of unreacted oxygen is 2.015 g. What mass of magnesium oxide is produced?

Analyze

The total mass is unchanged. The total mass is the sum of the masses of the substances present initially. The mass of magnesium oxide is the total mass minus the mass of unreacted oxygen.

Solve

First, determine the total mass before the reaction.

$$\begin{aligned}\text{mass before reaction} &= 0.455\text{ g magnesium} + 2.315\text{ g oxygen}\\ &= 2.770\text{ g mass before reaction}\end{aligned}$$

The total mass after the reaction is the same as before the reaction.

$$\begin{aligned}2.770\text{ g mass after reaction} &= \text{? g magnesium oxide after reaction}\\ &\quad + 2.015\text{ g oxygen after reaction}\end{aligned}$$

Solve for the mass of magnesium oxide.

$$\begin{aligned}\text{? g magnesium oxide after reaction} &= 2.770\text{ g mass after reaction}\\ &\quad - 2.015\text{ g oxygen after reaction}\\ &= 0.755\text{ g magnesium oxide after reaction}\end{aligned}$$

Assess

Here is another approach. The mass of oxygen that reacted is 2.315 g − 2.015 g = 0.300 g. Thus, 0.300 g oxygen combined with 0.455 g magnesium to give 0.300 g + 0.455 g = 0.755 g magnesium oxide.

PRACTICE EXAMPLE A: A 0.382 g sample of magnesium is allowed to react with 2.652 g of nitrogen gas. The sole product is magnesium nitride. After the reaction, the mass of unreacted nitrogen is 2.505 g. What mass of magnesium nitride is produced?

PRACTICE EXAMPLE B: A 7.12 g sample of magnesium is heated with 1.80 g of bromine. All the bromine is used up, and 2.07 g of magnesium bromide is the only product. What mass of magnesium remains *unreacted*?

Law of constant composition / Law of definite proportions

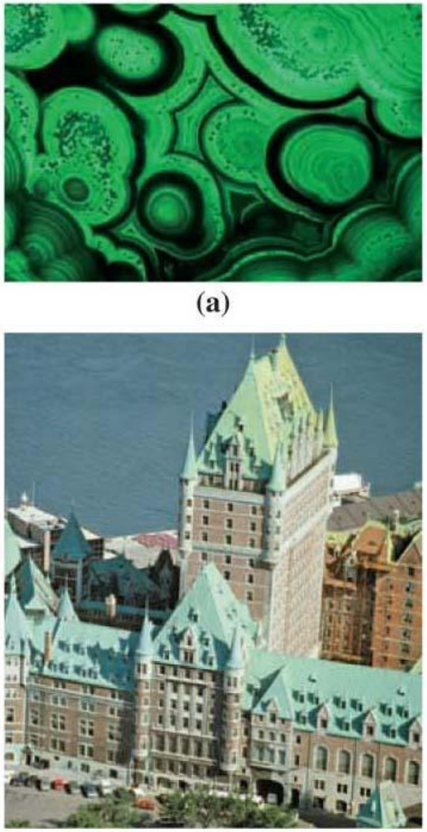
Proust 1799:

* One hundred pounds of copper, dissolved in sulfuric or nitric acids and precipitated by the carbonates of soda or potash, invariably gives 180 pounds of green carbonate. *

All samples of a compound have the same composition - the same proportions by mass of the constituent elements.

Sample A and Its Composition
10.000 g
1.119 g H, %H = 11.19
8.881 g O %O = 88.81

Sample B and Its composition
27.000 g
3.021 g H % H = 11.19
23.979 g O % O = 88.81



(a) (b)

▲ The mineral malachite **(a)** and the green patina on a copper roof **(b)** are both basic copper carbonate, just like the basic copper carbonate prepared by Proust in 1799.

EXAMPLE 2-2 Using the Law of Constant Composition

In Example 2-1 we found that when 0.455 g of magnesium reacted with 2.315 g of oxygen, 0.755 g of magnesium oxide was obtained. Determine the mass of magnesium contained in a 0.500 g sample of magnesium oxide.

Analyze

We know that 0.755 g of magnesium oxide contains 0.455 g of magnesium. According to the law of constant composition, the mass ratio 0.455 g magnesium/0.755 g magnesium oxide should exist in all samples of magnesium oxide.

Solve

Application of the law of constant composition gives

$$\frac{0.455\ \text{g magnesium}}{0.755\ \text{g magnesium oxide}} = \frac{?\ \text{g magnesium}}{0.500\ \text{g magnesium oxide}}$$

Solving the expression above, we obtain

$$?\ \text{g magnesium} = 0.500\ \cancel{\text{g magnesium oxide}} \times \frac{0.455\ \text{g magnesium}}{0.755\ \cancel{\text{g magnesium oxide}}}$$

$$= 0.301\ \text{g magnesium}$$

Assess

You can also work this problem by using mass percentages. If 0.755 g of magnesium oxide contains 0.455 g of magnesium, then magnesium oxide is $(0.455\ \text{g}/0.755\ \text{g}) \times 100\% = 60.3\%$ magnesium by mass and $(100\% - 60.3\%) = 39.7\%$ oxygen by mass. Thus, a 0.500 g sample of magnesium oxide must contain $0.500\ \text{g} \times 60.3\% = 0.301$ g of magnesium and $0.500\ \text{g} \times 39.7\% = 0.199$ g of oxygen.

PRACTICE EXAMPLE A: What masses of magnesium and oxygen must be combined to make exactly 2.000 g of magnesium oxide?

PRACTICE EXAMPLE B: What substances are present, and what are their masses, after the reaction of 10.00 g of magnesium and 10.00 g of oxygen?

Dalton's Atomic Theory:

1. Each element is composed of minute, indivisible particles called **atoms**. Atoms are **neither created nor destroyed** in chemical reactions.

2. All atoms of a given element are **alike** in mass (weight) and other properties but differ from all other elements

3. In each of their **compounds,** different elements combine in simple numerical ratios. One atom of A to one of B (AB), or one atom of A to two of B (AB_2).

▲ John Dalton (1766–1844), developer of the atomic theory. Dalton has not been considered a particularly good experimenter, perhaps because of his color blindness (a condition sometimes called daltonism). However, he did skillfully use the data of others in formulating his atomic theory. (The Granger Collection)

Dalton’s theory explains the law of conservation of mass.

- If atoms of an element are indestructible (assumption 1), then the *same atoms* must be present after a chemical reaction as before. The total mass remains unchanged.

Dalton’s theory also explains the law of constant composition.

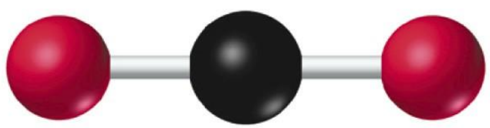
- If all atoms of an element are alike in mass (assumption 2) and if atoms unite in *fixed* numerical ratios (assumption 3), the percent composition of a compound must have a unique value, regardless of the origin of the sample analyzed.

Like all good theories, Dalton s atomic theory led to a prediction of the **law of multiple proportions.**

- The law of multiple proportions.

If two elements form more than a single compound, the masses of one element combined with a fixed mass of the second are in the ratio of small whole numbers.

- In forming carbon monoxide, 1.0 g of carbon combines with 1.33 g of oxygen.
- In forming carbon dioxide, 1.0 g of carbon combines with 2.66 g of oxygen.



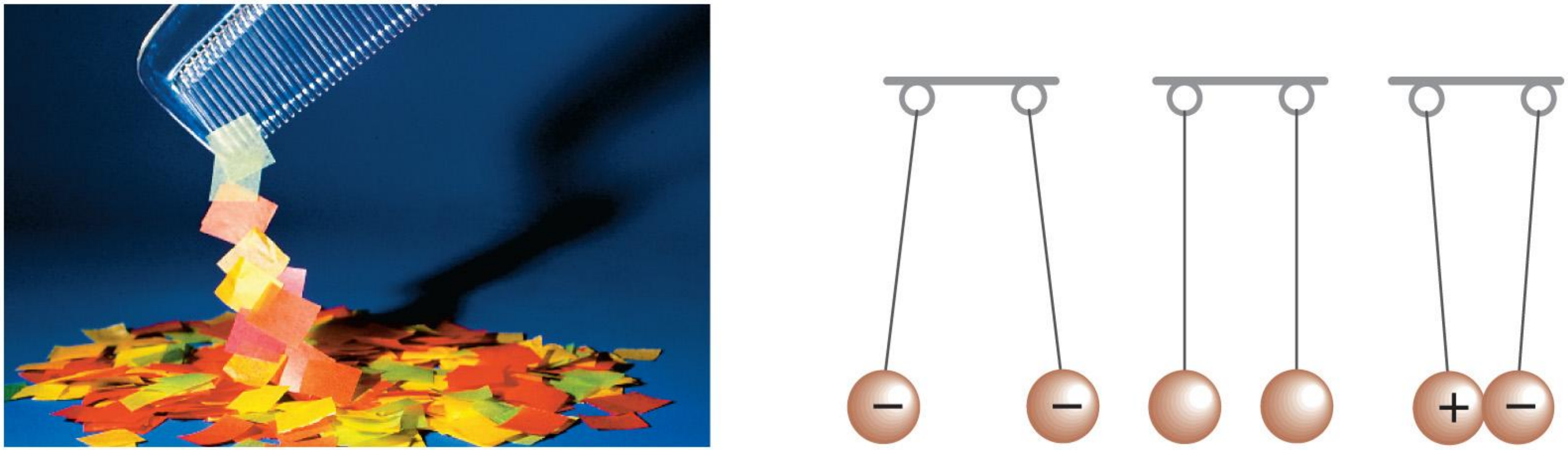
CO and CO_2 1:1 and 1:2

▲ Figure 2-3

Consequences of Dalton's theory

2-2 Electrons and Other Discoveries in Atomic Physics

Electricity and magnetism were used in the experiments that led to the current theory of atomic structure.



(a) (b)

(a) Electrostatically charged comb. If you comb your hair on a dry day, a static charge develops on the comb and causes bits of paper to be attracted to the comb. (b) Both objects on the left carry a negative electric charge. Objects with like charge repel each other. The objects in the center lack any electric charge and exert no forces on each other. The objects on the right carry opposite charges one positive and one negative and attract each other.

Effect of magnetic field on charged particles:

* When charged particles travel through a magnetic field so that their path is perpendicular to the field, they are deflected by the field.
* Negatively charged particles are deflected in one direction, and positively charged particles in the opposite direction.
* Several phenomena described in this section depend on this behavior.

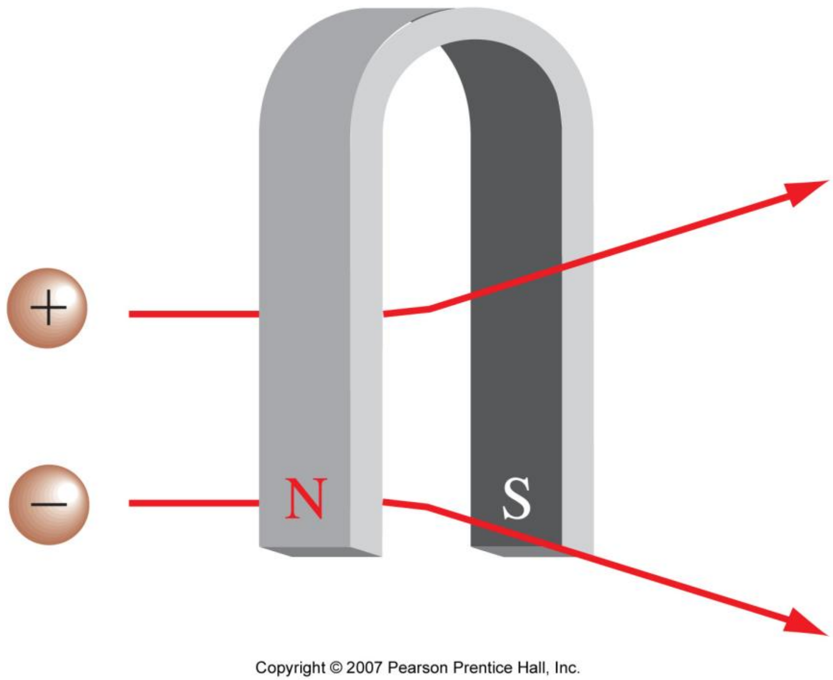


FIGURE 2-5

Effect of a magnetic field on charged particles

The Discovery of Electrons:

✓CRT, the abbreviation for cathode-ray tube, was once a familiar acronym. Before liquid crystal display (LCD) was available, the CRT was the heart of computer monitors and TV sets.

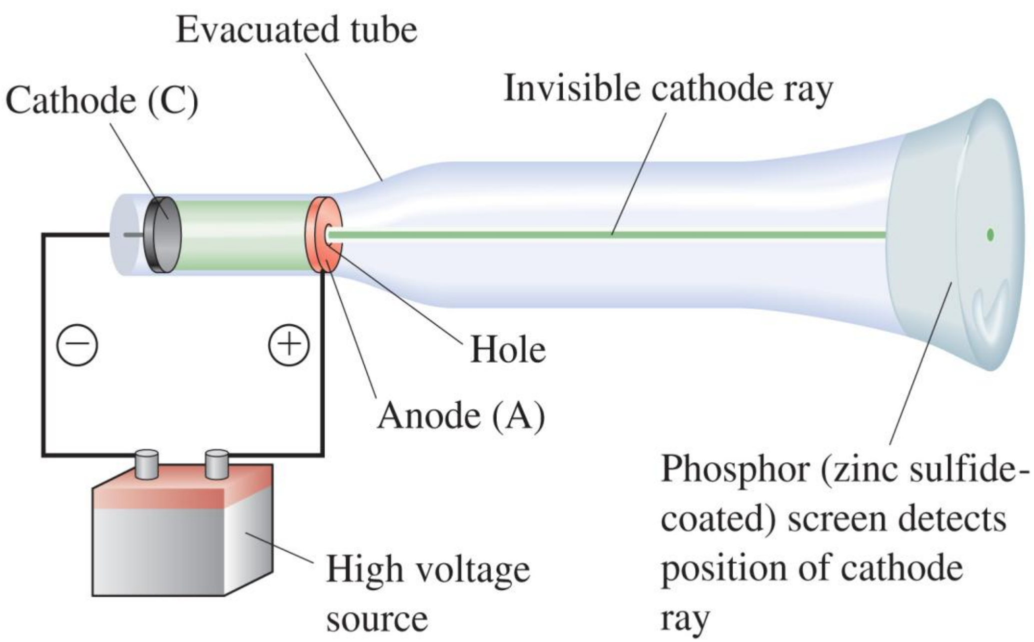


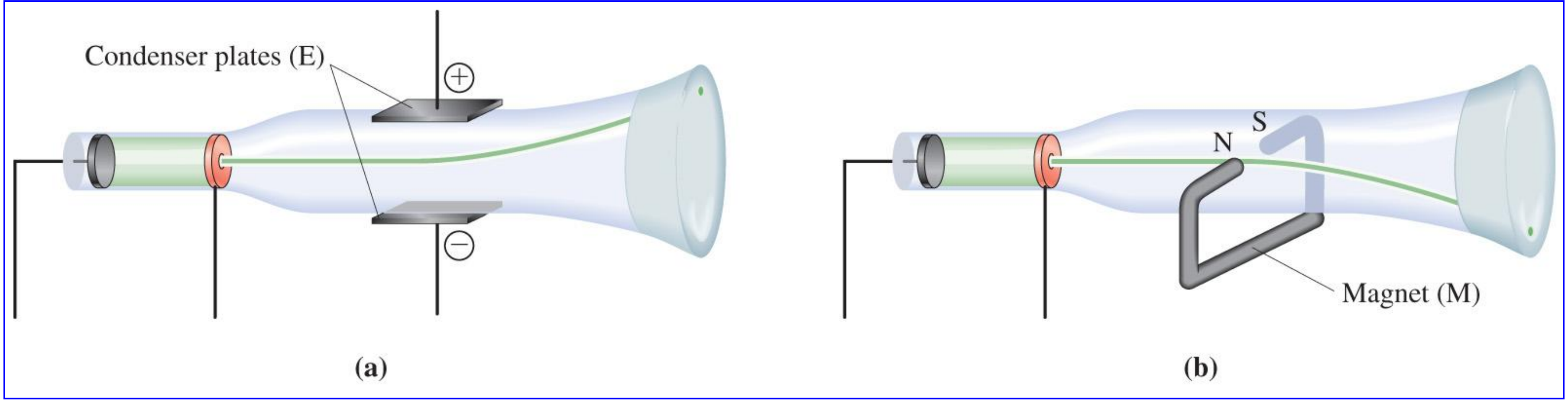
FIGURE 2-6 Cathode ray tube

✓ Michael Faraday (1791- 1867) discovered Cathode Rays, a type of radiation emitted by the negative terminal, or *cathode,* when he passed electricity through glass tubes from which most of the air had been evacuated,

✓ Scientists found that cathode rays travel in straight lines and have properties that are independent of the cathode material (whether it is iron, platinum, and so on).

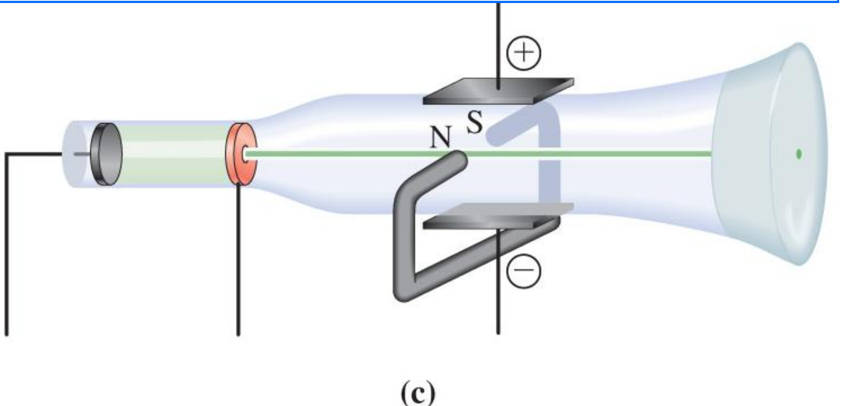
✓ The cathode rays produced in the CRT are invisible, and they can be detected only by the light emitted by materials that they strike. These materials, called *phosphors, are painted on the end of the CRT so* that the path of the cathode rays can be revealed.

Another significant observation about cathode rays is that they are deflected by electric and magnetic fields in the manner expected for negatively charged particles



J. J. Thomson in 1897 established the ratio of mass to electric charge m/e *for cathode rays.* The cathode-ray beam strikes the end screen undeflected if the forces exerted on it by the electric and magnetic fields are counterbalanced.

- Also, concluded that cathode rays are negatively charged fundamental particles of matter found in all atoms.
- Cathode rays subsequently became known as electrons, a term first proposed by George Stoney in 1874.



Precise measurements yield Electron m/e $= -5.6857 \times 10^{-9}$ g coulomb^{-1}

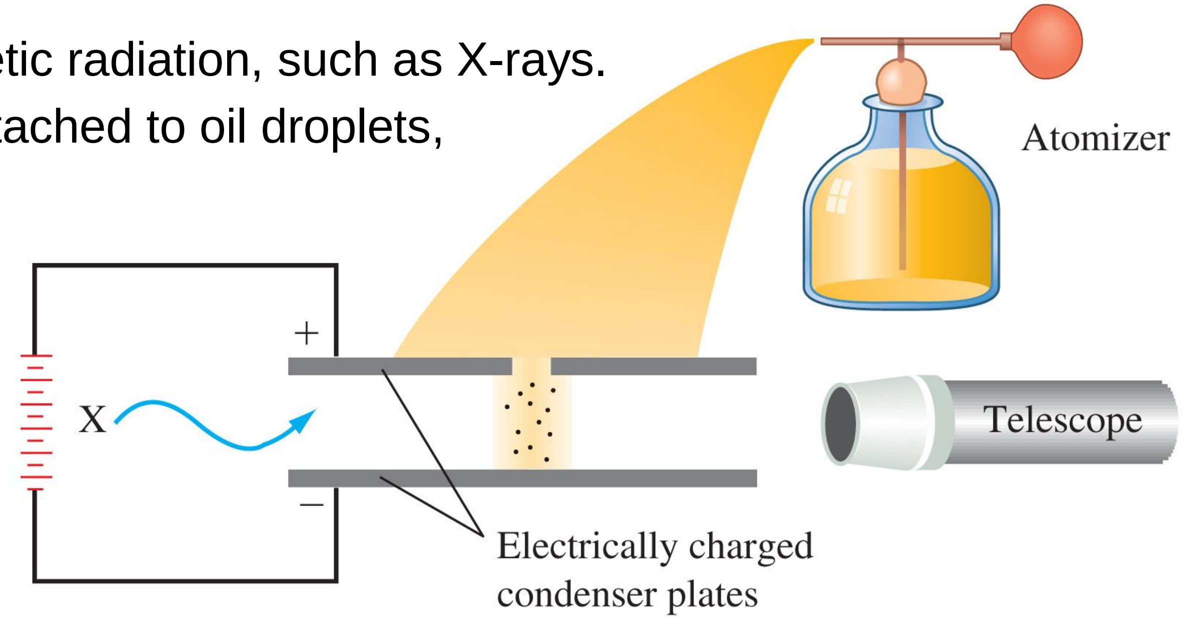
Determination of electronic charge, e

☆ Robert Millikan (1868-1953) showed ionized oil drops can be balanced against the pull of gravity by an electric field.

Millikan's oil-drop experiment: Figure 2-8

➢ Ions are produced by energetic radiation, such as X-rays. Some of these ions become attached to oil droplets, giving them a net charge.

➢ The fall of a droplet in the electric field between the condenser plates is speeded up or slowed down, depending on the magnitude and sign of the charge on the droplet.



➢ By analyzing data from a large number of droplets, Millikan concluded that the magnitude of the charge, q, on a droplet is an integral multiple of the electric charge,e.

➢ The currently accepted value of $e = -1.6022 * 10^{-19}$ C .

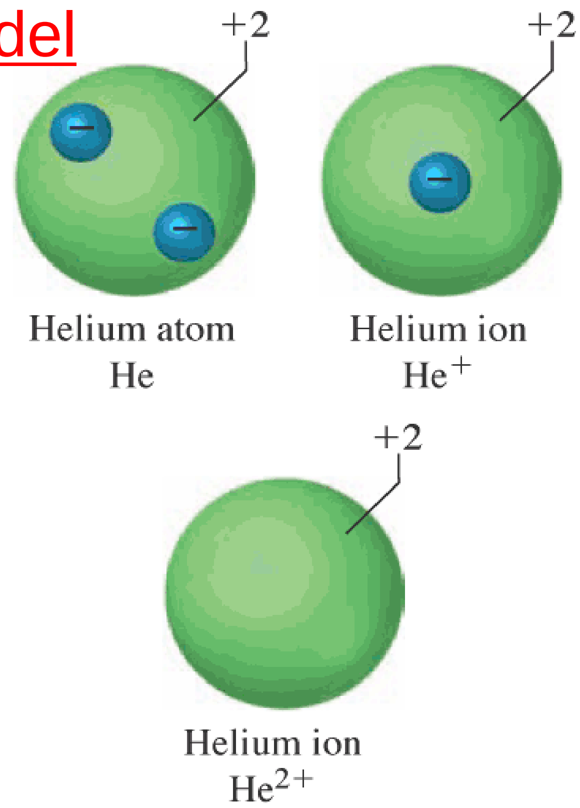
➢ By combining this value with an accurate value of the mass-to-charge ratio for an electron, $m/e = -5.6857 \times 10^{-9}$ g . C^{-1} , we find that the mass of an electron as

mass of an electron = $9.1094 * 10^{-28}$ g.

Once the electron was seen to be a fundamental particle of matter found in all atoms, atomic physicists began to speculate on how these particles were incorporated into atoms?

J. J. Thoson's the plum-pudding atomic model

- the positive charge necessary to counterbalance the negative charges of electrons in a neutral atom was in the form of a nebulous cloud.
- Electrons, he suggested, floated in a diffuse cloud of positive charge (rather like a lump of gelatin with electron fruit embedded in it).
- This model became known as the plum-pudding model because of its similarity to a popular English dessert.



▲ FIGURE 2-9
The plum-pudding atomic model

X-Rays and Radioactivity

✓ Cathode-ray research had many important spin-offs. In particular, two natural phenomena of immense theoretical and practical significance (*x-rays and radioactivity*) were discovered in the course of other investigations.

✓ In 1895, Wilhelm Roentgen (1845 - 1923) noticed that when cathode-ray tubes were operating, certain materials *outside the tubes glowed or fluoresced.*

- ✓ He showed that this fluorescence was caused by radiation emitted by the cathode-ray tubes.
- ✓ Because of the unknown nature of this radiation, coined the term *X-ray. We now recognize the X-ray as a form of high-energy* electromagnetic radiation

Antoine Henri Becquerel (1852 - 1908) associated X-rays with fluorescence and wondered if naturally fluorescent materials produce X-rays.

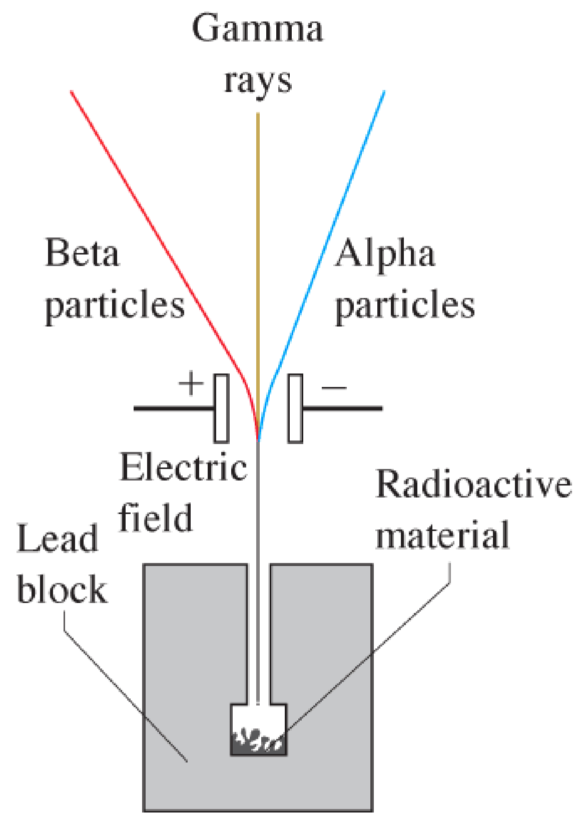
To test this idea, he wrapped a photographic plate with black paper, placed a coin on the paper, covered the coin with a uranium-containing fluorescent material, and exposed the entire assembly to sunlight. When he developed the film, a clear image of the coin could be seen. The fluorescent material had emitted radiation (presumably X-rays) that penetrated the paper and exposed the film. On one occasion, because the sky was overcast, Becquerel placed the experimental assembly inside a desk drawer for a few days while waiting for the weather to clear. On resuming the experiment, Becquerel decided to replace the original photographic film, expecting that it may have become slightly exposed. He developed the original film and found that instead of the expected feeble image, there was a very sharp one. The film had become strongly exposed because the uranium-containing material had emitted radiation continuously, even when it was not fluorescing. Becquerel had discovered ***radioactivity*** by *chance/serendipity*

✓ Ernest Rutherford (1871 - 1937) identified two types of radiation from radioactive materials, alpha and beta.

✓Alpha particles, α, carry two fundamental units of positive charge and have essentially the same mass as helium atoms. In fact, alpha particles are identical to ions.

✓ Beta particles, β, are negatively charged particles produced by changes occurring within the nuclei of radioactive atoms and have the same properties as electrons.

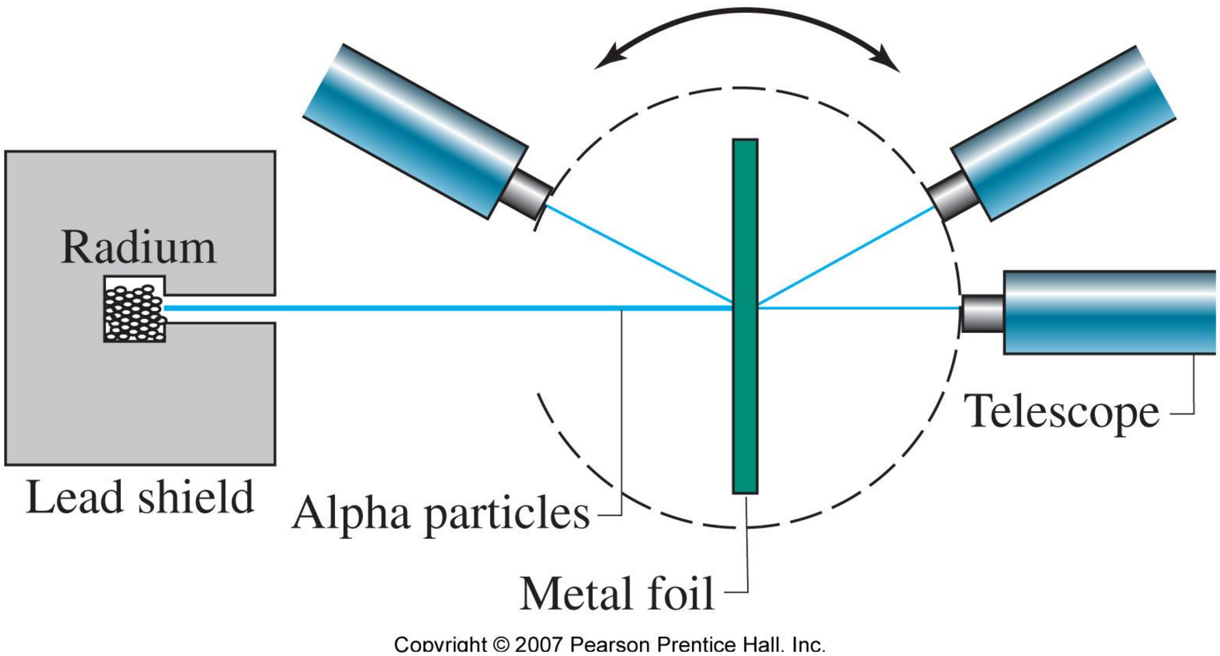
✓ A third form of radiation, which is not affected by electric or magnetic fields, was discovered in 1900 by Paul Villard. This radiation, called gamma rays, γ, is not made up of particles; it is electromagnetic radiation of extremely high penetrating power



▲ FIGURE 2-10
Three types of radiation from radioactive materials

2-3 The Nuclear Atom

- In 1909, Rutherford, with his assistant Hans Geiger, began a line of research using particles as probes to study the inner structure of atoms.
- Rutherford expected that most particles would pass through thin sections of matter largely undeflected, but that some particles would be *slightly* scattered or deflected as they encountered electrons. By studying these scattering patterns, he hoped to deduce something about the distribution of electrons in atoms.



- The telescope travels in a circular track around an evacuated chamber containing the metal foil. Alpha particles were detected by the flashes of light they produced when they struck a zinc sulfide screen mounted on the end of a telescope.

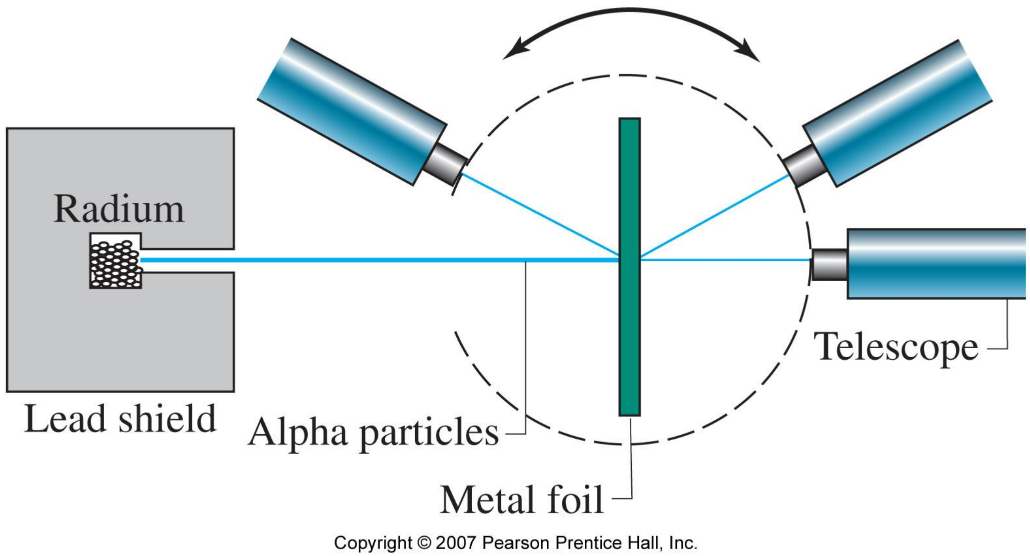


Figure 2-11
The α-particle scattering experiment

When they bombarded very thin foils of gold with alpha particles, they observed the following;

- The majority of particles penetrated the foil undeflected.
- Some particles experienced slight deflections.
- A few (about 1 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

Rutherford's nuclear atom model:

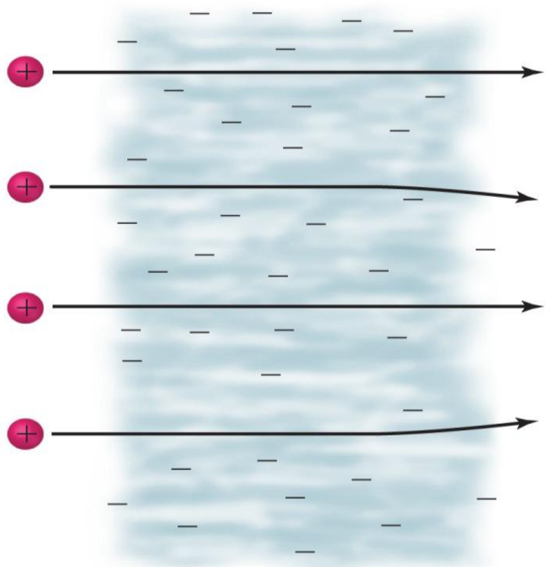
➢ The large-angle scattering greatly puzzled Rutherford. As he commented some years later, this observation was about as credible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you.

➢ By 1911, Rutherford had an explanation. He based his explanation on 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 remainder of 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.

▲

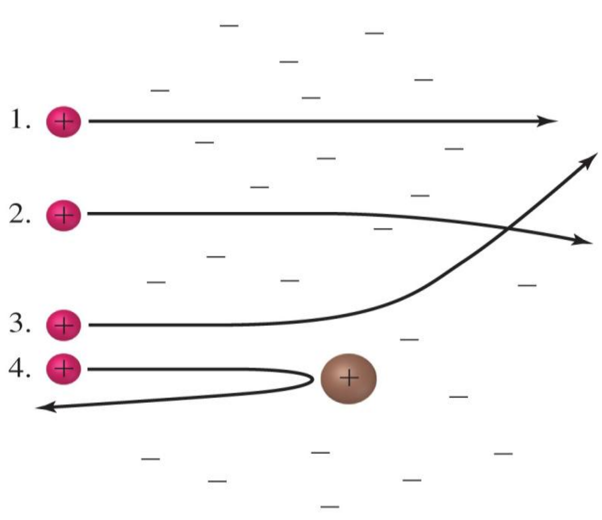
Explaining the results of alpha particle scattering experiments



(a) Rutherfords expectation was that small, positively charged particles should pass through the nebulous, positively charged cloud of the Thomson plum-pudding model largely undeflected. Some would be slightly deflected by passing near electrons (present to neutralize the positive charge of the cloud).

(b) Rutherford s explanation was based on a nuclear atom. With an atomic model having a small, dense, positively charged nucleus and extranuclear electrons, we would expect the four different types of paths actually observed:

1.undeflected straight-line paths exhibited by most of the particles
2.slight deflections of particles passing close to electrons
3.severe deflections of particles passing close to a nucleus
4.reflections from the foil of particles approaching a nucleus head-on



(b)

Discovery of Protons and Neutrons

✓ Rutherford, protons 1919
✓ James Chadwick , neutrons 1932
✓ Thus, it has been only for about the past 100 years that we have had the atomic model suggested by Figure 2-13.

✓ In this drawing, electrons are shown much closer to the nucleus than is the case. The actual situation is more like this: If the entire atom were represented by a room, 5m x 5m x 5m, the nucleus would occupy only about as much space as the period at the end of this sentence.
✓ The heaviest atom has a mass of only 4.8 x 10^{-22} g and a diameter of only 5 x 10^{-10} m.
✓ Biggest atom is 240 amu and is 50 **Å** across.
✓ Typical C-C bond length 154 pm (1.54 **Å**)
Molecular models are 1 **Å** /inch or about 0.4 **Å** /cm

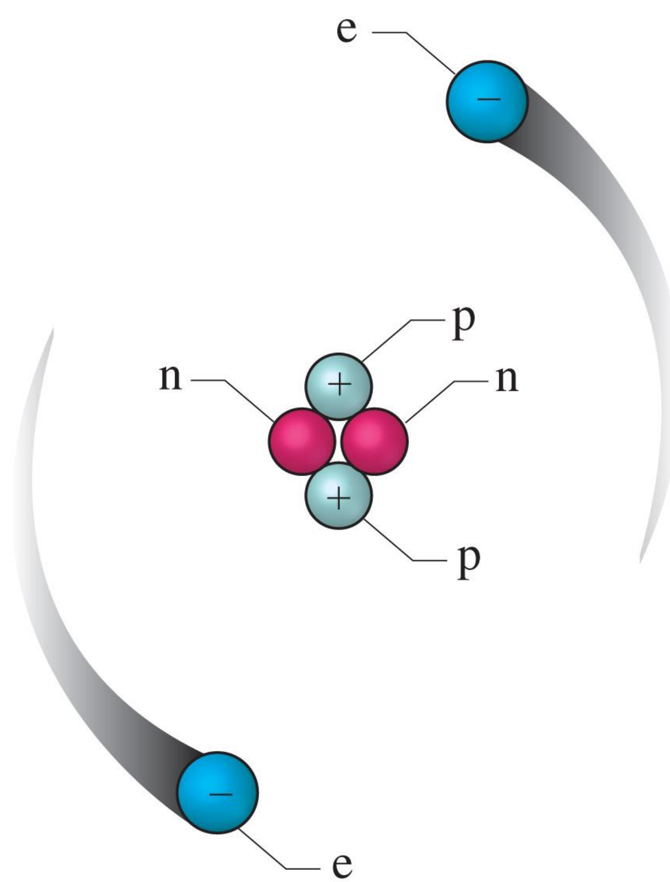


Figure 2.13
The nuclear atom-
İllustrated by the Helium atom

TABLE 2.1 Properties of Three Fundamental Particles

	Electric Charge		Mass	
	SI (C)	Atomic	SI (g)	Atomic (u)[a]
Proton	$+1.6022 \times 10^{-19}$	$+1$	1.6726×10^{-24}	1.0073
Neutron	0	0	1.6749×10^{-24}	1.0087
Electron	-1.6022×10^{-19}	-1	9.1094×10^{-28}	0.00054858

[a]u is the SI symbol for atomic mass unit (abbreviated as amu).

- The masses of the proton and neutron are different in the fourth significant figure.
- The charges of the proton and electron, however, are believed to be exactly equal in magnitude (but opposite in sign).
- The charges and masses are known much more precisely than suggested here. More precise values are given on the inside back cover.

2-4 Chemical Elements

All atoms of a particular element have the same atomic number, Z, and, conversely, all atoms with the same number of protons are atoms of the same element. The elements shown on the inside front cover have atomic numbers from Z:1 to Z:112.

∗To represent a particular atom we use symbolism:

- Chemical symbols are one- or two-letter abbreviations of the name (usually the English name).
- The first (but never the second) letter of the symbol is capitalized; for example: carbon, C; oxygen, O; neon, Ne; and silicon, Si.
- Some elements known since ancient times have symbols based on their Latin names, such as Fe for iron (ferrum) and Pb for lead (plumbum).
- The element sodium has the symbol Na, based on the Latin Natrium for sodium carbonate. Potassium has the symbol K, based on the Latin Kalium for potassium carbonate. The symbol for tungsten, W, is based on the German wolfram.
- Elements beyond uranium do not occur naturally and must be synthesized in particle accelerators

$$\text{number p} + \text{number n} \longrightarrow {}^{A}_{Z}\mathrm{E}^{\pm ?} \longleftarrow \text{number p} - \text{number e}$$

$$\text{number p} \longrightarrow$$

A= mass number   Z = atomic number

This symbolism indicates that the atom is element E and that it has atomic number Z and mass number A .
For example, an atom of aluminum represented as ${}_{13}Al^{27}$
has 13 protons and 14 neutrons in its nucleus and 13 electrons outside the nucleus. (Recall that an atom has the same number of electrons as protons.

Contrary to what Dalton thought, we now know that atoms of an element do not necessarily all have the same mass.

Isotopes:

✓ Atoms that have the same atomic number (Z) but different mass numbers (A) are called isotopes.

✓ All neon atoms have 10 protons in their nuclei, and most have 10 neutrons as well. A very few neon atoms, however, have 11 neutrons and some have 12.

✓ Of all Ne atoms on Earth, 90.51% are $^{20}Ne_{10}$. These percentages 90.51%, 0.27%, 9.22% are the percent natural abundances of the three neon isotopes.

✓ Percent natural abundances are always based on numbers, not masses. Thus, 9051 of every 10,000 neon atoms are neon-20 atoms.

✓ Some elements, as they exist in nature, consist of just a single type of atom and therefore do not have naturally occurring isotopes. Aluminum, for example, consists only of aluminum-27 atoms.

Ions:

- When atoms lose or gain electrons, for example, in the course of a chemical reaction, the species formed are called **ions and carry net charges. Because an** electron is negatively charged, adding electrons to an electrically neutral atom produces a negatively charged ion. Removing electrons results in a positively charged ion.
- The number of protons does not change when an atom becomes an ion.

EXAMPLE 2-3 Relating the Numbers of Protons, Neutrons, and Electrons in Atoms and Ions

Through an appropriate symbol, indicate the number of protons, neutrons, and electrons in **(a)** an atom of barium-135 and **(b)** the double negatively charged ion of selenium-80.

Analyze

Given the name of an element, we can find the symbol and the atomic number, Z, for that element from a list of elements or a periodic table. To determine the number of protons, neutrons, and electrons, we make use of the following relationships:

$$Z = \text{number p} \qquad A = \text{number p} + \text{number n} \qquad \text{charge} = \text{number p} - \text{number e}$$

The relationships above are summarized in expression (2.2).

Solve

(a) We are given the name (barium) and the mass number of the atom (135). From a list of the elements or a periodic table we obtain the symbol (Ba) and the atomic number ($Z = 56$), leading to the symbolic representation

$$^{135}_{56}\text{Ba}$$

From this symbol one can deduce that the neutral atom has 56 protons; a neutron number of $A - Z = 135 - 56 = 79$ neutrons; and a number of electrons equal to Z, that is, 56 electrons.

(b) We are given the name (selenium) and the mass number of the ion (80). From a list of the elements or a periodic table we obtain the symbol (Se) and the atomic number (34). Together with the fact that the ion carries a charge of 2−, we have the data required to write the symbol

$$^{80}_{34}\text{Se}^{2-}$$

From this symbol, we can deduce that the ion has 34 protons; a neutron number of $A - Z = 80 - 34 = 46$ neutrons; and 36 electrons, leading to a net charge of $+34 - 36 = -2$.

Assess

When writing the symbol for a particular atom or ion, we often omit the atomic number. For example, for $^{135}_{56}\text{Ba}$ and $^{80}_{34}\text{Se}^{2-}$, we often use the simpler representations ^{135}Ba and $^{80}\text{Se}^{2-}$.

PRACTICE EXAMPLE A: Use the notation $^{A}_{Z}\text{E}$ to represent the isotope of silver having a neutron number of 62.

PRACTICE EXAMPLE B: Use the notation $^{A}_{Z}E$ to represent a tin ion having the same number of electrons as an atom of the isotope cadmium-112. Explain why there can be more than one answer.

Isotopic Masses:

- We cannot determine the mass of an individual atom just by adding up the masses of its fundamental particles. When protons and neutrons combine to form a nucleus, a very small portion of their original mass is converted to energy and released. However, we cannot predict exactly how much this so called nuclear binding energy will be. Determining the masses of individual atoms, then, is something that must be done by experiment, in the following way.
- By international agreement, one type of atom has been chosen and assigned a specific mass. This standard is an atom of the isotope carbon-12, which is assigned a mass of exactly 12 atomic mass units, that is, 12 u. Next,the masses of other atoms relative to carbon-12 are determined with a **mass spectrometer.**

In this device, a beam of gaseous ions passing through electric and magnetic fields separates into components of differing masses. The separated ions are focused on a measuring instrument, which records their presence and amounts. Figure 2-14 illustrates mass spectrometry and a typical mass spectrum.

Although mass numbers are whole numbers, the actual masses of individual atoms (in atomic mass units, u) are never whole numbers, except for carbon-12. However, they are very close in value to the corresponding mass numbers, as we can see for the isotope oxygen-16. From mass spectral data the ratio of the mass of ^{16}O to $^{12}C$is found to be 1.33291. Thus, the mass of the oxygen-16 atom is

$$1.33291 * 12\ u = 15.9949\ u$$

which is very nearly equal to the mass number of 16.

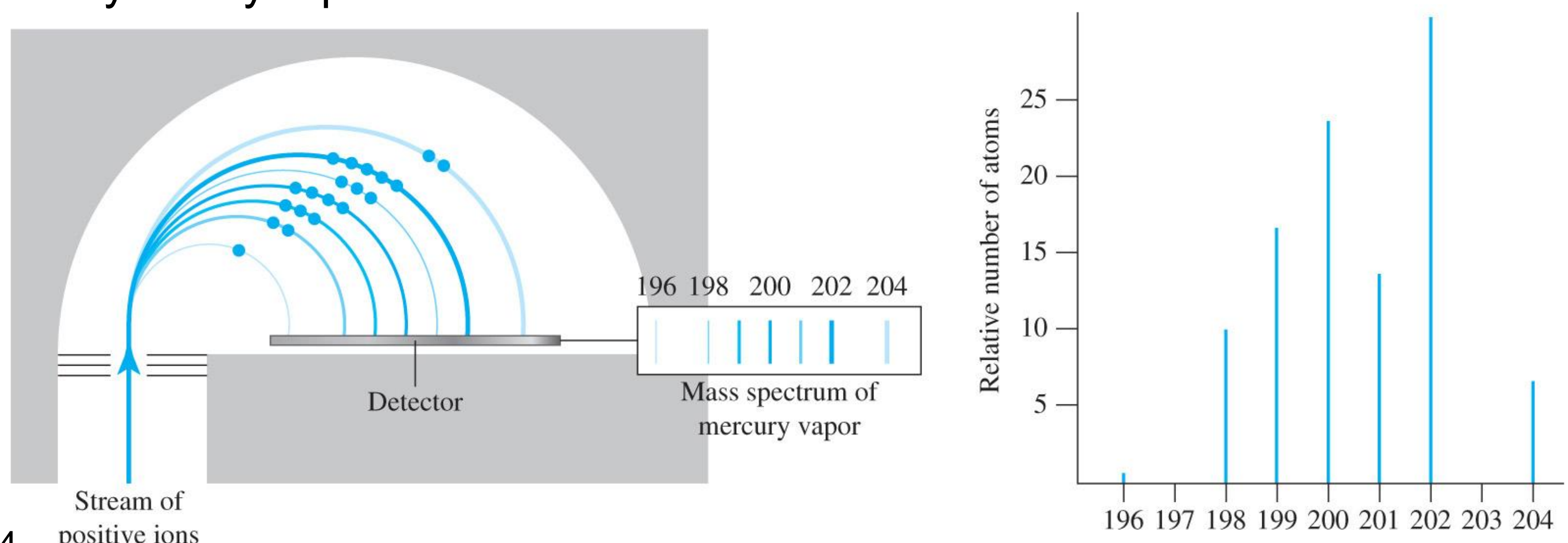


Figure 2-14

A mass spectrometer and a mass spectrum

2-5 Atomic Mass

For Carbon, the existence of two isotopes causes the observed atomic mass to be greater than 12.

Weighted Average Atomic Mass of an Element

= fractional abundance of isotope 1 x atomic mass of isotope 1 + fractional abundance of isotope 2 x atomic mass of isotope 2 + ……

$$A_{ave} = \xi_1 \times A_1 + \xi_2 \times A_2 + \ldots\ldots \xi_n \times A_n$$

where $\xi_1 + \xi_2 + \square.. + \xi_n = 1.0$

at. mass of naturally occurring carbon

= 0.9893 * 12 u + (1 - 0.9893) * 13.0033548378 u

= 13.0033548378 u - 0.9893

= 12.0107 u

The Periodic Table:

- We can distinguish one element from all others by its particular set of observable physical properties.
- Periodic Table is the classification system, in which we can;

  * read atomic masses,

  * read the ions formed by main group elements,

  * read the electron configuration.

  * learn trends in physical and chemical properties.

✓ In the periodic table, elements are listed according to increasing atomic number starting at the upper left and arranged in a series of horizontal rows,called periods.

✓ This arrangement places similar elements in *vertical* **groups, or families** The elements can also be subdivided into broad categories.

✓ One categorization is that of **metals, nonmetals, metalloids, and noble gases.**

✓ **Another is that of maingroup elements and transition elements (transition metals). Included among the transition elements are the** two subcategories **lanthanides and actinides.**

The Periodic table

Alkali Metals

Noble Gases

Alkaline Earths

Halogens

Main Group

Transition Metals

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 131.29
55 Cs 132.905	56 Ba 137.327	57 *La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	89 †Ac 227.028	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (272)		114 (287)		116 (289)		118 (293)

*Lanthanide series	58 Ce 140.115	59 Pr 140.908	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
†Actinide series	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np 237.048	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Main Group

Lanthanides and Actinides

EXAMPLE 2-7 Describing Relationships Based on the Periodic Table

Refer to the periodic table on the inside front cover, and indicate

(a) the element that is in group 14 and the fourth period;
(b) two elements with properties similar to those of molybdenum (Mo);
(c) the ion most likely formed from a strontium atom.

Analyze

For **(a)**, the key concept is that the rows (periods) are numbered 1 through 7, starting from the top of the periodic table, and the groups are numbered 1 through 18, starting from the left side. For **(b)**, the key concept is that elements in the same group have similar properties. For **(c)**, the key concept is that main-group metal atoms in groups 1 and 2 form positive ions with charges of +1 and +2, respectively.

Solve

(a) The elements in the fourth period range from K ($Z = 19$) to Kr ($Z = 36$). Those in group 14 are C, Si, Ge, Sn, and Pb. The only element that is common to both of these groupings is Ge ($Z = 32$).
(b) Molybdenum is in group 6. Two other members of this group that should resemble it are chromium (Cr) and tungsten (W).
(c) Strontium (Sr) is in group 2. It should form the ion Sr^{2+}.

Assess

In Chapter 8, we will examine in greater detail reasons for the arrangement of the periodic table.

PRACTICE EXAMPLE A: Write a symbol for the ion most likely formed by an atom of each of the following: Li, S, Ra, F, I, and Al.

PRACTICE EXAMPLE B: Classify each of the following elements as a main-group or transition element. Also, specify whether they are metals, metalloids, or nonmetals: Na, Re, S, I, Kr, Mg, U, Si, B, Al, As, H.

Metals; (except for mercury, a liquid,)

- solids at room temperature.
- malleable (capable of being flattened into thin sheets),
- ductile (capable of being drawn into fine wires),
- good conductors of heat and electricity,
- have a lustrous or shiny appearance.
- maingroup metal atoms in groups 1 and 2 form ions, they lose the same number of electrons as the IUPAC group number. Na+ and Ca++

Nonmetals ;

- are poor conductors of heat and electricity.
- several of the nonmetals, such as nitrogen, oxygen, and chlorine, are gases at room temperature.
- Some, such as silicon and sulfur, are brittle solids. One bromine is a liquid.
- nonmetal atoms form ions, they gain electrons. O^{2-} , Cl^{-}

2-7 The Concept of the Mole and the Avogadro Constant

- Physically counting atoms is impossible.
- We must be able to relate measured mass to numbers of atoms.

- The SI quantity that describes an amount of substance by relating it to a number of particles of that substance is called the *mole (abbreviated mol).*

•A **mole is the amount of a substance that contains the same number of elementary** entities as there are atoms in exactly 12 g of pure carbon-12.
• The number of elementary entities (atoms, molecules, and so on) in a mole is the **Avogadro constant, N_A**

$$N_A = 6.02214179 \times 10^{23}\ \text{mol}^{-1}$$

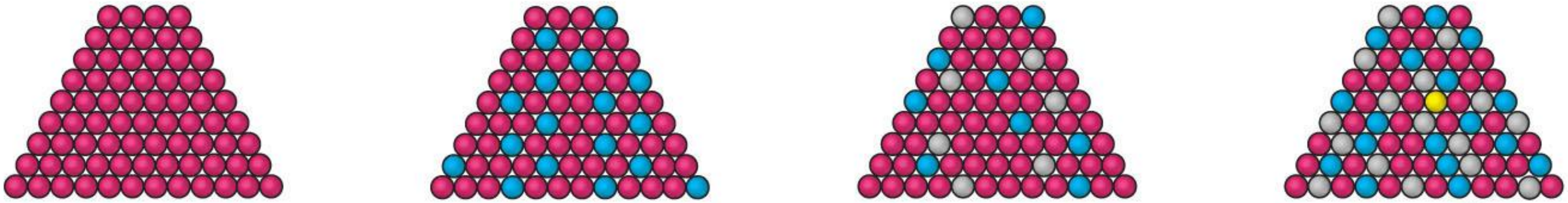
If a substance contains atoms of only a single isotope, then
1 mol ^{12}C = 6.02214 * 10^{23} ^{12}C atoms = 12.0000 g
1 mol ^{16}O = 6.02214 * 10^{23} ^{16}O atoms = 15.9949 g

- - However, most elements are composed of mixtures of two or more isotopes so that the atoms in a sample of the element are not all of the same mass but are present in their naturally occurring proportions: Thus, in one mole of carbon, most of the atoms are carbon-12, but some are carbon-13. In one mole of oxygen, most of the atoms are oxygen-16, but some are oxygen-17 and some are oxygen-18. As a result,

1 mol of C = 6.02214 * 10^{23} C atoms = 12.0107 g

1 mol of O = 6.02214 * 10^{23} O atoms = 15.9994 g, and so on.

Distribution of isotopes in four elements:



(a) 6.02214×10^{23} F atoms $= 18.9984$ g

(b) 6.02214×10^{23} Cl atoms $= 35.453$ g

(c) 6.02214×10^{23} Mg atoms $= 24.3050$ g

(d) 6.02214×10^{23} Pb atoms $= 207.2$ g

▲ FIGURE 2-16
Distribution of isotopes in four elements

(a) There is only one type of fluorine atom, ^{19}F (shown in red).
(b) In chlorine, 75.77% of the atoms are ^{35}Cl (red) and the remainder are ^{37}Cl (blue).
(c) Magnesium has one principal isotope, ^{24}Mg (red), and two minor ones, ^{25}Mg(gray) and ^{26}Mg (blue).
(d) Lead has four naturally occurring isotopes: 1.4% ^{204}Pb(yellow), 24.1% ^{206}Pb (blue), 22.1% ^{207}Pb (gray) and 52.4% ^{208}Pb (red).

Molar Mass:

The molar mass, M, is the mass of one mole of a substance.

$$M\,(\text{g/mol } ^{12}\text{C}) = A\,(\text{g/atom } ^{12}\text{C}) \times N_A\,(\text{atoms } ^{12}\text{C /mol } ^{12}\text{C})$$

One mole of an element :
The watch glasses contain one mole of copper atoms (left) and one mole of sulfur atoms (right). The beaker contains one mole of mercury atoms as liquid mercury, and the balloon, of which only a small portion is visible here, contains one mole of helium atoms in the gaseous state.

2-8 Using the Mole Concept in Calculations

1 mol S =

$6.022 * 10^{23}$ S atoms

1 mol S=
32.065 g S

EXAMPLE 2-8 **Relating Number of Atoms, Amount in Moles, and Mass in Grams**

In the sample of sulfur weighing 4.07 g pictured in Figure 2-18, **(a)** how many moles of sulfur are present, and **(b)** what is the total number of sulfur atoms in the sample?

Analyze

For **(a)**, the conversion pathway is g S $\rightarrow$ mol S. To carry out this conversion, we multiply 4.07 g S by the conversion factor (1 mol S/32.07 g S). The conversion factor is the molar mass inverted. For **(b)**, the conversion pathway is mol S $\rightarrow$ atoms S. To carry out this conversion, we multiply the quantity in moles from part **(a)** by the conversion factor (6.022×10^{23} atoms S/1 mol S).

Solve

(a) For the conversion g S $\rightarrow$ mol S, using $(1/M)$ as a conversion factor achieves the proper cancellation of units. The result of this calculation should be stored without rounding it off because it is required in part (b).

$$? \text{ mol S} = 4.07 \cancel{\text{g S}} \times \frac{1 \text{ mol S}}{32.07 \cancel{\text{g S}}} = 0.127 \text{ mol S}$$

(b) The conversion mol S $\rightarrow$ atoms S is carried out using the Avogadro constant as a conversion factor.

$$? \text{ atoms S} = 0.127 \cancel{\text{mol S}} \times \frac{6.022 \times 10^{23} \text{ atoms S}}{1 \cancel{\text{mol S}}} = 7.64 \times 10^{22} \text{ atoms S}$$

Assess

By including units in our calculations, we can check that proper cancellation of units occurs. Also, if our only concern is to calculate the number of sulfur atoms in the sample, the calculations carried out in parts **(a)** and **(b)** could be combined into a single calculation, as shown below.

$$? \text{ atoms S} = 4.07 \cancel{\text{g S}} \times \frac{1 \cancel{\text{mol S}}}{32.07 \cancel{\text{g S}}} \times \frac{6.022 \times 10^{23} \text{ atoms S}}{1 \cancel{\text{mol S}}} = 7.64 \times 10^{22} \text{ atoms S}$$

Had we rounded 4.07 g S $\times$ (1 mol S/32.07 g S) to 0.127 mol S and used the rounded result in part (b), we would have obtained a final answer of 7.65×10^{22} atoms S. With a single line calculation, we do not have to write down an intermediate result and we avoid round-off errors.

PRACTICE EXAMPLE A: What is the mass of 2.35×10^{24} atoms of Cu?

PRACTICE EXAMPLE B: How many lead-206 atoms are present in a 22.6 g sample of lead metal? [*Hint:* See Figure 2-16.]

EXAMPLE 2-9 Combining Several Factors in a Calculation—Molar Mass, the Avogadro Constant, Percent Abundances

Potassium-40 is one of the few naturally occurring radioactive isotopes of elements of low atomic number. Its percent natural abundance among K isotopes is 0.012%. How many ^{40}K atoms are present in 225 mL of whole milk containing 1.65 mg K/mL?

Analyze

Ultimately we need to complete the conversion mL milk $\rightarrow$ atoms ^{40}K. There is no single conversion factor that allows us to complete this conversion in one step, so we anticipate having to complete several steps or conversions. We are told the milk contains 1.65 mg K/mL $= 1.65 \times 10^{-3}$ g K/mL, and this information can be used to carry out the conversion mL milk $\rightarrow$ g K. We can carry out the conversions g K $\rightarrow$ mol K $\rightarrow$ atoms K by using conversion factors based on the molar mass of K and the Avogadro constant. The final conversion, atoms K $\rightarrow$ atoms ^{40}K, can be carried out by using a conversion factor based on the percent natural abundance of ^{40}K. A complete conversion pathway is shown below:

$$\text{mL milk} \rightarrow \text{mg K} \rightarrow \text{g K} \rightarrow \text{mol K} \rightarrow \text{atoms K} \rightarrow \text{atoms } ^{40}\text{K}$$

Solve

The required conversions can be carried out in a stepwise fashion, or they can be combined into a single line calculation. Let's use a stepwise approach. First, we convert from mL milk to g K.

$$?\text{ g K} = 225\text{ mL milk} \times \frac{1.65\text{ mg K}}{1\text{ mL milk}} \times \frac{1\text{ g K}}{1000\text{ mg K}} = 0.371\text{ g K}$$

Next, we convert from g K to mol K,

$$?\text{ mol K} = 0.371\text{ g K} \times \frac{1\text{ mol K}}{39.10\text{ g K}} = 9.49 \times 10^{-3}\text{ mol K}$$

and then we convert from mol K to atoms K.

$$?\text{ atoms K} = 9.49 \times 10^{-3}\text{ mol K} \times \frac{6.022 \times 10^{23}\text{ atoms K}}{1\text{ mol K}} = 5.71 \times 10^{21}\text{ atoms K}$$

Finally, we convert from atoms K to atoms ^{40}K.

$$?\text{ atoms } ^{40}\text{K} = 5.71 \times 10^{21}\text{ atoms K} \times \frac{0.012\text{ atoms } ^{40}\text{K}}{100\text{ atoms K}} = 6.9 \times 10^{17}\text{ atoms } ^{40}\text{K}$$

Assess

The final answer is rounded to two significant figures because the least precisely known quantity in the calculation, the percent natural abundance of ^{40}K, has two significant figures. It is possible to combine the steps above into a single line calculation.

$$?\text{ atoms } ^{40}\text{K} = 225\text{ mL milk} \times \frac{1.65\text{ mg K}}{1\text{ mL milk}} \times \frac{1\text{ g K}}{1000\text{ mg K}} \times \frac{1\text{ mol K}}{39.10\text{ g K}}$$

$$\times \frac{6.022 \times 10^{23}\text{ atoms K}}{1\text{ mol K}} \times \frac{0.012\text{ atoms } ^{40}\text{K}}{100\text{ atoms K}}$$

$$= 6.9 \times 10^{17}\text{ atoms } ^{40}\text{K}$$

PRACTICE EXAMPLE A: How many Pb atoms are present in a small piece of lead with a volume of 0.105 cm^3? The density of Pb $= 11.34\text{ g/cm}^3$.

PRACTICE EXAMPLE B: Rhenium-187 is a radioactive isotope that can be used to determine the age of meteorites. A 0.100 mg sample of Re contains 2.02×10^{17} atoms of ^{187}Re. What is the percent abundance of rhenium-187 in the sample?