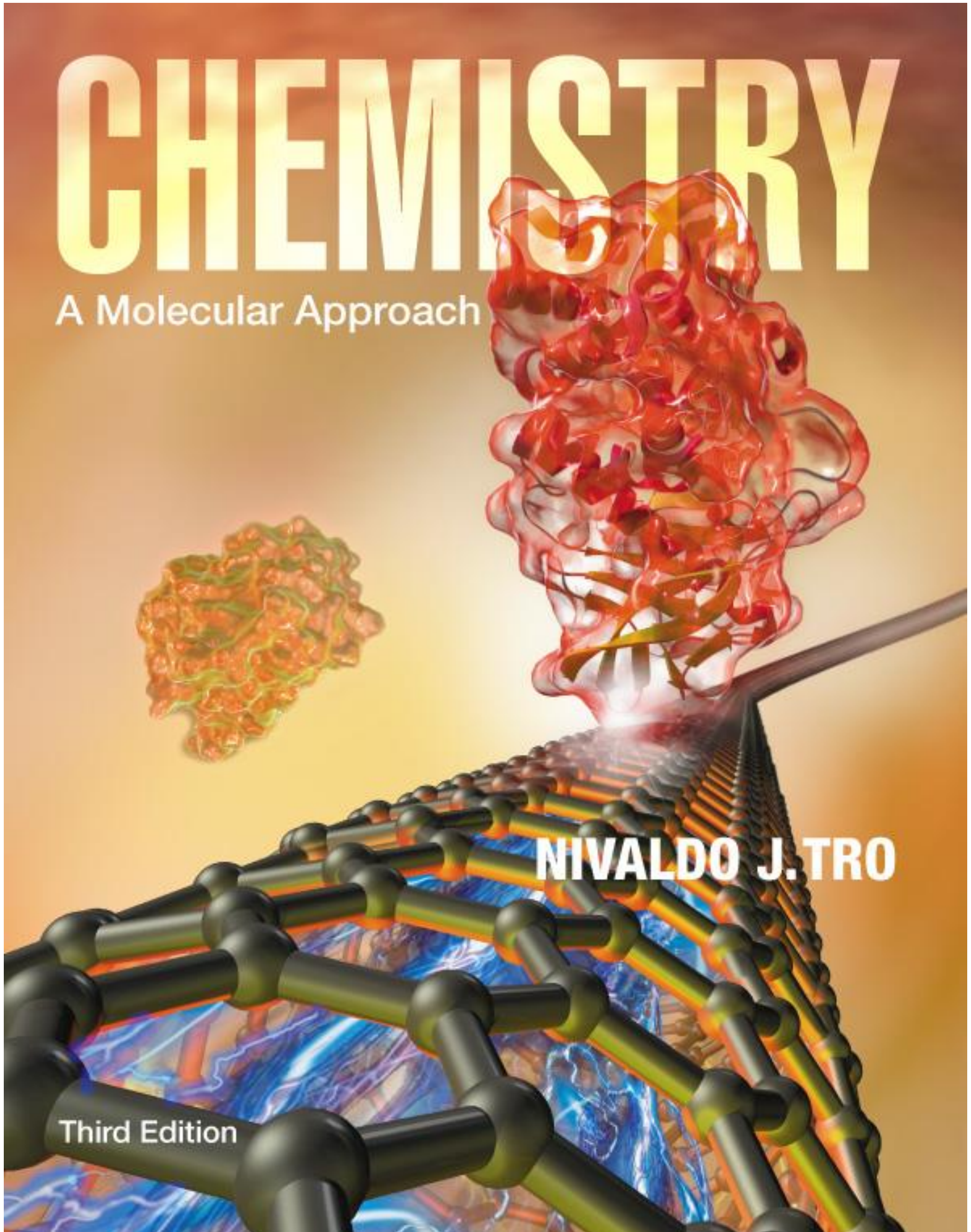


Lecture Presentation

Chapter 2

Atoms and Elements

Christian Madu, Ph.D.
Collin College

- “Continuity” How are we connected: to the Earth, to each other, and to the Universe?
- https://www.youtube.com/watch?v=XGK84Poeynk
- https://www.youtube.com/watch?v=yqLlgIaz1L0
- https://www.youtube.com/watch?v=-4Us5PTb4J8

Atomic Theory Timeline

- **Purpose:** To organize important historical details that led to our understanding of the modern day atom

- **Instruction:**

1. Individual Work
2. Include Dalton, Thomson, Rutherford, and Bohr
3. Include description of their experiment that lead to their contribution
4. Include present day model/research

If You Cut a Piece of Graphite

- If you cut a piece of graphite from the tip of a pencil into smaller and smaller pieces, how far could you go? Could you divide it forever?

- Cutting the graphite from a pencil tip into smaller and smaller pieces (far smaller than the eye could see), would eventually end up with individual carbon atoms.

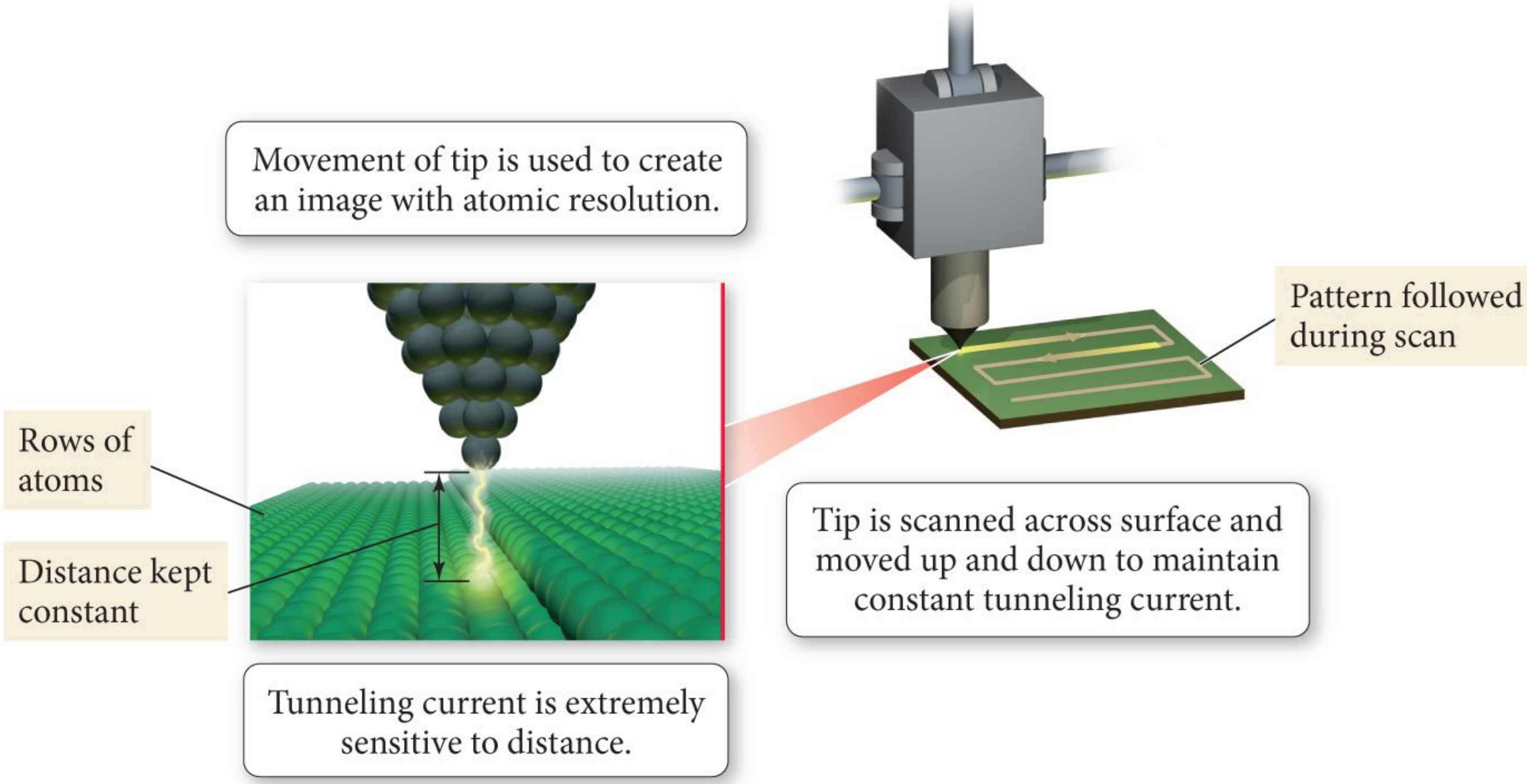
If You Cut a Piece of Graphite

- The word atom comes from the Greek ***atomos***, meaning “indivisible.”
- You cannot divide a carbon atom into smaller pieces and still have carbon.
- Atoms compose all ordinary matter—if you want to understand matter, you must begin by understanding atoms.

Imaging and Moving Individual Atoms

- On March 16, 1981, Gerd Binnig and Heinrich Rohrer worked late into the night in their laboratory.
- Their work led to the development of *scanning tunneling microscopy (STM)*.
- STM is a technique that can image, and even move, individual atoms and molecules.

Scanning Tunneling Microscopy



- Binnig and Rohrer developed a type of microscope that could “see” atoms.

Imaging and Moving Individual Atoms

- In spite of their small size, atoms are the key to connecting the macroscopic and microscopic worlds.
- An ***atom*** is the smallest identifiable unit of an *element*.
- There are about
  - 91 different naturally occurring elements, and
  - over 20 synthetic elements (elements not found in nature).

Early Ideas about the Building Blocks of Matter

- Leucippus (fifth century B.C.) and his student Democritus (460–370 B.C.) were first to propose that matter was composed of small, indestructible particles.
  - Democritus wrote, "Nothing exists except atoms and empty space; everything else is opinion."
- They proposed that many different kinds of atoms existed, each different in shape and size, and that they moved randomly through empty space.

Early Building Blocks of Matter Ideas

- Plato and Aristotle did not embrace the atomic ideas of Leucippus and Democritus.

- They held that
  – matter had no smallest parts.
  – different substances were composed of various proportions of fire, air, earth, and water.

Early Building Blocks of Matter Ideas

- Later scientific approach became the established way to learn about the physical world.
- An English chemist, John Dalton (1766–1844) offered convincing evidence that supported the early atomic ideas of Leucippus and Democritus.

Modern Atomic Theory and the Laws That Led to It

- The theory that all matter is composed of atoms grew out of observations and laws.

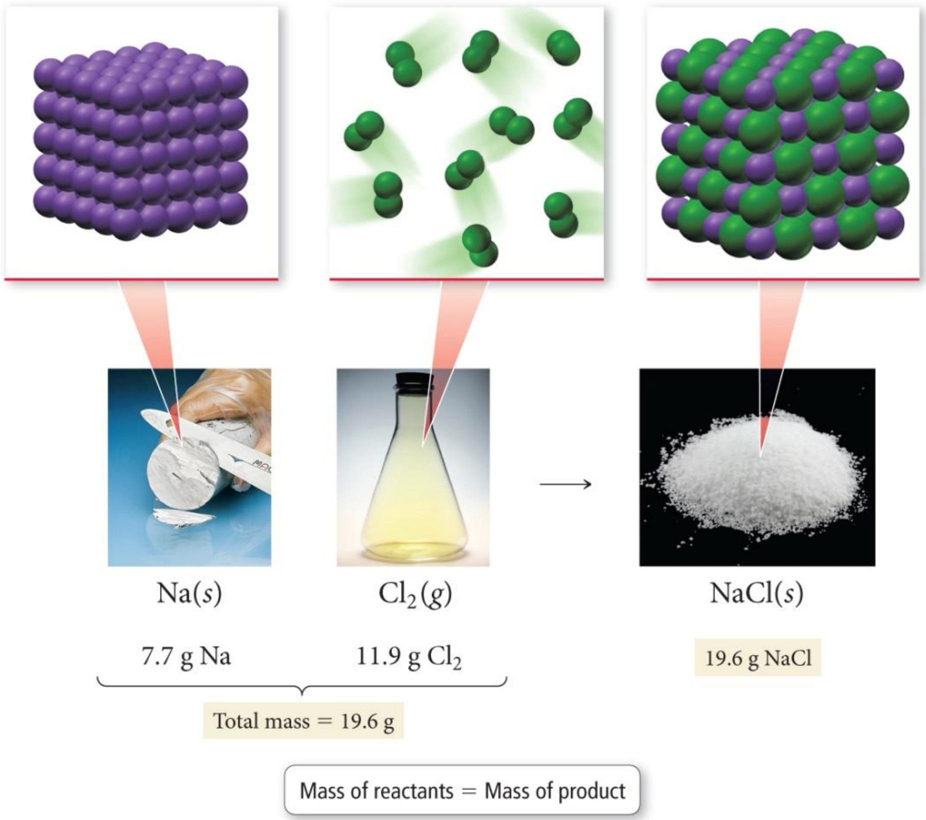
- The three most important laws that led to the development and acceptance of the atomic theory are as follows:
  - The law of conservation of mass
  - The law of definite proportions
  - The law of multiple proportions

The Law of Conservation of Mass

- Antoine Lavoisier formulated the **law of conservation of mass**, which states the following:
  - ***In a chemical reaction, matter is neither created nor destroyed.***

- Hence, when a chemical reaction occurs, the total mass of the substances involved in the reaction does not change.

The Law of Conservation of Mass



- This law is consistent with the idea that matter is composed of small, indestructible particles.

The Law of Definite Proportions

- In 1797, a French chemist, Joseph Proust made observations on the composition of compounds.
- He summarized his observations in the **law of definite proportions**:
  - ***All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.***

The Law of Definite Proportions

- The law of definite proportions is sometimes called the law of constant composition.
  - For example, the decomposition of 18.0 g of water results in 16.0 g of oxygen and 2.0 g of hydrogen, or an oxygen-to-hydrogen mass ratio of:

$$\text{Mass ratio} = \frac{16.0 \text{ g O}}{2.0 \text{ g H}} = 8.0 \text{ or } 8{:}1$$

Example 2.1 Law of Definite Proportions

Two samples of carbon dioxide are decomposed into their constituent elements. One sample produces 25.6 g of oxygen and 9.60 g of carbon, and the other produces 21.6 g of oxygen and 8.10 g of carbon. Show that these results are consistent with the law of definite proportions.

Solution

To show this, calculate the mass ratio of one element to the other for both samples by dividing the mass of one element by the mass of the other. For convenience, divide the larger mass by the smaller one.

For the first sample:

$$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{25.6}{9.60} = 2.67 \text{ or } 2.67{:}1$$

For the second sample:

$$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{21.6}{8.10} = 2.67 \text{ or } 2.67{:}1$$

The ratios are the same for the two samples, so these results are consistent with the law of definite proportions.

For Practice 2.1

Two samples of carbon monoxide are decomposed into their constituent elements. One sample produces 17.2 g of oxygen and 12.9 g of carbon, and the other sample produces 10.5 g of oxygen and 7.88 g of carbon. Show that these results are consistent with the law of definite proportions.

The Law of Multiple Proportions

- In 1804, John Dalton published his **law of multiple proportions.**
  - *When two elements (call them A and B) form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.*
- An atom of A combines with either one, two, three, or more atoms of B (AB_1, AB_2, AB_3, etc.).

The Law of Multiple Proportions

- Carbon monoxide and carbon dioxide are two compounds composed of the same two elements: carbon and oxygen.
  - The mass ratio of oxygen to carbon in carbon dioxide is 2.67:1; therefore, 2.67 g of oxygen reacts with 1 g of carbon.
  - In carbon monoxide, however, the mass ratio of oxygen to carbon is 1.33:1, or 1.33 g of oxygen to every 1 g of carbon.

Carbon dioxide

Mass oxygen that combines with 1 g carbon = 2.67 g

Carbon monoxide

Mass oxygen that combines with 1 g carbon = 1.33 g

- The ratio of these two masses is itself a small whole number.

$$\frac{\text{Mass oxygen to 1 g carbon in carbon dioxide}}{\text{Mass oxygen to 1 g carbon in carbon monoxide}} = \frac{2.67}{1.33} = 2$$

Example 2.2 Law of Multiple Proportions

Nitrogen forms several compounds with oxygen, including nitrogen dioxide and dinitrogen monoxide. Nitrogen dioxide contains 2.28 g oxygen to every 1.00 g nitrogen, while dinitrogen monoxide contains 0.570 g oxygen to every 1.00 g nitrogen. Show that these results are consistent with the law of multiple proportions.

Solution

To show this, calculate the ratio of the mass of oxygen from one compound to the mass of oxygen in the other. Always divide the larger of the two masses by the smaller one.

$$\frac{\begin{array}{c}\text{Mass oxygen to 1 g}\\ \text{nitrogen in nitrogen dioxide}\end{array}}{\begin{array}{c}\text{Mass oxygen to 1 g}\\ \text{nitrogen in dinitrogen monoxide}\end{array}} = \frac{2.28}{0.570} = 4.00$$

The ratio is a small whole number (4); these results are consistent with the law of multiple proportions.

For Practice 2.2

Hydrogen and oxygen form both water and hydrogen peroxide. The decomposition of a sample of water forms 0.125 g hydrogen to every 1.00 g oxygen. The decomposition of a sample of hydrogen peroxide forms 0.250 g hydrogen to every 1.00 g oxygen. Show that these results are consistent with the law of multiple proportions.

John Dalton and the Atomic Theory

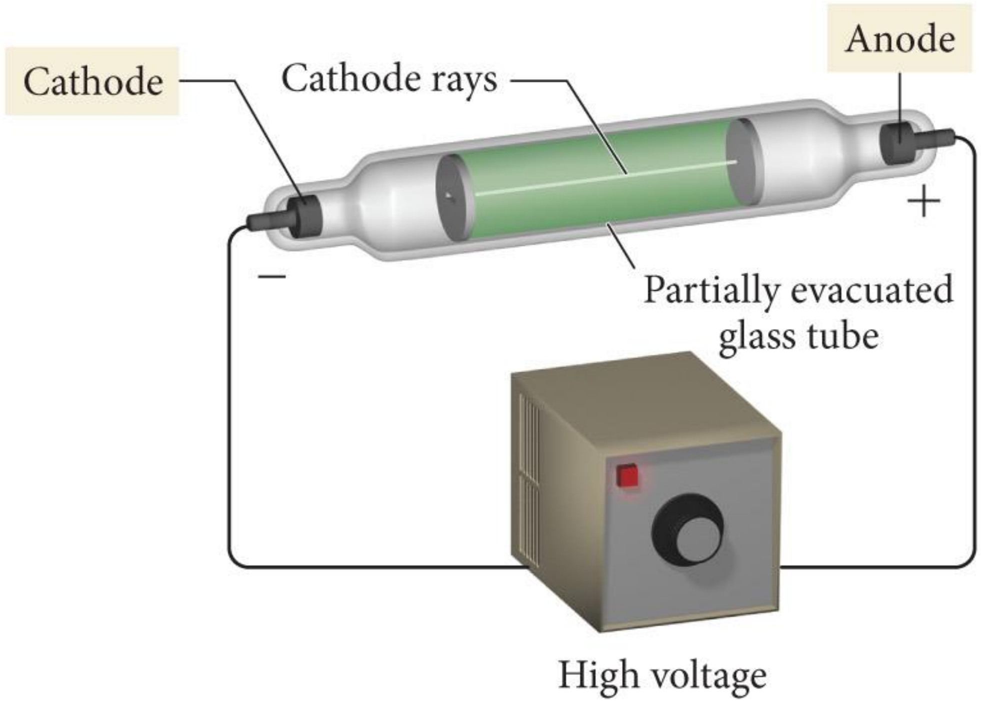
- Dalton's **atomic theory** explained the laws as follows**:**
  1. Each element is composed of tiny, indestructible particles called atoms.
  2. All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements.
  3. Atoms combine in simple, whole-number ratios to form compounds.
  4. Atoms of one element cannot change into atoms of another element. In a chemical reaction, atoms only change the way that they are bound together with other atoms.

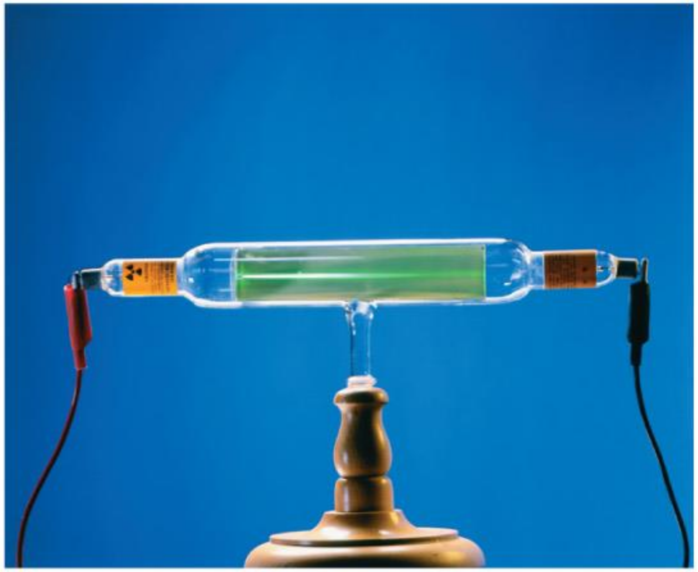
The Discovery of the Electron

- J. J. Thomson (1856–1940) **cathode rays** experiments
- Thomson constructed a partially evacuated glass tube called a **cathode ray tube.**
- He found that a beam of particles, called cathode rays, traveled from the negatively charged electrode (called the cathode) to the positively charged one (called the anode).

- https://www.youtube.com/watch?v=2xKZRpAsWL8
- **Thomson's Cathode Ray**

The Discovery of the Electron





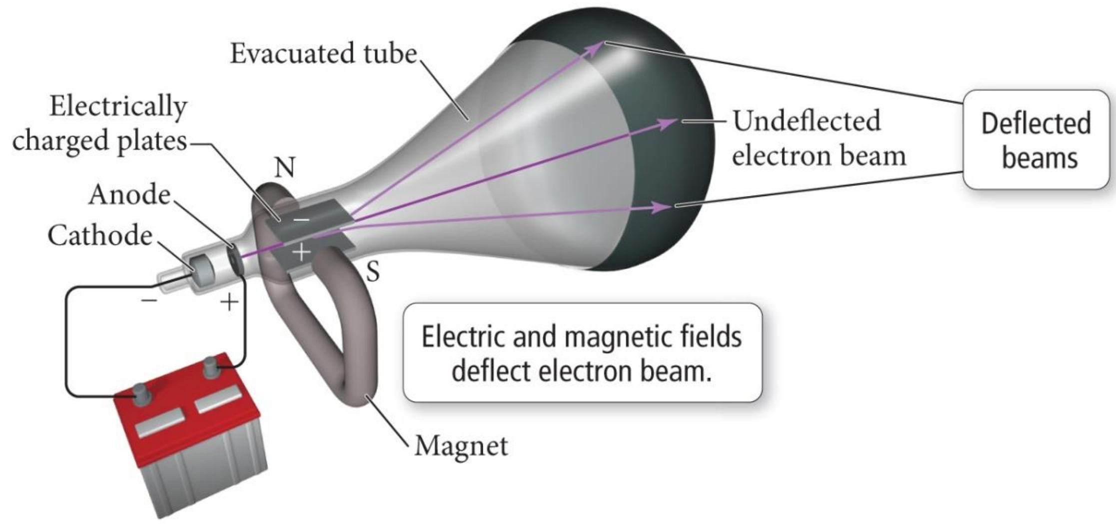
The Discovery of the Electron

- Thomson found that the particles that compose the cathode ray have the following properties:
  - They travel in straight lines.
  - They are independent of the composition of the material from which they originate (the cathode).
  - They carry a negative **electrical charge**.

The Discovery of the Electron

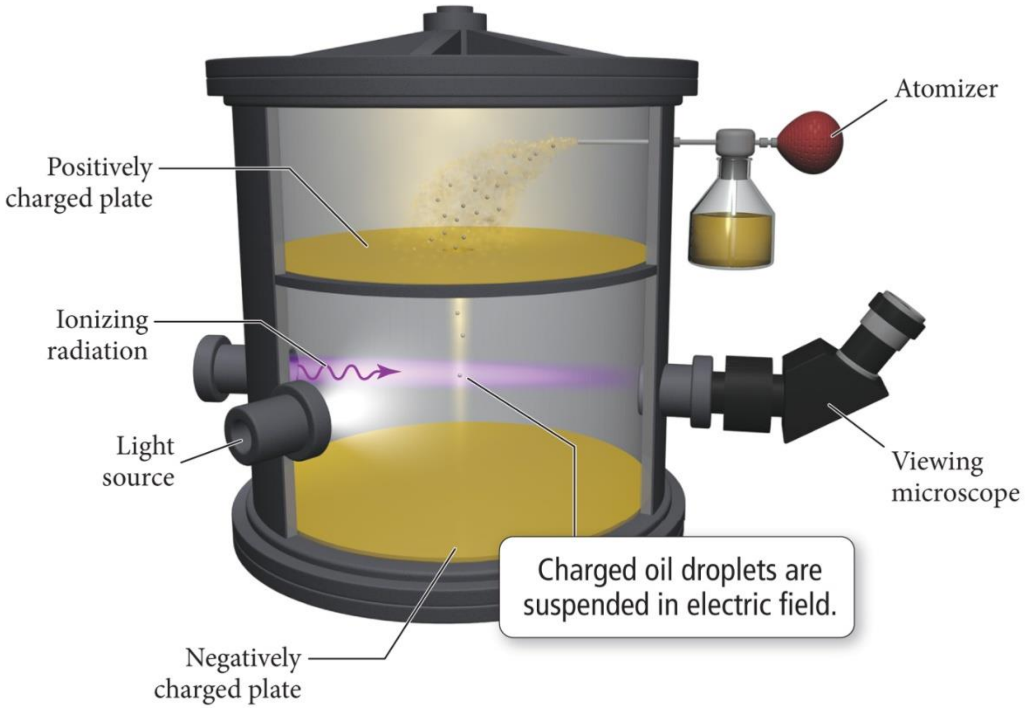
- J. J. Thomson measured the charge-to-mass ratio of the cathode ray particles by deflecting them using electric and magnetic fields, as shown in the figure.
- The value he measured was -1.76×10^3 coulombs (C) per gram.

Charge-to-Mass Ratio of the Electron



Millikan's Oil Drop Experiment: The Charge of the Electron

- American physicist Robert Millikan (1868–1953), performed his now famous oil drop experiment in which he deduced the charge of a single electron.



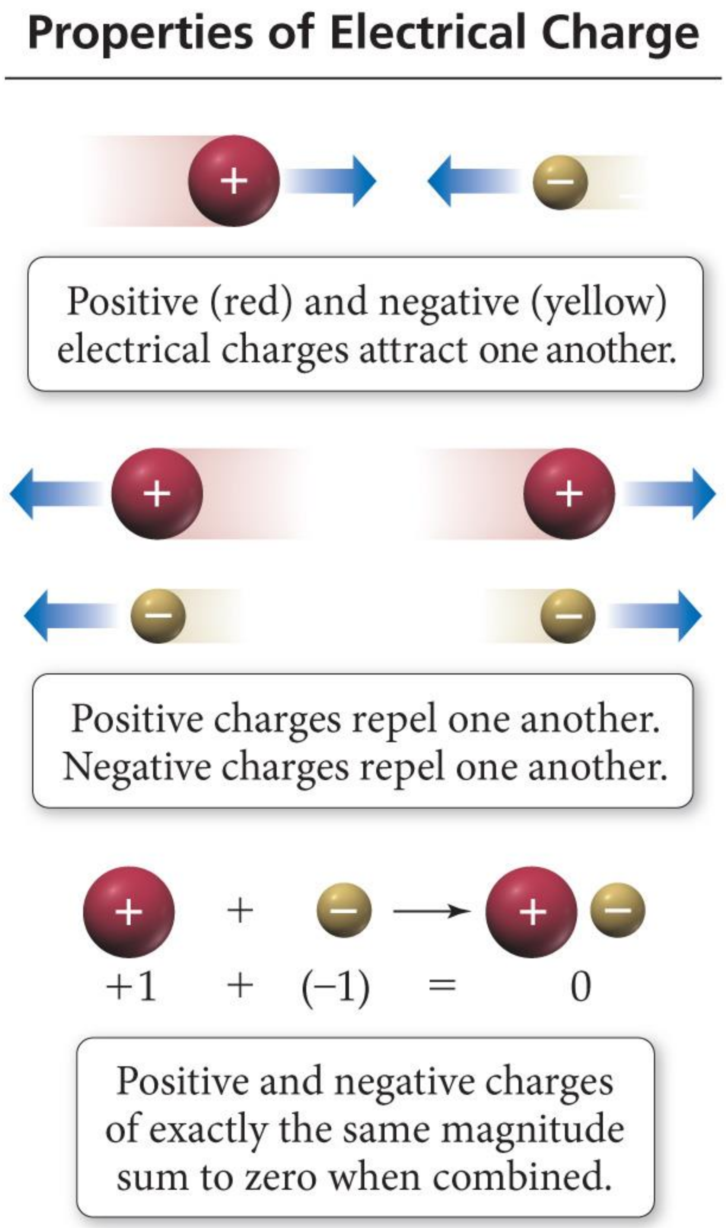
- https://www.youtube.com/watch?v=ijHKu6iXiRk
- **Millikan's Oil Drop Experiment**

Millikan's Oil Drop Experiment

- By measuring the strength of the electric field required to halt the free fall of the drops, and by figuring out the masses of the drops themselves (determined from their radii and density), Millikan calculated the charge of each drop.
- The measured charge on any drop was always a whole-number multiple of -1.96×10^{-19}, the fundamental charge of a single electron.

The Discovery of the Electron

- J. J. Thomson had discovered the **electron**, a negatively charged, low mass particle present within all atoms.



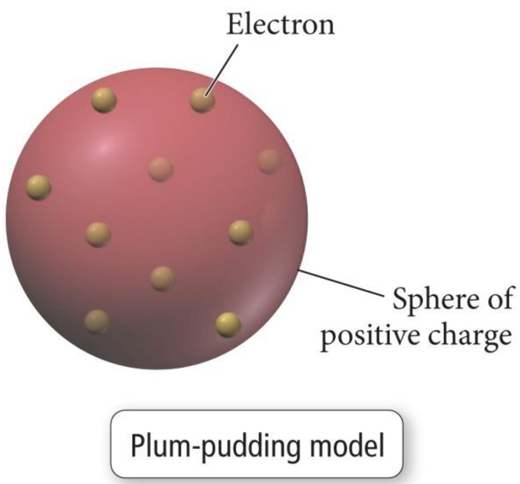
Millikan's Oil Drop Experiment

- With this number in hand, and knowing Thomson's mass-to-charge ratio for electrons, we can deduce the mass of an electron:

$$\cancel{\text{Charge}} \times \frac{\text{mass}}{\cancel{\text{charge}}} = \text{mass}$$

The Structure of the Atom

- J. J. Thomson proposed that the negatively charged electrons were small particles held within a positively charged sphere.



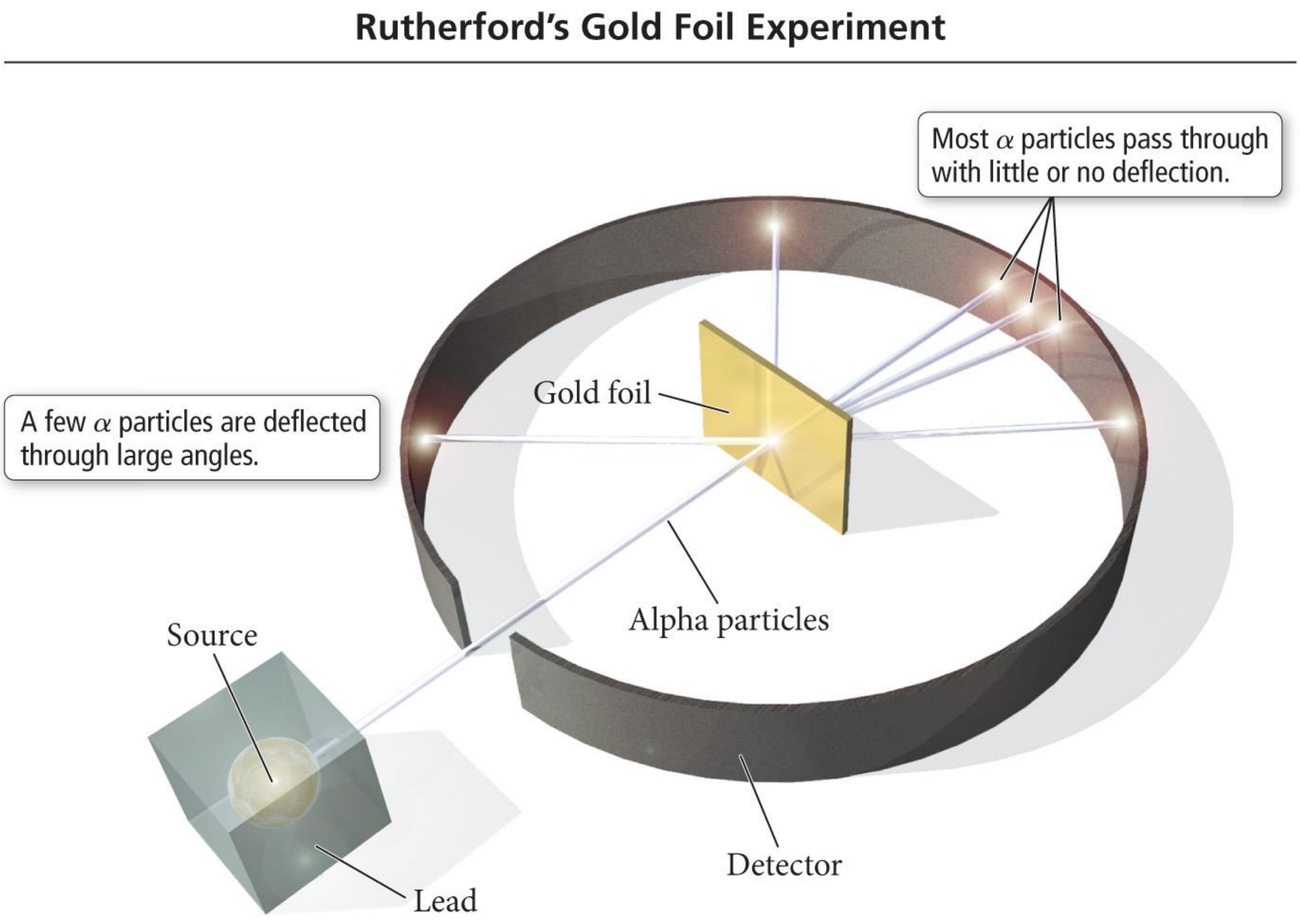
- This model, the most popular of the time, became known as the plum-pudding model.

Rutherford's Gold Foil Experiment

- In 1909, Ernest Rutherford (1871–1937), who had worked under Thomson and subscribed to his plum-pudding model, performed an experiment in an attempt to confirm Thomson's model.
- In the experiment, Rutherford directed the positively charged particles at an ultra thin sheet of gold foil.

- https://www.youtube.com/watch?v=XBqHkraf8iE
- **Rutherford's Gold Foil Experiment**

Rutherford’s Gold Foil Experiment

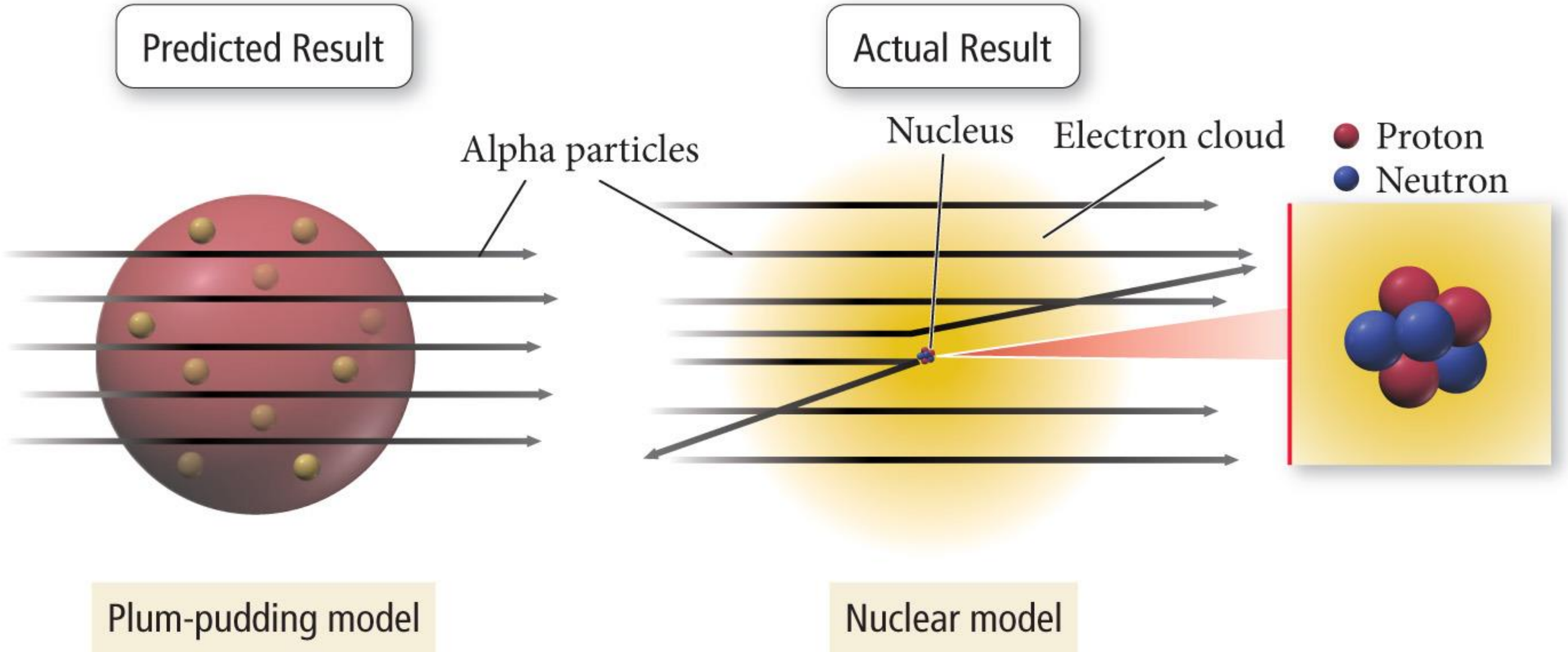


Rutherford's Gold Foil Experiment

- The Rutherford experiment gave an unexpected result. A majority of the particles did pass directly through the foil, but some particles were deflected, and some (approximately 1 in 20,000) even bounced back.

- Rutherford created a new model—a modern version of which is shown in Figure 2.7 alongside the plum-pudding model—to explain his results.

Rutherford’s Gold Foil Experiment



- He concluded that matter must not be as uniform as it appears. It must contain large regions of empty space dotted with small regions of very dense matter.

Rutherford's Gold Foil Experiment

- Building on this idea, he proposed the **nuclear theory** of the atom, with three basic parts:
  1. Most of the atom's mass and all of its positive charge are contained in a small core called a **nucleus**.
  2. Most of the volume of the atom is empty space, throughout which tiny, negatively charged electrons are dispersed.
  3. There are as many negatively charged electrons outside the nucleus as there are positively charged particles (named **protons**) within the nucleus, so that the atom is electrically neutral.

The Neutrons

- Although Rutherford's model was highly successful, scientists realized that it was incomplete.

- Later work by Rutherford and one of his students, British scientist James Chadwick (1891–1974), demonstrated that the previously unaccounted for mass was due to **neutrons**, neutral particles within the nucleus.

The Neutrons

- The mass of a neutron is similar to that of a proton.
- However, a neutron has no electrical charge.
  - The helium atom is four times as massive as the hydrogen atom because
    - it contains two protons
    - *and two neutrons.*
- Hydrogen, on the other hand, contains only one proton and no neutrons.

Subatomic Particles

- All atoms are composed of the same subatomic particles:
  - Protons
  - Neutrons
  - Electrons
- Protons and neutrons, as we saw earlier, have nearly identical masses.
  - The mass of the proton is 1.67262×10^{-27} kg.
  - The mass of the neutron is 1.67493×10^{-27} kg.
  - The mass of the electron is 9.1×10^{-31} kg.

Subatomic Particles

TABLE 2.1 Subatomic Particles

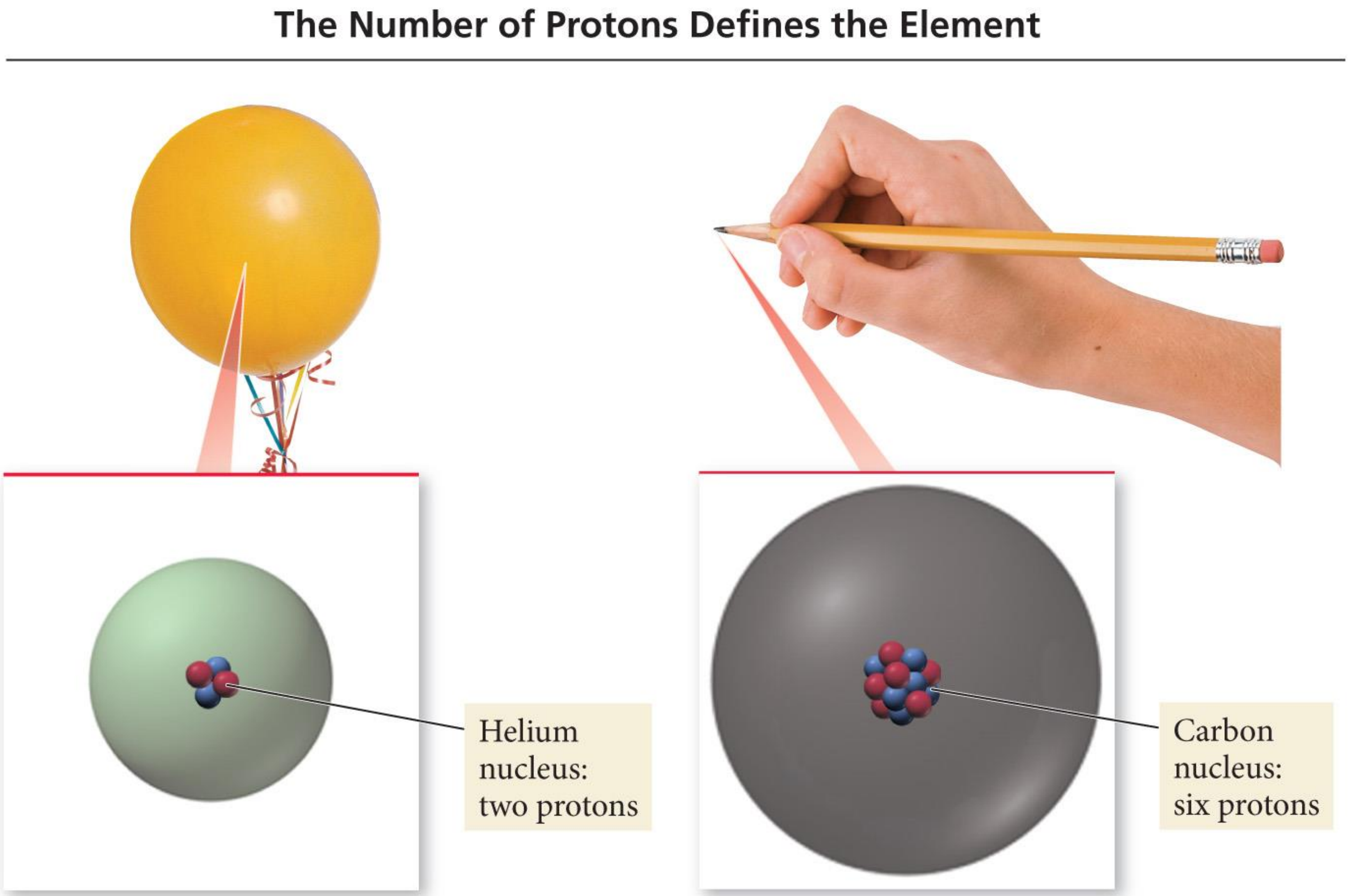
	Mass (kg)	Mass (amu)	Charge (relative)	Charge (C)
Proton	1.67262×10^{-27}	1.00727	$+1$	$+1.60218 \times 10^{-19}$
Neutron	1.67493×10^{-27}	1.00866	0	0
Electron	0.00091×10^{-27}	0.00055	-1	-1.60218×10^{-19}

- *The charge of the proton and the electron are equal in magnitude but opposite in sign*. The neutron has no charge.

Elements: Defined by Their Numbers of Protons

- The most important number to the identity of an atom is the number of protons in its nucleus.
- *The number of protons defines the element*.
- The number of protons in an atom's nucleus is its **atomic number** and is given the symbol ***Z***.

Elements: Defined by Their Numbers of Protons



Periodic Table

The Periodic Table

4 Be beryllium	
4	Atomic number (Z)
Be	Chemical symbol
beryllium	Name

1 H hydrogen																	2 He helium
3 Li lithium	4 Be beryllium											5 B boron	6 C carbon	7 N nitrogen	8 O oxygen	9 F fluorine	10 Ne neon
11 Na sodium	12 Mg magnesium											13 Al aluminum	14 Si silicon	15 P phosphorus	16 S sulfur	17 Cl chlorine	18 Ar argon
19 K potassium	20 Ca calcium	21 Sc scandium	22 Ti titanium	23 V vanadium	24 Cr chromium	25 Mn manganese	26 Fe iron	27 Co cobalt	28 Ni nickel	29 Cu copper	30 Zn zinc	31 Ga gallium	32 Ge germanium	33 As arsenic	34 Se selenium	35 Br bromine	36 Kr krypton
37 Rb rubidium	38 Sr strontium	39 Y yttrium	40 Zr zirconium	41 Nb niobium	42 Mo molybdenum	43 Tc technetium	44 Ru ruthenium	45 Rh rhodium	46 Pd palladium	47 Ag silver	48 Cd cadmium	49 In indium	50 Sn tin	51 Sb antimony	52 Te tellurium	53 I iodine	54 Xe xenon
55 Cs cesium	56 Ba barium	57 La lanthanum	72 Hf hafnium	73 Ta tantalum	74 W tungsten	75 Re rhenium	76 Os osmium	77 Ir iridium	78 Pt platinum	79 Au gold	80 Hg mercury	81 Tl thallium	82 Pb lead	83 Bi bismuth	84 Po polonium	85 At astatine	86 Rn radon
87 Fr francium	88 Ra radium	89 Ac actinium	104 Rf rutherfordium	105 Db dubnium	106 Sg seaborgium	107 Bh bohrium	108 Hs hassium	109 Mt meitnerium	110 Ds darmstadtium	111 Rg roentgenium	112 Cn copernicium	113 **	114 Fl flerovium	115 **	116 Lv livermorium	117 **	118 **

58 Ce cerium	59 Pr praseodymium	60 Nd neodymium	61 Pm promethium	62 Sm samarium	63 Eu europium	64 Gd gadolinium	65 Tb terbium	66 Dy dysprosium	67 Ho holmium	68 Er erbium	69 Tm thulium	70 Yb ytterbium	71 Lu lutetium
90 Th thorium	91 Pa protactinium	92 U uranium	93 Np neptunium	94 Pu plutonium	95 Am americium	96 Cm curium	97 Bk berkelium	98 Cf californium	99 Es einsteinium	100 Fm fermium	101 Md mendelevium	102 No nobelium	103 Lr lawrencium

Periodic Table

- Each element is identified by a unique atomic number and with a unique **chemical symbol.**

- The chemical symbol is either a one- or two-letter abbreviation listed directly below its atomic number on the periodic table.
  - The chemical symbol for helium is He.
  - The chemical symbol for carbon is C.
  - The chemical symbol for Nitrogen is N.

Isotopes: Varied Number of Neutrons

- All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons.
    - For example, all neon atoms contain 10 protons, but they may contain 10, 11, or 12 neutrons. All three types of neon atoms exist, and each has a slightly different mass.
- Atoms with the same number of protons but a different number of neutrons are called **isotopes**.

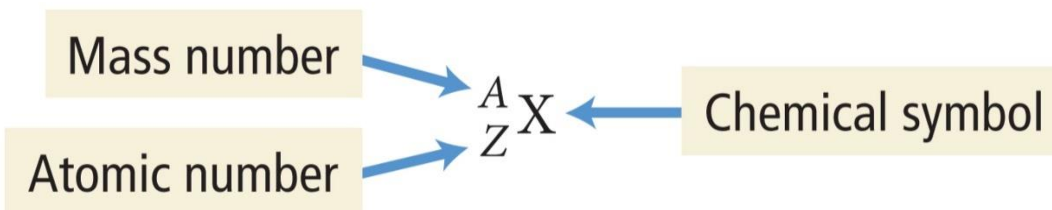
Isotopes: Varied Number of Neutrons

- The relative amount of each different isotope in a naturally occurring sample of a given element is roughly constant.
- These percentages are called the **natural abundance** of the isotopes.
  - Advances in mass spectrometry have allowed accurate measurements that reveal small but significant variations in the natural abundance of isotopes for many elements.

Isotopes: Varied Number of Neutrons

- The sum of the number of neutrons and protons in an atom is its **mass number** and is represented by the symbol ***A***

A = number of protons (p) + number of neutrons (n)



- where X is the chemical symbol, *A* is the mass number, and *Z* is the atomic number.

$$^{20}_{10}\mathrm{Ne} \quad ^{21}_{10}\mathrm{Ne} \quad ^{22}_{10}\mathrm{Ne}$$

Isotopes: Varied Number of Neutrons

- A second common notation for isotopes is the chemical symbol (or chemical name) followed by a dash and the mass number of the isotope.

Chemical symbol or name → X-*A* ← Mass number

Ne-20	Ne-21	Ne-22
neon-20	neon-21	neon-22

Isotopes: Varied Number of Neutrons

Symbol	Number of Protons	Number of Neutrons	*A* (Mass Number)	Natural Abundance (%)
Ne−20 or $^{20}_{10}Ne$	10	10	20	90.48
Ne−21 or $^{21}_{10}Ne$	10	11	21	0.27
Ne−22 or $^{22}_{10}Ne$	10	12	22	9.25

Ions: Losing and Gaining Electrons

- The number of electrons in a neutral atom is equal to the number of protons in its nucleus (designated by its atomic number *Z*).
- In a chemical changes, however, atoms can lose or gain electrons and become charged particles called **ions**.
  - Positively charged ions, such as Na^+, are called **cations.**
  - Negatively charged ions, such as F^-, are called **anions**.

Example 2.3 Atomic Numbers, Mass Numbers, and Isotope Symbols

a. What are the atomic number (Z), mass number (A), and symbol of the chlorine isotope with 18 neutrons?

b. How many protons, electrons, and neutrons are present in an atom of $^{52}_{24}\text{Cr}$?

Solution

a. Look up the atomic number (Z) for chlorine on the periodic table. The atomic number specifies the number of protons.

Z = 17, so chlorine has 17 protons.

The mass number (A) for an isotope is the sum of the number of protons and the number of neutrons.

$$\begin{aligned} A &= \text{number of protons} + \text{number of neutrons} \\ &= 17 + 18 = 35 \end{aligned}$$

The symbol for an isotope is its two-letter abbreviation with the atomic number (Z) in the lower left corner and the mass number (A) in the upper left corner.

$$^{35}_{17}\text{Cl}$$

b. For any isotope (in this case $^{52}_{24}\text{Cr}$) the number of protons is indicated by the atomic number located at the lower left. Since this is a neutral atom, the number of electrons equals the number of protons.

Number of protons = Z = 24
Number of electrons = 24 (neutral atom)

The number of neutrons is equal to the mass number (upper left) minus the atomic number (lower left).

Number of neutrons = 52 – 24 = 28

Example 2.3 Atomic Numbers, Mass Numbers, and Isotope Symbols

Continued

For Practice 2.3

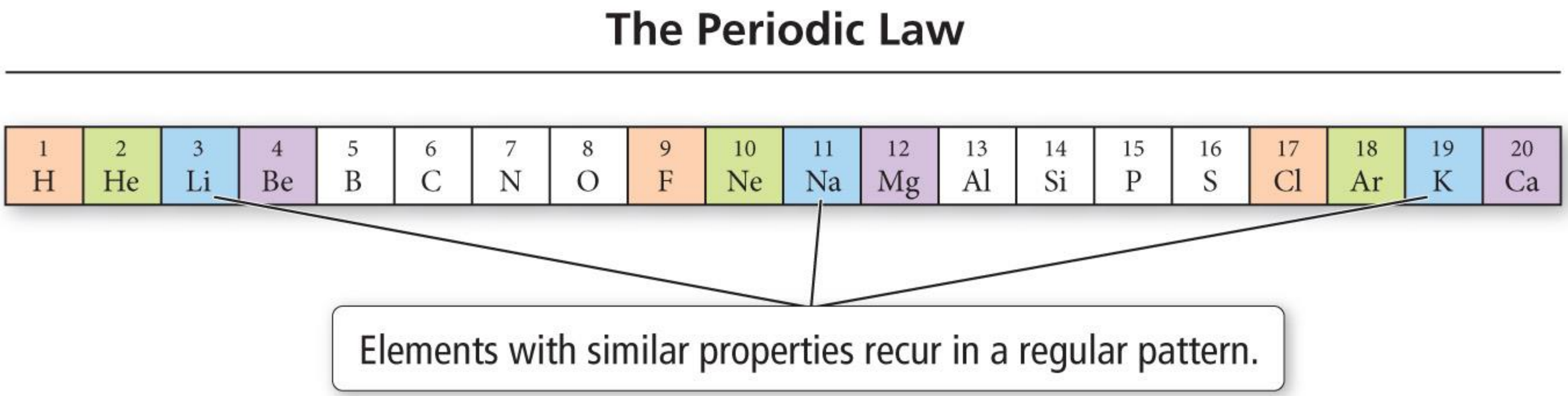
a. What are the atomic number, mass number, and symbol for the carbon isotope with seven neutrons?

b. How many protons and neutrons are present in an atom of $^{39}_{19}K$?

Finding Patterns: The Periodic Law and the Periodic Table

- In 1869, Mendeleev noticed that certain groups of elements had similar properties.
- He found that when elements are listed in order of increasing mass, these similar properties recurred in a periodic pattern.
  - To be periodic means to exhibit a repeating pattern.

The Periodic Law

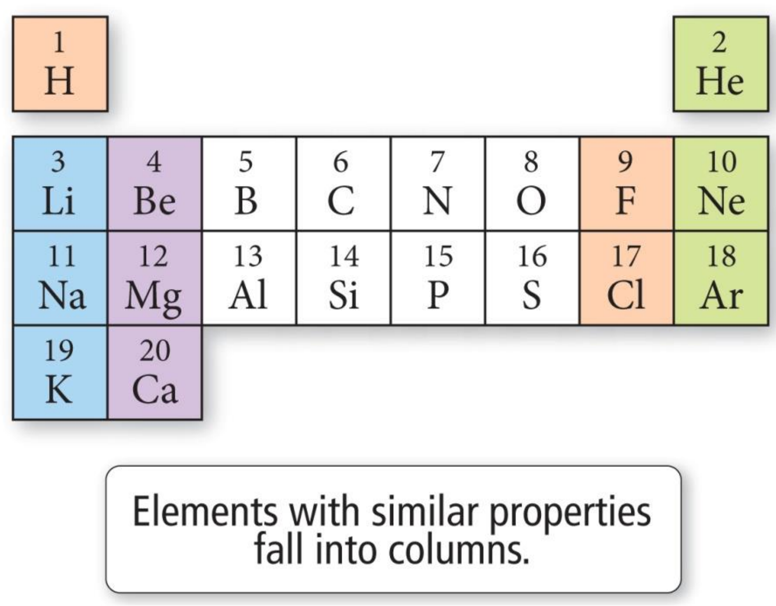


- Mendeleev summarized these observations in the **periodic law**:
  - **When the elements are arranged in order of increasing mass, certain sets of properties recur periodically.**

Periodic Table

- Mendeleev organized the known elements in a table.
- He arranged the rows so that elements with similar properties fall in the same vertical columns.

A Simple Periodic Table



Periodic Table

- Mendeleev’s table contained some gaps, which allowed him to predict the existence (and even the properties) of yet undiscovered elements.
  - Mendeleev predicted the existence of an element he called eka-silicon.
  - In 1886, eka-silicon was discovered by German chemist Clemens Winkler (1838–1904), who named it germanium.

Modern Periodic Table

- In the modern table, elements are listed in order of increasing atomic number rather than increasing relative mass.
- The modern periodic table also contains more elements than Mendeleev’s original table because more have been discovered since his time.

Modern Periodic Table

Major Divisions of the Periodic Table

Metals | Metalloids | Nonmetals

Silicon

Arsenic

Carbon

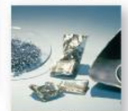
Strontium

Chromium

Gold

Copper

Lead



Sulfur

Bromine

Iodine

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118

Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Classification of Elements

- Elements in the periodic table are classified as the following:
  - Metals
  - Nonmetals
  - Metalloids

Metals

- **Metals** lie on the lower left side and middle of the periodic table and share some common properties:
    - They are good conductors of heat and electricity.
    - They can be pounded into flat sheets (malleability).
    - They can be drawn into wires (ductility).
    - They are often shiny.
    - They tend to lose electrons when they undergo chemical changes.
  - Chromium, copper, strontium, and lead are typical metals.

Nonmetals

- **Nonmetals** lie on the upper right side of the periodic table.
- There are a total of 17 nonmetals:
  - Five are solids at room temperature (C, P, S, Se, and I)
  - One is a liquid at room temperature (Br)
  - Eleven are gases at room temperature (H, He, N, O, F, Ne, Cl, Ar, Kr, Xe, and Rn)

Nonmetals

- Nonmetals as a whole tend to
  - be poor conductors of heat and electricity.
  - be not ductile and not malleable.
  - gain electrons when they undergo chemical changes.

Oxygen, carbon, sulfur, bromine, and iodine are nonmetals.

Metalloids

- Metalloids are sometimes called semimetals.
- They are elements that lie along the zigzag diagonal line that divides metals and nonmetals.
- They exhibit mixed properties.
- Several metalloids are also classified as **semiconductors** because of their intermediate (and highly temperature-dependent) electrical conductivity.

Periodic Table

- The periodic table can also be divided into
  - **main-group elements**, whose properties tend to be largely predictable based on their position in the periodic table.
  - **transition elements** or **transition metals**, whose properties tend to be less predictable based simply on their position in the periodic table.

Periodic Table

Main-group elements | Transition elements | Main-group elements

Group number

Periods	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118

Periodic Table

- The periodic table is divided into vertical columns and horizontal rows.
  - Each vertical column is called a group (or family).
  - Each horizontal row is called a period.
- There are a total of 18 groups and 7 periods.
- The groups are numbered 1–18 (or the A and B grouping).

Periodic Table

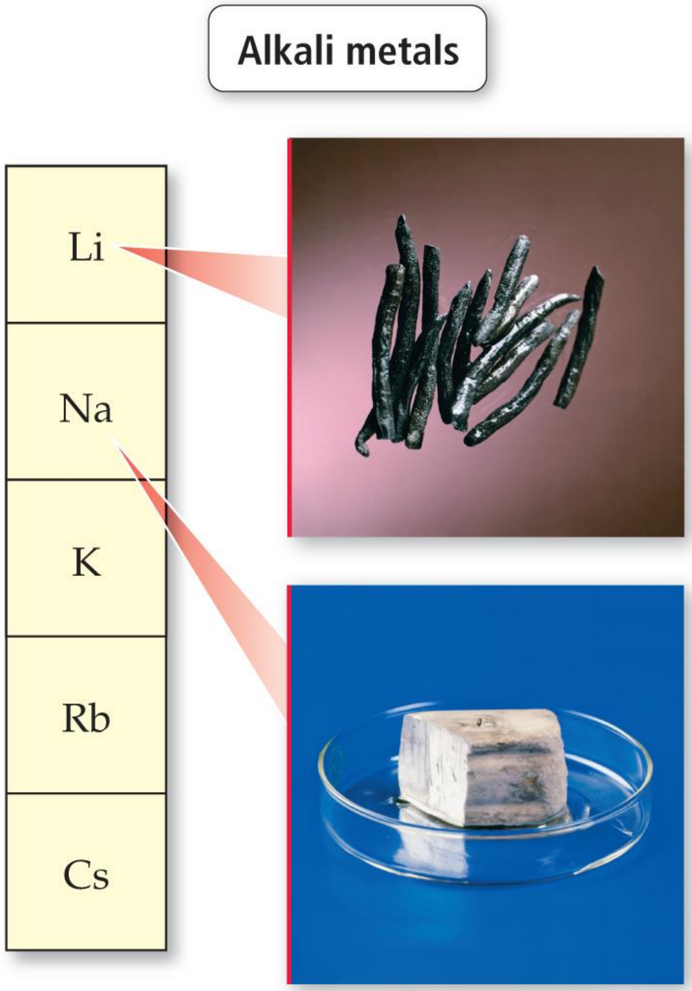
- Main-group elements are in columns labeled with a number and the letter A (1A–8A or groups 1, 2, and 13–18).
- Transition elements are in columns labeled with a number and the letter B (or groups 3–12).

Noble Gas

- The elements within a group usually have similar properties.
- The group 8A elements, called the **noble gases**, are mostly unreactive.
  - The most familiar noble gas is probably helium, used to fill buoyant balloons. Helium is chemically stable—it does not combine with other elements to form compounds—and is therefore safe to put into balloons.
  - Other noble gases are neon (often used in electronic signs), argon (a small component of our atmosphere), krypton, and xenon.

Alkali

- The group 1A elements, called the **alkali metals**, are all reactive metals.
- A marble-sized piece of sodium explodes violently when dropped into water.
- Lithium, potassium, and rubidium are also alkali metals.

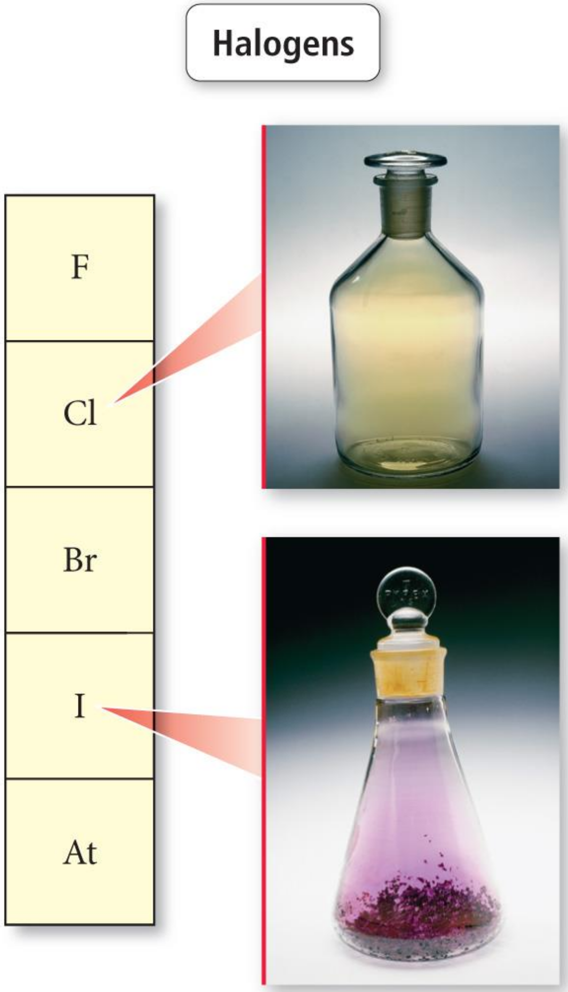


Alkaline Earth Metals

- The group 2A elements are called the **alkaline earth metals**.
- They are fairly reactive, but not quite as reactive as the alkali metals.
  - Calcium, for example, reacts fairly vigorously with water.
  - Other alkaline earth metals include magnesium (a common low-density structural metal), strontium, and barium.

Halogens

- The group 7A elements, the **halogens**, are very reactive nonmetals.
- They are always found in nature as a salt.
  - Chlorine, a greenish-yellow gas with a pungent odor
  - Bromine, a red-brown liquid that easily evaporates into a gas
  - Iodine, a purple solid
  - Fluorine, a pale-yellow gas



Ions and the Periodic Table

- A main-group metal tends to lose electrons, forming a cation with the same number of electrons as the nearest noble gas.
- A main-group nonmetal tends to gain electrons, forming an anion with the same number of electrons as the nearest noble gas.

Ions and the Periodic Table

- In general, the alkali metals (group 1A) have a tendency to lose one electron and form 1+ ions.
- The alkaline earth metals (group 2A) tend to lose two electrons and form 2+ ions.
- The halogens (group 7A) tend to gain one electron and form 1– ions.
- The oxygen family nonmetals (group 6A) tend to gain two electrons and form 2– ions.

Ions and the Periodic Table

- For the main-group elements that form cations with predictable charge, the charge is equal to the group number.
- For main-group elements that form anions with predictable charge, the charge is equal to the group number minus eight.
- Transition elements may form various different ions with different charges.

Example 2.4 Predicting the Charge of Ions

Predict the charges of the monoatomic (single atom) ions formed by these main-group elements.

a. Al

b. S

Solution

a. Aluminum is a main-group metal and tends to lose electrons to form a cation with the same number of electrons as the nearest noble gas. Aluminum atoms have 13 electrons and the nearest noble gas is neon, which has 10 electrons. Aluminum therefore loses 3 electrons to form a cation with a 3+ charge (Al^{3+}).

b. Sulfur is a nonmetal and tends to gain electrons to form an anion with the same number of electrons as the nearest noble gas. Sulfur atoms have 16 electrons and the nearest noble gas is argon, which has 18 electrons. Sulfur therefore gains 2 electrons to form an anion with a 2– charge (S^{2-}).

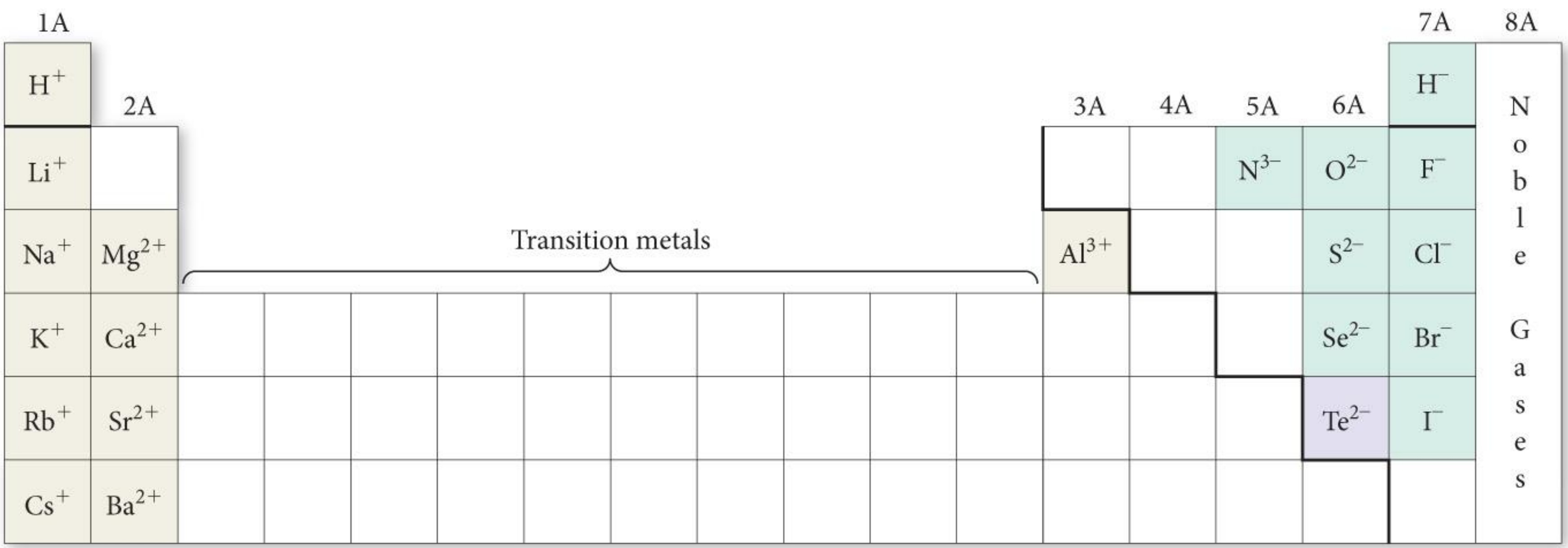
For Practice 2.4

a. N

b. Rb

Ions and the Periodic Table

Elements That Form Ions with Predictable Charges



Atomic Mass: The Average Mass of an Element's Atoms

- Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*
- The atomic mass of each element is directly beneath the element's symbol in the periodic table.
- It represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope*.

Atomic Mass

- Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). We can calculate its atomic mass:

- Solution:
  - Convert the percent abundance to decimal form and multiply it with its isotopic mass:

    Cl-37 = 0.2423(36.97 amu) = 8.9578 amu

    Cl-35 = 0.7577(34.97 amu) = 26.4968 amu

    Atomic Mass Cl = 8.9578 + 26.4968 = 35.45 amu

Atomic Mass

17
Cl
35.45
chlorine

- In general, we calculate the atomic mass with the equation:

$$
\begin{aligned}
\text{Atomic mass} &= \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n) \\
&= (\text{fraction of isotope 1} \times \text{mass of isotope 1}) \\
&+ (\text{fraction of isotope 2} \times \text{mass of isotope 2}) \\
&+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) + \ldots
\end{aligned}
$$

Example 2.5 Atomic Mass

Copper has two naturally occurring isotopes: Cu-63 with a mass of 62.9396 amu and a natural abundance of 69.17%, and Cu-65 with a mass of 64.9278 amu and a natural abundance of 30.83%. Calculate the atomic mass of copper.

Solution

Convert the percent natural abundances into decimal form by dividing by 100.

$$\text{Fraction Cu-63} = \frac{69.17}{100} = 0.6917$$

$$\text{Fraction Cu-65} = \frac{30.83}{100} = 0.3083$$

Calculate the atomic mass using the equation given in the text.

$$\begin{aligned}\text{Atomic mass} &= 0.6917(62.9396\text{ amu}) + 0.3083(64.9278\text{ amu})\\ &= 43.5\underline{3}53\text{ amu} + 20.0\underline{1}72\text{ amu} = 63.5\underline{5}25 = 63.55\text{ amu}\end{aligned}$$

For Practice 2.5

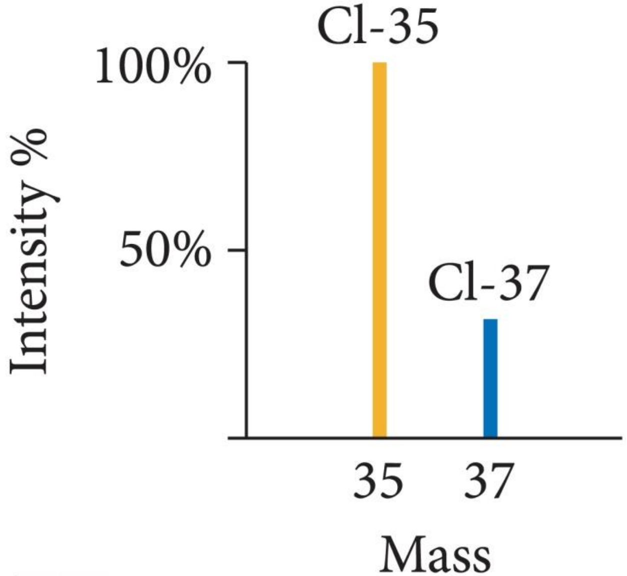
Magnesium has three naturally occurring isotopes with masses of 23.99 amu, 24.99 amu, and 25.98 amu and natural abundances of 78.99%, 10.00%, and 11.01%, respectively. Calculate the atomic mass of magnesium.

For More Practice 2.5

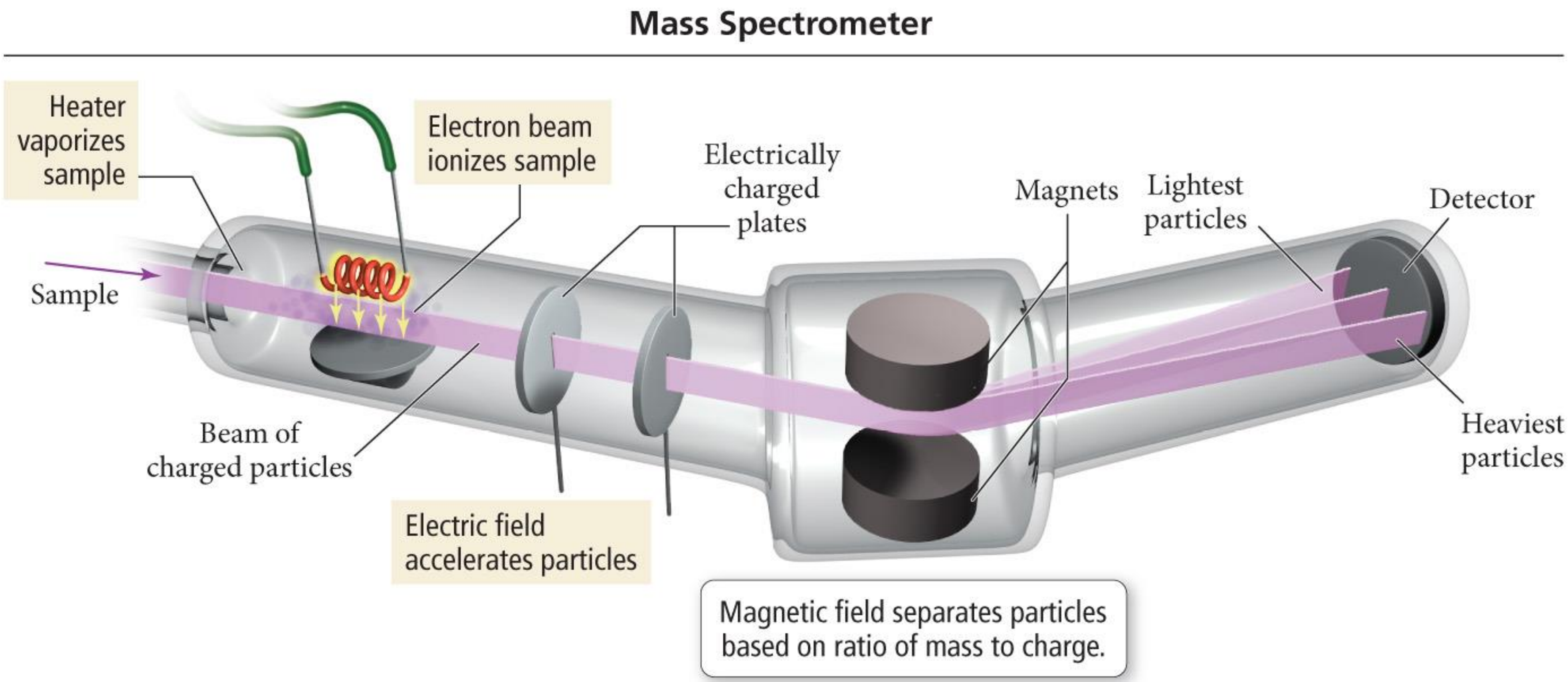
Gallium has two naturally occurring isotopes: Ga-69 with a mass of 68.9256 amu and a natural abundance of 60.11%, and Ga-71. Use the atomic mass of gallium from the periodic table to find the mass of Ga-71.

Mass Spectrometry: Measuring the Mass of Atoms and Molecules

- The masses of atoms and the percent abundances of isotopes of elements are measured using **mass spectrometry**—a technique that separates particles according to their mass.



Mass Spectrometry



Molar Mass: Counting Atoms by Weighing Them

- As chemists, we often need to know the number of atoms in a sample of a given mass. Why? *Because chemical processes happen between particles*.
- Therefore, if we want to know the number of atoms in anything of ordinary size, we count them by weighing.

The Mole: A Chemist's "Dozen"

- When we count large numbers of objects, we often use units such as
  - 1 dozen objects = 12 objects.
  - 1 gross objects = 144 objects.
- The chemist's "dozen" is the **mole** (abbreviated mol). A mole is the measure of material containing 6.02214×10^{23} particles:

$$1 \text{ mole} = 6.02214 \times 10^{23} \text{ particles}$$

- This number is **Avogadro's number**.

The Mole

- First thing to understand about the mole is that it can specify Avogadro's number of anything.
- For example, 1 mol of marbles corresponds to 6.02214×10^{23} marbles.
- 1 mol of sand grains corresponds to 6.02214×10^{23} sand grains.
- *One mole of anything is* 6.02214×10^{23} *units of that thing.*

The Mole

- The second, and more fundamental, thing to understand about the mole is how it gets its specific value.
- **The value of the mole is equal to the number of atoms in exactly 12 grams of pure C-12.**
- **12 g C = 1 mol C atoms = 6.022×10^{23} C atoms**

Converting between Number of Moles and Number of Atoms

- Converting between number of moles and number of atoms is similar to converting between dozens of eggs and number of eggs.
- For atoms, you use the conversion factor 1 mol atoms = 6.022×10^{23} atoms.
- The conversion factors take the following forms:

$$\frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}} \quad \text{or} \quad \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}}$$

Example 2.6 Converting between Number of Moles and Number of Atoms

Calculate the number of copper atoms in 2.45 mol of copper.

Sort

You are given the amount of copper in moles and asked to find the number of copper atoms.

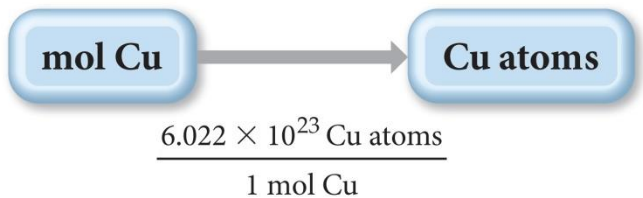
Given: 2.45 mol Cu

Find: Cu atoms

Strategize

Convert between number of moles and number of atoms by using Avogadro's number as a conversion factor.

Conceptual Plan



Relationships Used

$$6.022 \times 10^{23} = 1 \text{ mol (Avogadro's number)}$$

Solve

Follow the conceptual plan to solve the problem. Begin with 2.45 mol Cu and multiply by Avogadro's number to get to the number of Cu atoms.

Example 2.6 Converting between Number of Moles and Number of Atoms

Continued

Solution

$$2.45 \cancel{\text{mol Cu}} \times \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \cancel{\text{mol Cu}}} = 1.48 \times 10^{24} \text{ Cu atoms}$$

Check

Since atoms are small, it makes sense that the answer is large. The given number of moles of copper is almost 2.5, so the number of atoms is almost 2.5 times Avogadro's number.

For Practice 2.6

A pure silver ring contains 2.80×10^{22} silver atoms. How many moles of silver atoms does it contain?

Converting between Mass and Amount (Number of Moles)

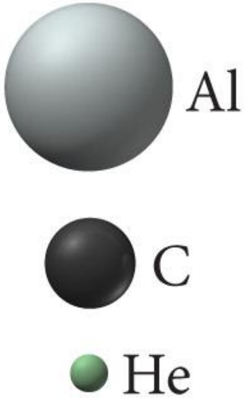
- To count atoms by weighing them, we need one other conversion factor—the mass of 1 mol of atoms.
- The mass of 1 mol of atoms of an element is the **molar mass**.
- **An element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units (amu).**

Converting between Mass and Moles

26.98 g aluminum = 1 mol aluminum = 6.022×10^{23} Al atoms

12.01 g carbon = 1 mol carbon = 6.022×10^{23} C atoms

4.003 g helium = 1 mol helium = 6.022×10^{23} He atoms



- The lighter the atom, the less mass in 1 mol of atoms.

Converting between Mass and Moles

- The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element. For carbon,

$$12.01 \text{ g C} = 1 \text{ mol C} \text{ or } \frac{12.01 \text{ g C}}{\text{mol C}} \text{ or } \frac{1 \text{ mol C}}{12.01 \text{ g C}}$$

Example 2.7 Converting between Mass and Amount (Number of Moles)

Calculate the amount of carbon (in moles) contained in a 0.0265 g pencil “lead.” (Assume that the pencil lead is made of pure graphite, a form of carbon.)

Sort

You are given the mass of carbon and asked to find the amount of carbon in moles.

Given: 0.0265 g C

Find: mol C

Strategize

Convert between mass and amount (in moles) of an element by using the molar mass of the element.

Conceptual Plan

$$\text{g C} \xrightarrow[\frac{1\ \text{mol}}{12.01\ \text{g}}]{} \text{mol C}$$

Relationships Used

$$12.01\ \text{g C} = 1\ \text{mol C (carbon molar mass)}$$

Solve

Follow the conceptual plan to solve the problem.

Example 2.8 The Mole Concept—Converting between Mass and Number of Atoms

Continued

Solve

Follow the conceptual plan to solve the problem. Begin with 3.10 g Cu and multiply by the appropriate conversion factors to arrive at the number of Cu atoms.

Solution

$$3.10\ \cancel{\text{g Cu}} \times \frac{1\ \cancel{\text{mol Cu}}}{63.55\ \cancel{\text{g Cu}}} \times \frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \cancel{\text{mol Cu}}} = 2.94 \times 10^{22}\ \text{Cu atoms}$$

Check

The answer (the number of copper atoms) is less than 6.022×10^{23} (1 mole). This is consistent with the given mass of copper atoms, which is less than the molar mass of copper.

For Practice 2.8

How many carbon atoms are there in a 1.3-carat diamond? Diamonds are a form of pure carbon. (1 carat = 0.20 grams)

For More Practice 2.8

Calculate the mass of 2.25×10^{22} tungsten atoms.

Example 2.9 The Mole Concept

An aluminum sphere contains 8.55×10^{22} aluminum atoms. What is the sphere's radius in centimeters? The density of aluminum is 2.70 g/cm^3.

Sort

You are given the number of aluminum atoms in a sphere and the density of aluminum. You are asked to find the radius of the sphere.

Given: 8.55×10^{22} Al atoms
Find: radius (r) of sphere

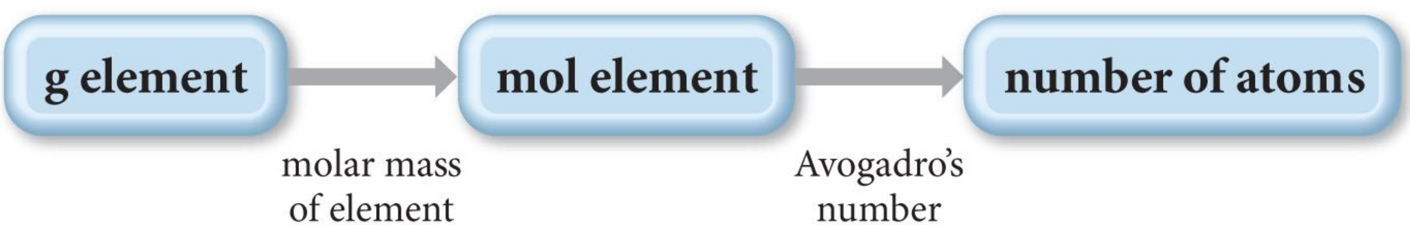
Strategize

The heart of this problem is density, which relates mass to volume; though you aren't given the mass directly, you are given the number of atoms, which you can use to find mass.

1. Convert from number of atoms to number of moles using Avogadro's number as a conversion factor.
2. Convert from number of moles to mass using molar mass as a conversion factor.
3. Convert from mass to volume (in cm^3) using density as a conversion factor.
4. Once you calculate the volume, find the radius from the volume using the formula for the volume of a sphere.

Conceptual Plan

- We now have all the tools to count the number of atoms in a sample of an element by weighing it.
  - First, we obtain the mass of the sample.
  - Then, we convert it to the amount in moles using the element’s molar mass.
  - Finally, we convert it to the number of atoms using Avogadro’s number.
- The conceptual plan for these kinds of calculations takes the following form:



Example 2.9 The Mole Concept

Continued

Conceptual Plan

$$\text{Number of Al atoms} \xrightarrow{\frac{1\ \text{mol Al}}{6.022 \times 10^{23}\ \text{Al atoms}}} \text{mol Al} \xrightarrow{\frac{26.98\ \text{g Al}}{1\ \text{mol Al}}} \text{g Al} \xrightarrow{\frac{1\ \text{cm}^3}{2.70\ \text{g Al}}} V\ (\text{in cm}^3)$$

$$V\ (\text{in cm}^3) \xrightarrow{V = \frac{4}{3}\pi r^3} r$$

Relationships Used and Equations Used

$$6.022 \times 10^{23} = 1\ \text{mol (Avogadro's number)}$$
$$26.98\ \text{g Al} = 1\ \text{mol Al (molar mass of aluminum)}$$
$$2.70\ \text{g/cm}^3\ \text{(density of aluminum)}$$
$$V = \frac{4}{3}\pi r^3\ \text{(volume of a sphere)}$$

Solve

Finally, follow the conceptual plan to solve the problem. Begin with 8.55×10^{22} Al atoms and multiply by the appropriate conversion factors to arrive at volume in cm^3.

Then solve the equation for the volume of a sphere for *r* and substitute the volume to calculate *r*.

Example 2.9 The Mole Concept

Continued

Solution

$$8.55 \times 10^{22} \cancel{\text{Al atoms}} \times \frac{1\ \cancel{\text{mol Al}}}{6.022 \times 10^{23}\ \cancel{\text{Al atoms}}}$$

$$\times \frac{26.98\ \cancel{\text{g Al}}}{1\ \cancel{\text{mol Al}}} \times \frac{1\ \text{cm}^3}{2.70\ \cancel{\text{g Al}}} = 1.41\underline{8}7\ \text{cm}^3$$

$$V = \frac{4}{3}\pi r^3$$

$$r = \sqrt[3]{\frac{3V}{4\pi}} = \sqrt[3]{\frac{3(1.41\underline{8}7\ \text{cm}^3)}{4\pi}} = 0.697\ \text{cm}$$

Check

The units of the answer (cm) are correct. The magnitude cannot be estimated accurately, but a radius of about one-half of a centimeter is reasonable for just over one-tenth of a mole of aluminum atoms.

For Practice 2.9

A titanium cube contains 2.86×10^{23} atoms. What is the edge length of the cube? The density of titanium is $4.50\ \text{g/cm}^3$.

For More Practice 2.9

Find the number of atoms in a copper rod with a length of 9.85 cm and a radius of 1.05 cm. The density of copper is $8.96\ \text{g/cm}^3$.