Unit 2: Includes the following OpenStax Sections

- 1.2 Phases and Classification of Matter
- 1.3 Physical and Chemical Properties
- 2.1 Early Ideas in Atomic Theory
- 2.2 Evolution of Atomic Theory
- 2.3 Atomic Structure and Symbolism
- 2.4 Chemical Formulas
- 2.5 The Periodic Table
- 2.6 Molecular and Ionic Compounds
- 2.7 Chemical Nomenclature

Learning Objectives

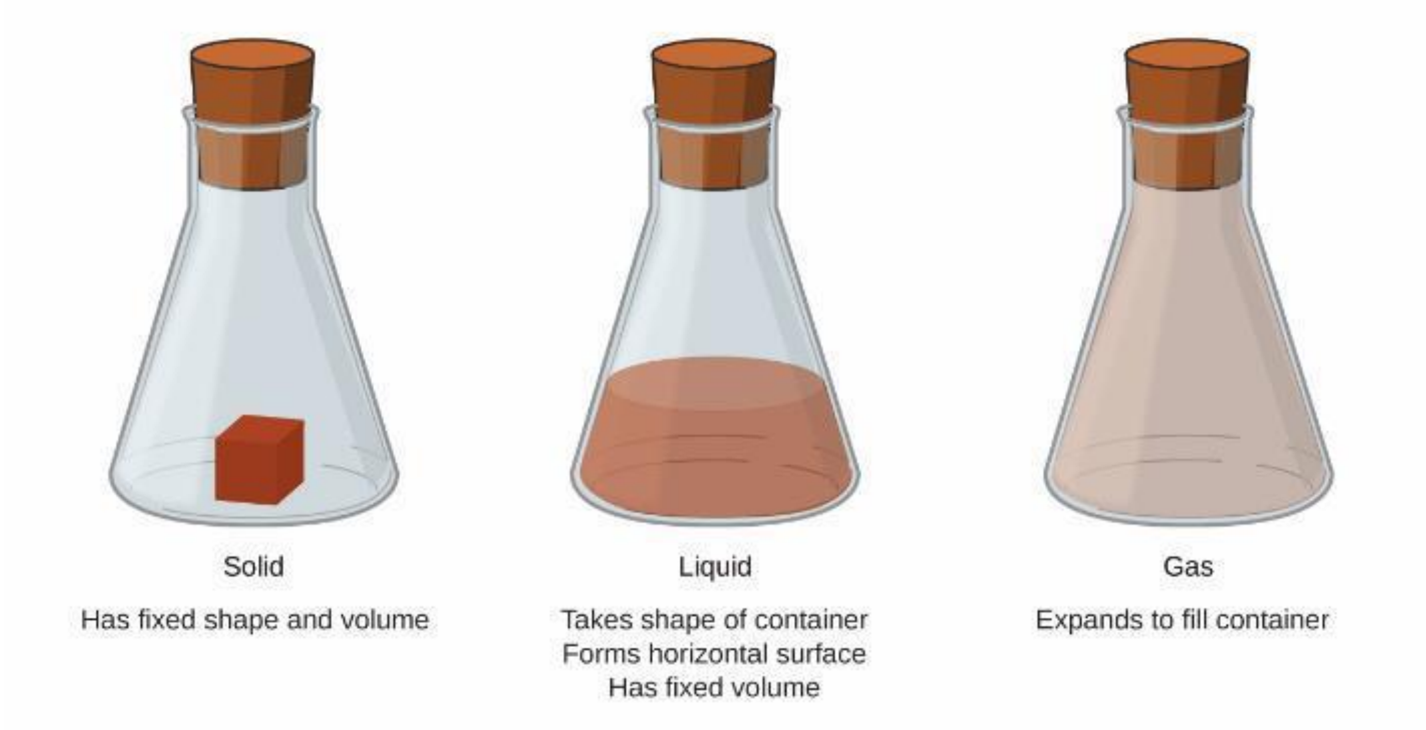
- 1.2 Phases of Matter
    - Describe the basic properties of each physical state of matter: solid, liquid, and gas
    - Distinguish between mass and weight
    - Apply the law of conservation of matter
    - Classify matter as an element, compound, homogeneous mixture, or heterogeneous mixture with regard to its physical state and composition
    - Define and give examples of atoms and molecules

Phases and Classification of Matter

- **Matter:** Anything that occupies space and has mass.

- The three most common states or phases of matter:
    1) A **solid** is rigid and possesses a definite shape.
    2) A **liquid** flows and takes the shape of its container.
    3) A **gas** takes both the shape and volume of its container.

Figure 1.6



The three most common states or phases of matter are solid, liquid, and gas.

Plasma: A Fourth State of Matter

- **Plasma:** A gaseous state of matter that contains an appreciable amount of electrically charged particles.

- Plasma has unique properties distinct from ordinary gases.

- Plasma is found in certain high-temperature environments.
  - Naturally: Stars, lightning
  - Man-made: Television screens

Figure 1.7

A plasma torch can be used to cut metal. (credit: “Hypertherm”/Wikimedia Commons)

Mass vs. Weight

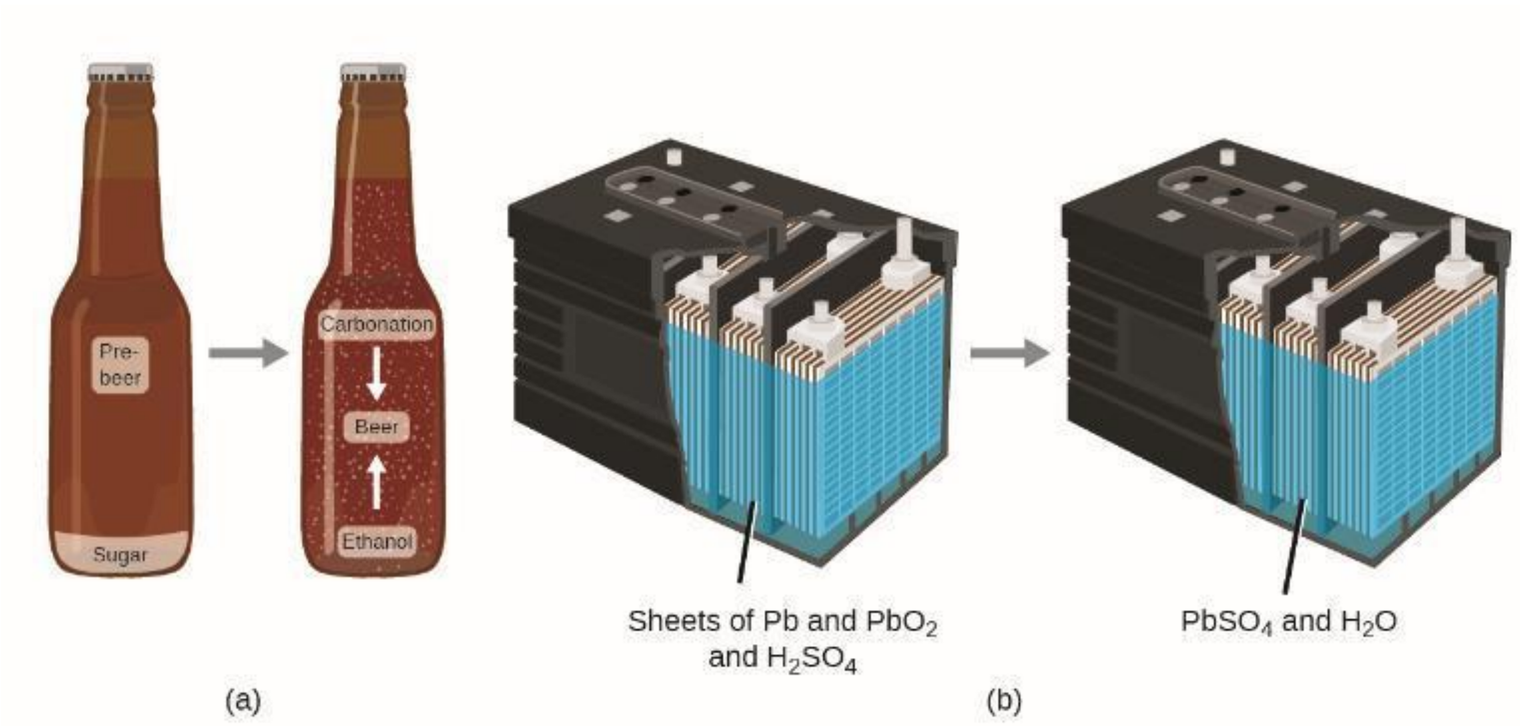
- Mass is a measure of the amount of matter in an object.
- Weight refers to the force that gravity exerts on an object.
- An object's mass is the same on the earth and the moon but its weight is different.

Law of Conservation of Matter

- **Law of conservation of matter:** There is no detectable change in the total quantity of matter present when matter converts from one type to another.

- This is true for both chemical and physical changes.

Figure 1.8



(a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbonation. (b) The mass of the lead, lead oxide plates, and sulfuric acid that goes into the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

Elements

- An **element** is a type of pure substance that cannot be broken down into simpler substances by chemical changes.

- The known elements are displayed in the periodic table.

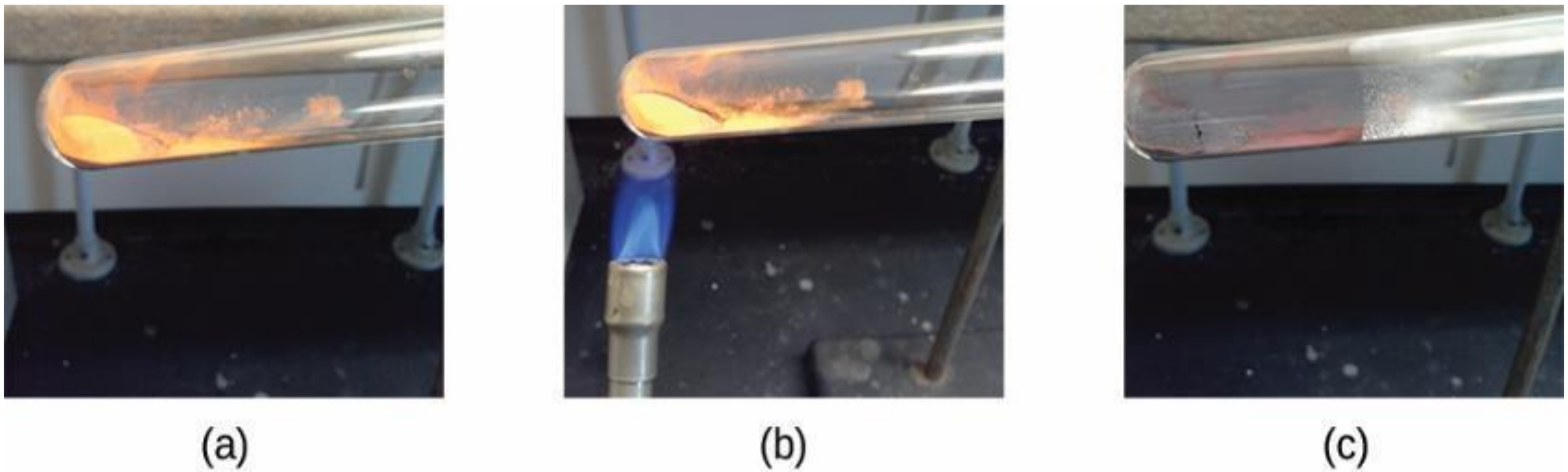
  - There are more than 100 known elements.
  - Ninety of these occur naturally.
  - Two dozen or so have been created in laboratories.

Pure Substances and Mixtures

- Pure substances have constant composition.
  - **Elements:** Pure substance that *cannot* be broken down into simpler substances by chemical changes.
    - Consist of one type of element
    - Examples: Gold (Au), Phosphorus (P), Oxygen (O)

  - **Compounds:** Pure substances that *can* be broken down into simpler substances by chemical changes.
    - Consist of two or more types of elements chemically bonded
    - Examples: H_2O, $C_6H_{12}O_6$, AgCl
    - The properties of compounds are different from the uncombined elements making up the compound.

Figure 1.9



(a) The compound mercury(II) oxide, (b) when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas. (credit: modification of work by Paul Flowers)

Pure Substances and Mixtures

- A mixture is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes.
- Evaporation is an example of a physical change.
- There are two types of mixtures: homogenous mixtures and heterogeneous mixtures.

Two Type of Mixtures

1) A **homogenous mixture** exhibits a uniform composition and appears visually the same throughout.
    - Another name for a homogenous mixture is a **solution**.

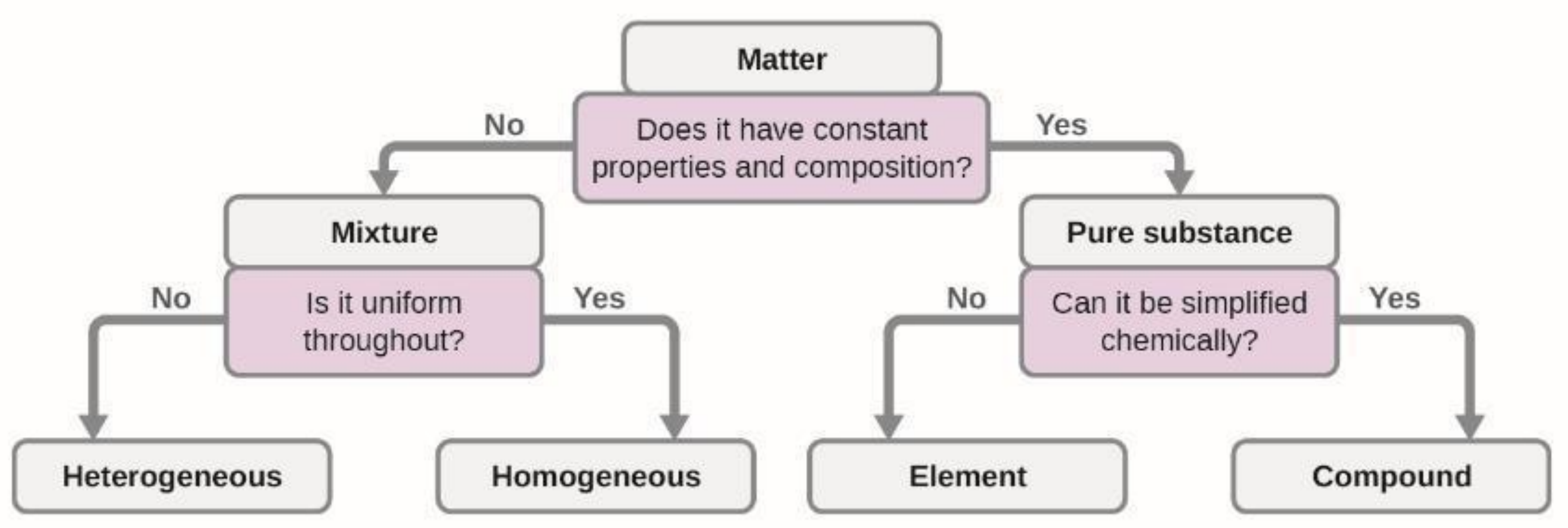
2) A **heterogeneous mixture** has a composition that varies from point to point.

Figure 1.10

(a) (b)

(a) Oil and vinegar salad dressing is a heterogeneous mixture because its composition is not uniform throughout. (b) A commercial sports drink is a homogeneous mixture because its composition is uniform throughout. (credit a “left”: modification of work by John Mayer; credit a “right”: modification of work by Umberto Salvagnin; creditb “left: modification of work by Jeff Bedford)

Figure 1.11

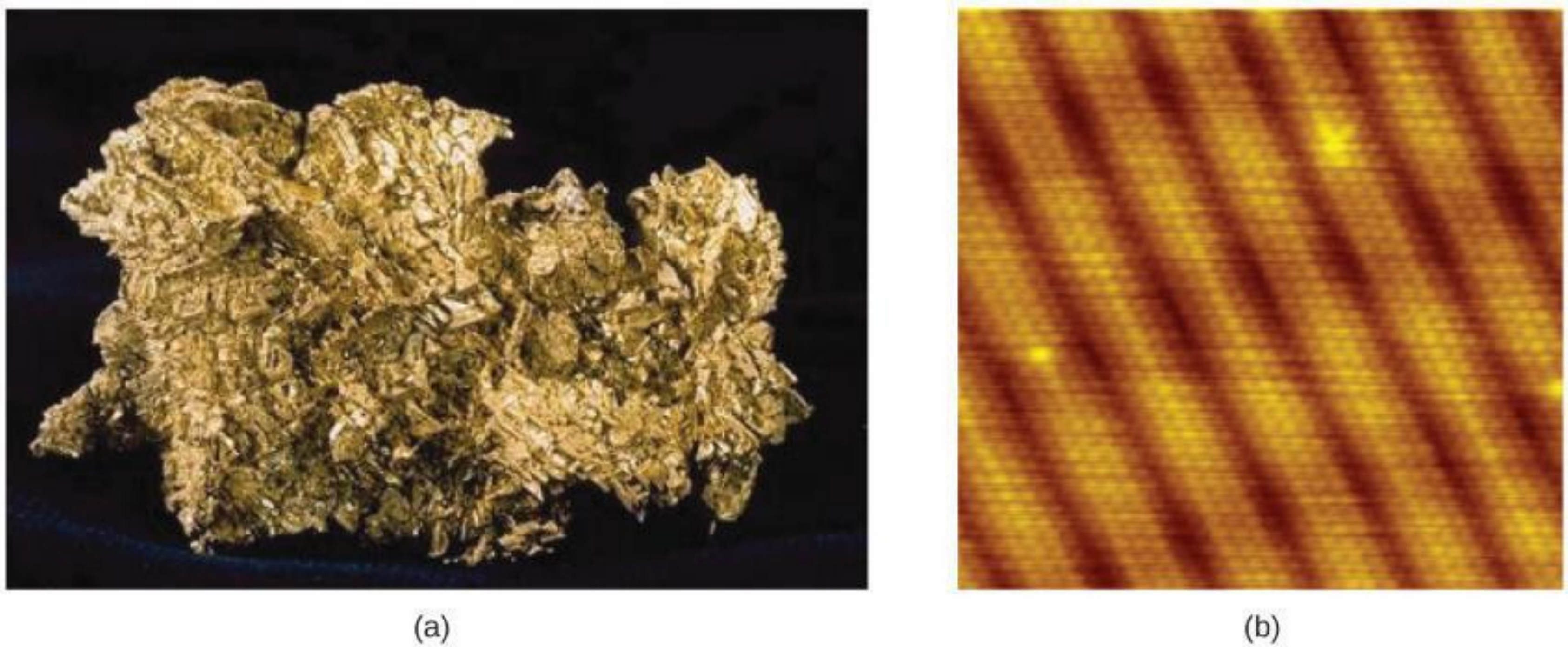


Depending on its properties, a given substance can be classified as a homogeneous mixture, a heterogeneous mixture, a compound, or an element.

- **Atom:** The smallest particle of an element that has the properties of that element and can enter into a chemical combination.
    - Idea first proposed by Greek philosophers, Leucippus and Democritus, in the 5th century BCE.
    - 19th century, John Dalton of England supported this hypothesis with quantitative measurements.

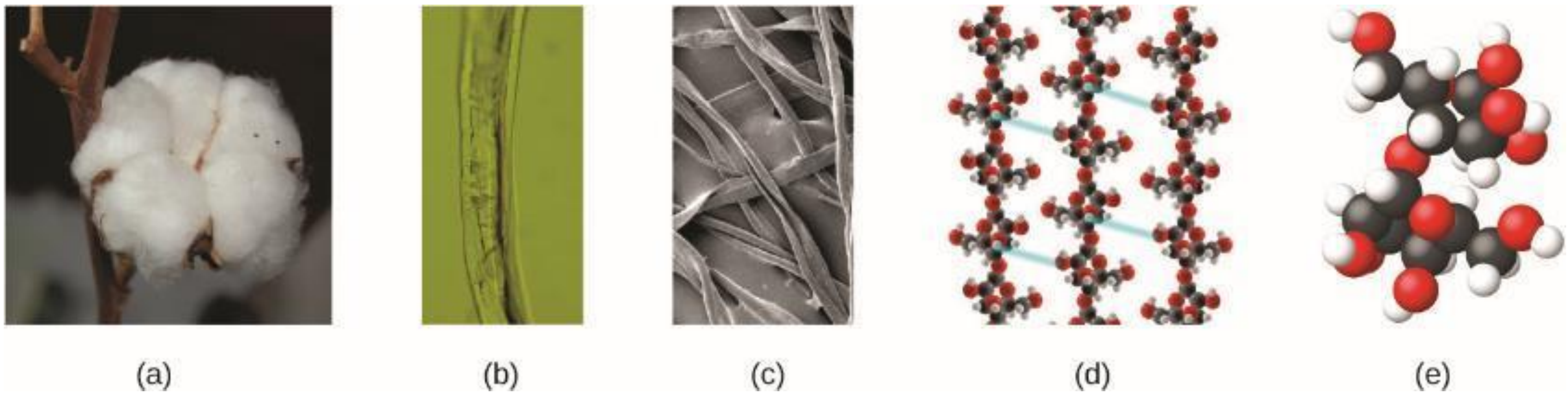
- **Molecule:** Consists of two or more atoms connected by strong forces known as chemical bonds.

Figure 1.12



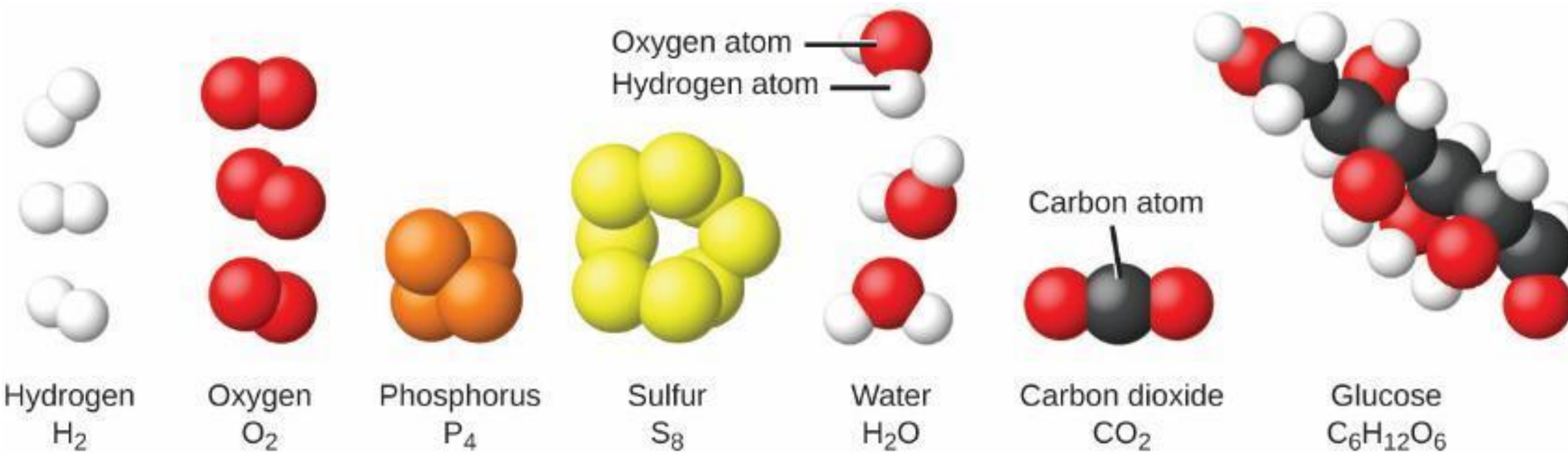
(a) This photograph shows a gold nugget. (b) A scanning-tunneling microscope (STM) can generate views of the surfaces of solids, such as this image of a gold crystal. Each sphere represents one gold atom. (credit a: modification of work by United States Geological Survey; credit b: modification of work by “Erwinrossen”/Wikimedia Commons)

Figure 1.13



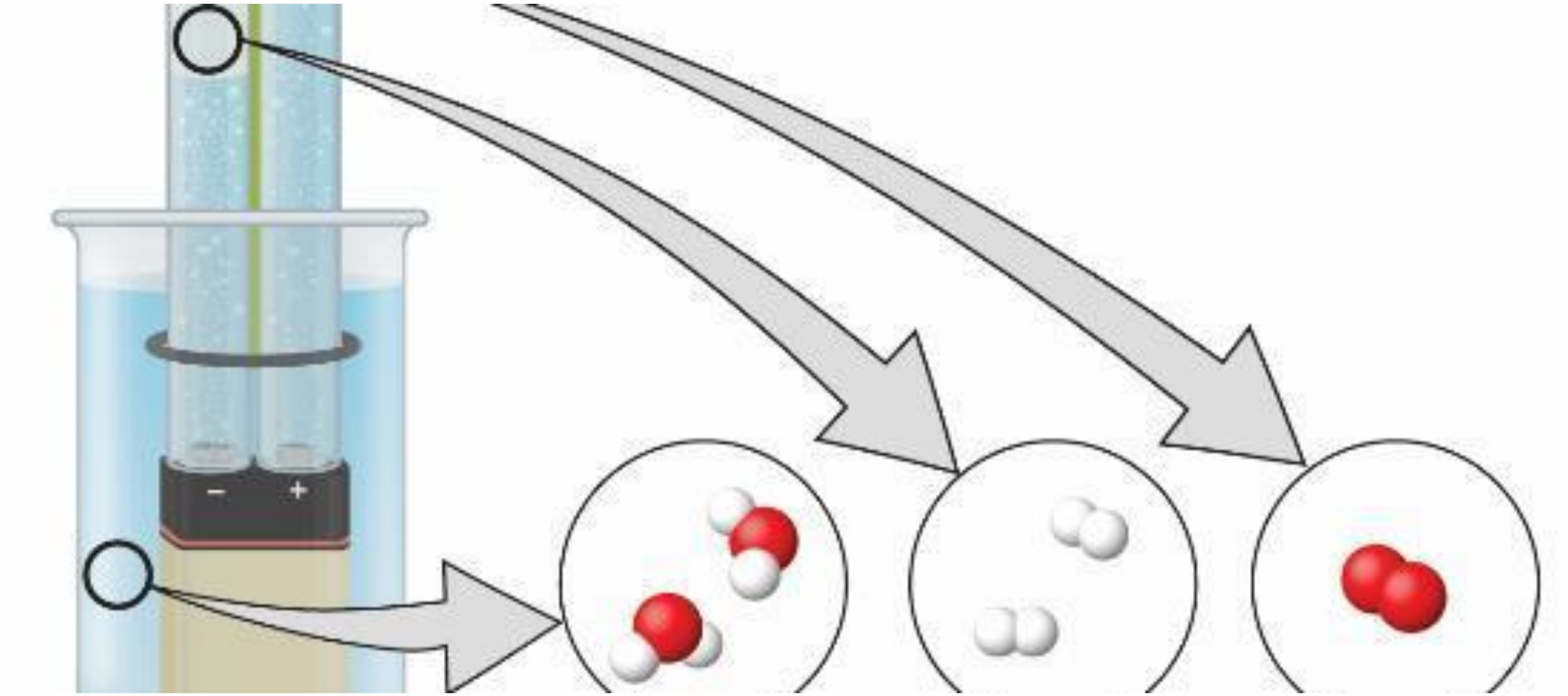
These images provide an increasingly closer view: (a) a cotton boll, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) an image of a cotton fiber obtained with an electron microscope (much higher magnification than with the optical microscope); and (d and e) atomic-level models of the fiber (spheres of different colors represent atoms of different elements). (credit c: modification of work by "Featheredtar"/Wikimedia Commons)

Figure 1.14



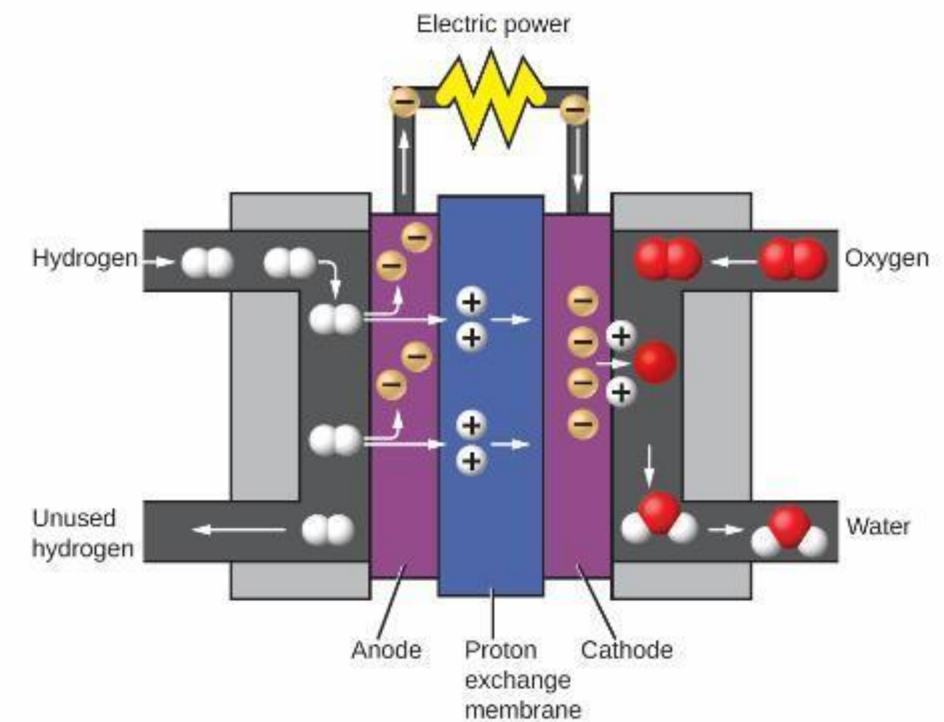
The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

Figure 1.15



The decomposition of water is shown at the macroscopic, microscopic, and symbolic levels. The battery provides an electric current (microscopic) that decomposes water. At the macroscopic level, the liquid separates into the gases hydrogen (on the left) and oxygen (on the right). Symbolically, this change is presented by showing how liquid H_2O separates into H_2 and O_2 gases.

Figure 1.16



A fuel cell generates electrical energy from hydrogen and oxygen via an electrochemical process and produces only water as the waste product.

Figure 1.17

Almost one-third of naturally occurring elements are used to make a cell phone. (credit: modification of work by John Taylor)

Learning Objectives

- 1.3 Physical and Chemical Properties
  - Identify properties of and changes in matter as physical or chemical
  - Identify properties of matter as extensive or intensive

Physical and Chemical Properties

- The characteristics that enable us to distinguish one substance from another are called properties.

- A **physical property** is a characteristic of matter that is not associated with a change in its chemical composition.

  - Examples: density, color, hardness, melting and boiling points, and electrical conductivity

- A *physical change* is a change in the state or properties of matter without any accompanying change in its chemical composition.

Figure 1.18

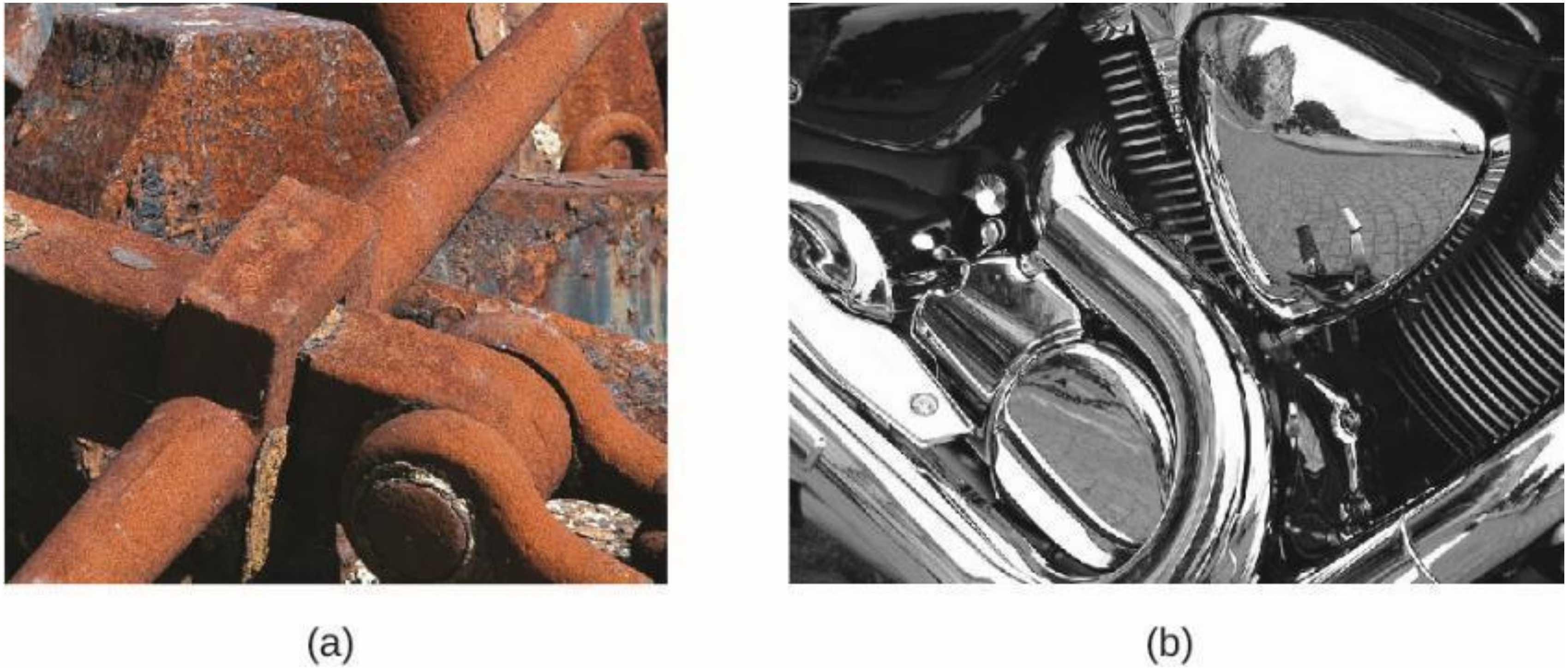
(a) (b)

(a) Butter undergoes a physical change when solid butter is heated and forms liquid melted butter. (b) Steam condensing inside a cooking pot is a physical change, as water vapor is changed into liquid water. (credit a: modification of work by “95jb14”/Wikimedia Commons; credit b: modification of work by “mjneuby”/Flickr)

- The change of one type of matter into another type (or the inability to change) is a **chemical property**.

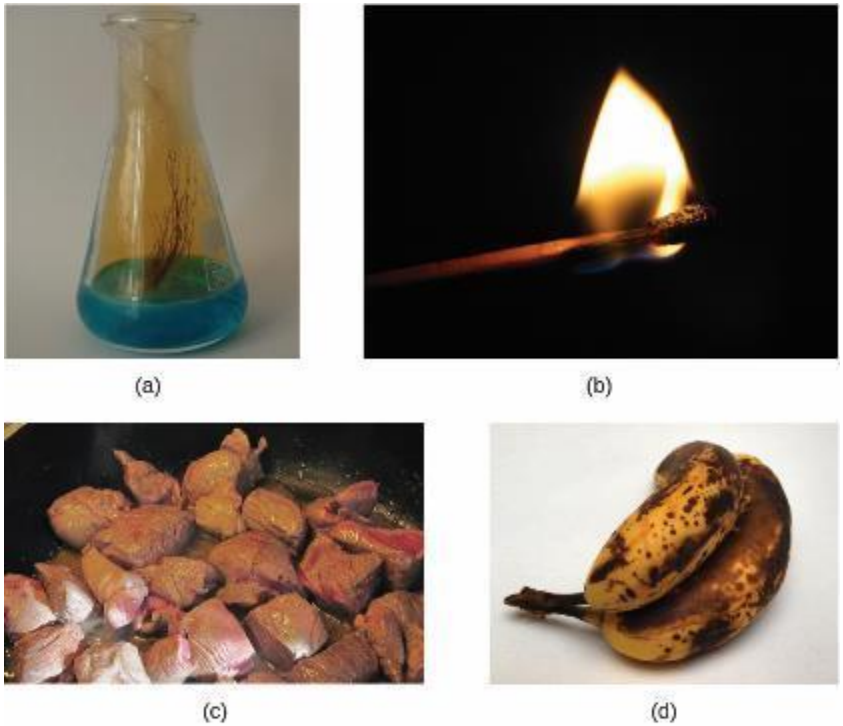
- Examples: flammability, toxicity, acidity, reactivity, and heat of combustion.

Figure 1.19



(a) One of the chemical properties of iron is that it rusts; (b) one of the chemical properties of chromium is that it does not. (credit a: modification of work by Tony Hisgett; credit b: modification of work by “Atoma”/Wikimedia Commons)

Figure 1.20



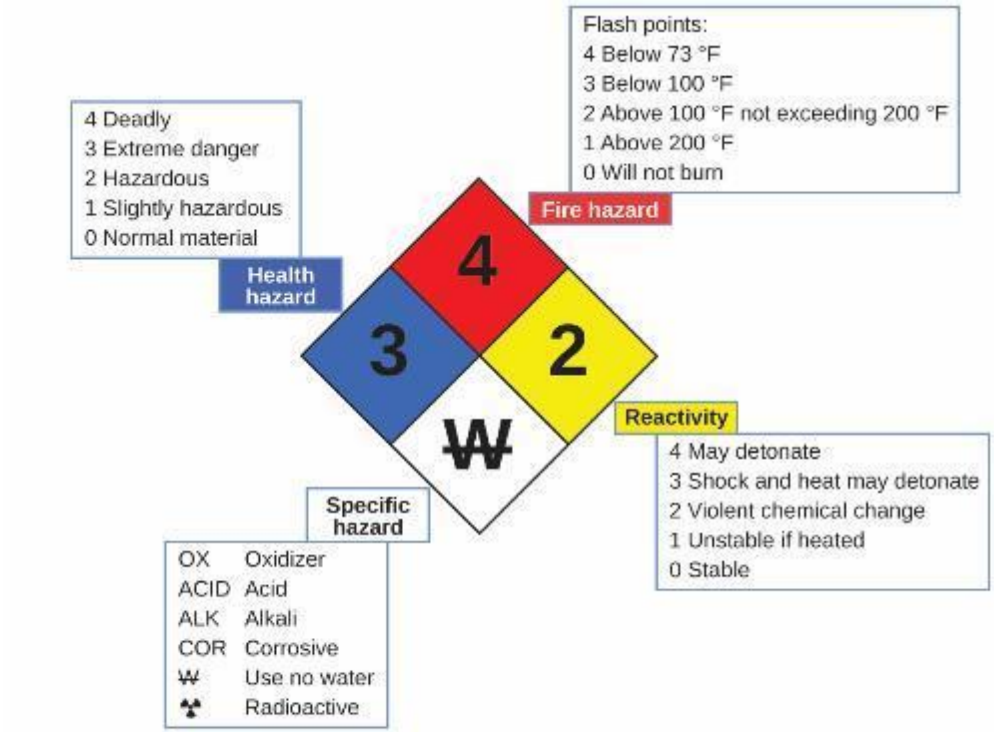
(a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide.

(b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor.

(c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change.

(d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo)

Figure 1.21



The National Fire Protection Agency (NFPA) hazard diamond summarizes the major hazards of a chemical substance.

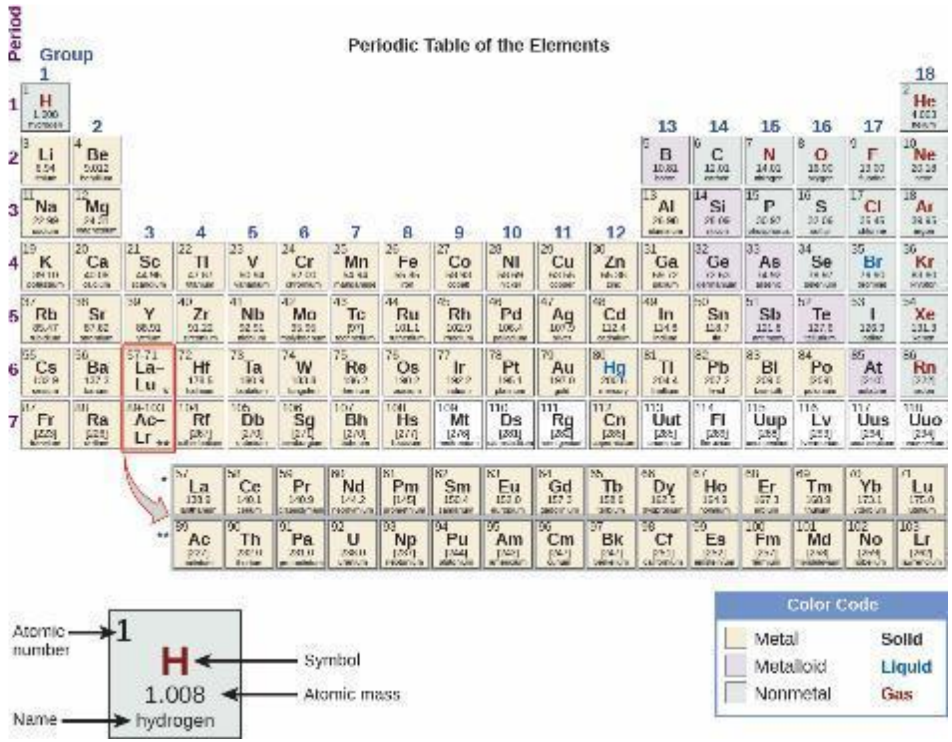
Extensive property

- Depends on the amount of matter present.
- Examples: mass, volume, heat

Intensive property

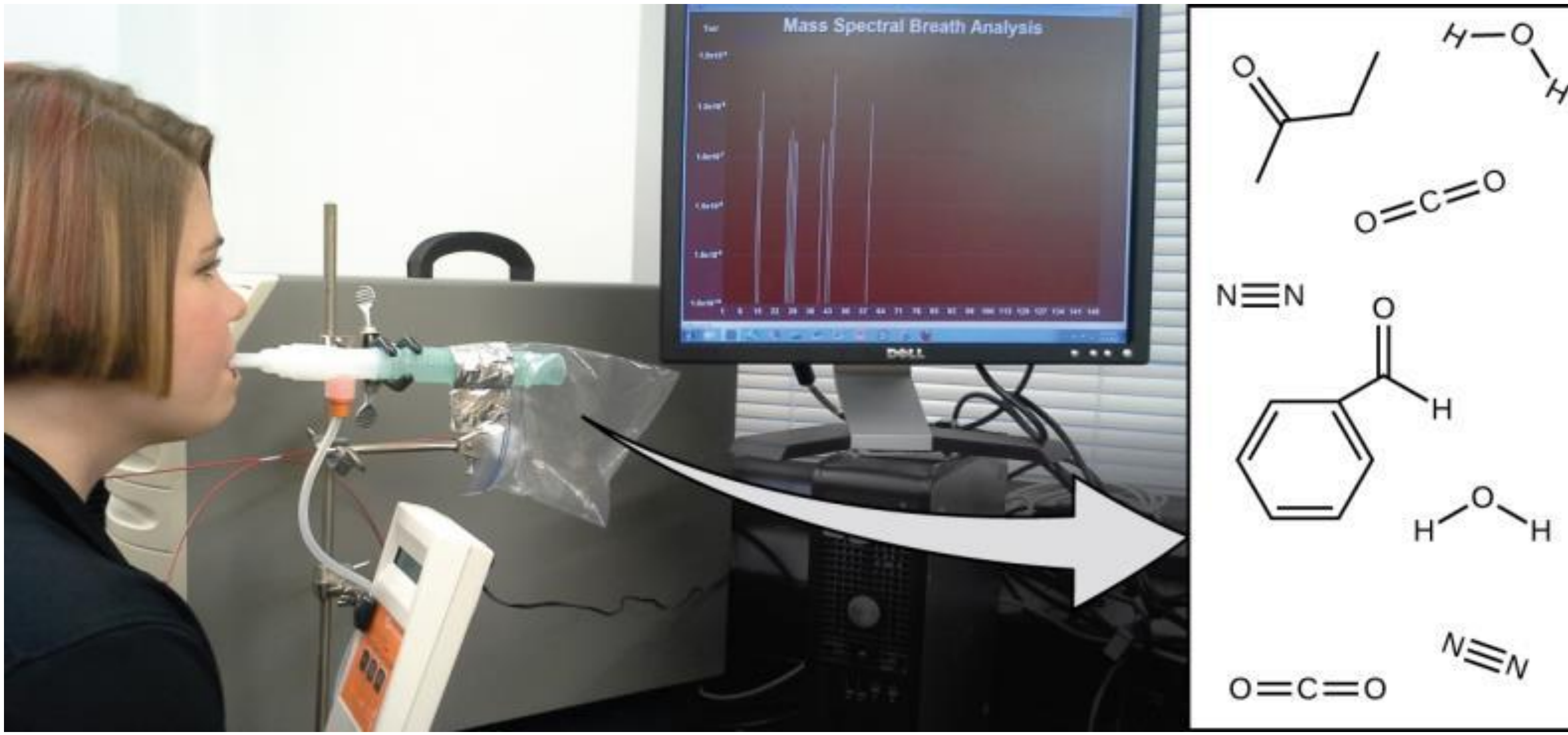
- Does not depend on the amount of matter present.
- Examples: density, temperature

Figure 1.22



The periodic table shows how elements may be grouped according to certain similar properties. Note the background color denotes whether an element is a metal, metalloid, or nonmetal, whereas the element symbol color indicates whether it is a solid, liquid, or gas.

Figure 2.1



Analysis of molecules in an exhaled breath can provide valuable information, leading to early diagnosis of diseases or detection of environmental exposure to harmful substances. (credit: modification of work by Paul Flowers)

Learning Objectives

- 2.1 Early Ideas in Atomic Theory
  - State the postulates of Dalton's atomic theory
  - Use postulates of Dalton's atomic theory to explain the laws of definite and multiple proportions

Early Ideas in Atomic Theory

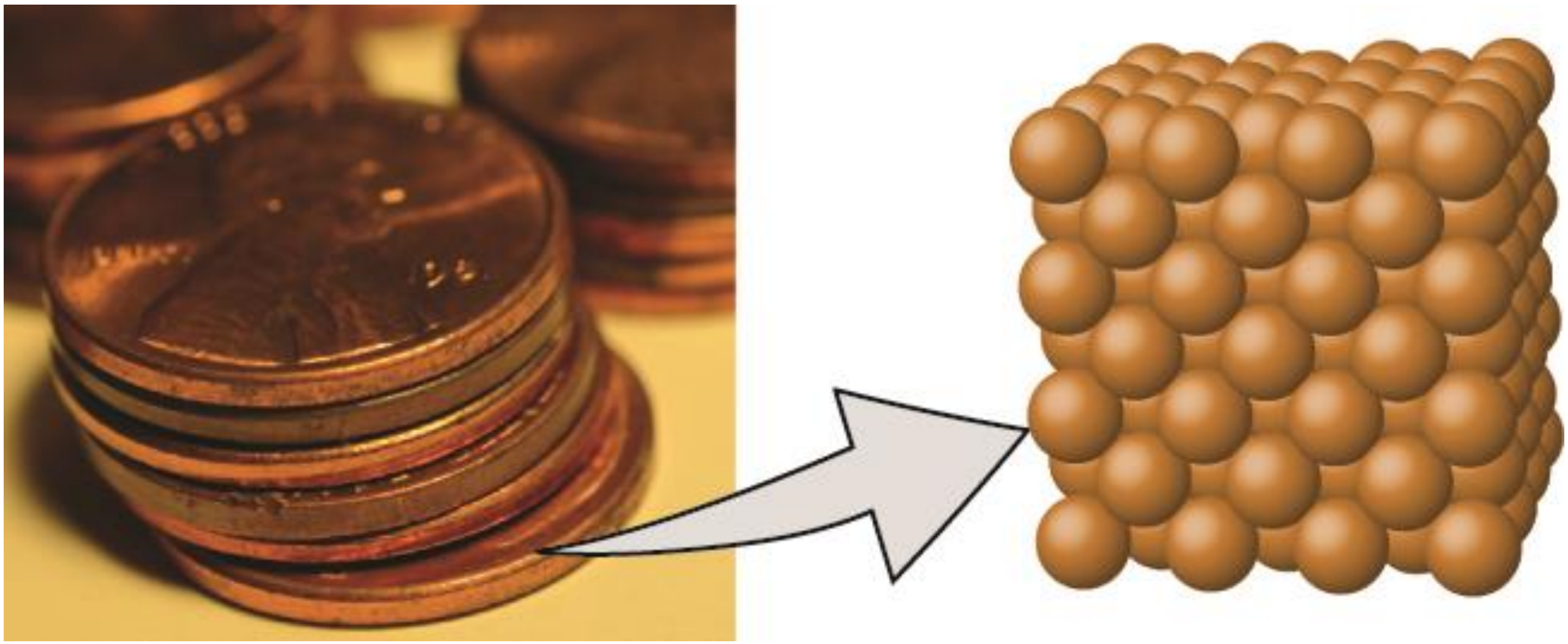
- The concept of atoms was first proposed by the Greek philosophers Leucippus and Democritus in the fifth century BC.
    - *atomos*, a term derived from the Greek word for "indivisible"
- Later, Aristotle and others believed that matter consisted of various combinations of the four "elements"—fire, earth, air, and water.
- In 1807, English schoolteacher John Dalton proposed his atomic theory.

Dalton's Atomic Theory

Dalton's atomic theory can be summarized in five postulates.

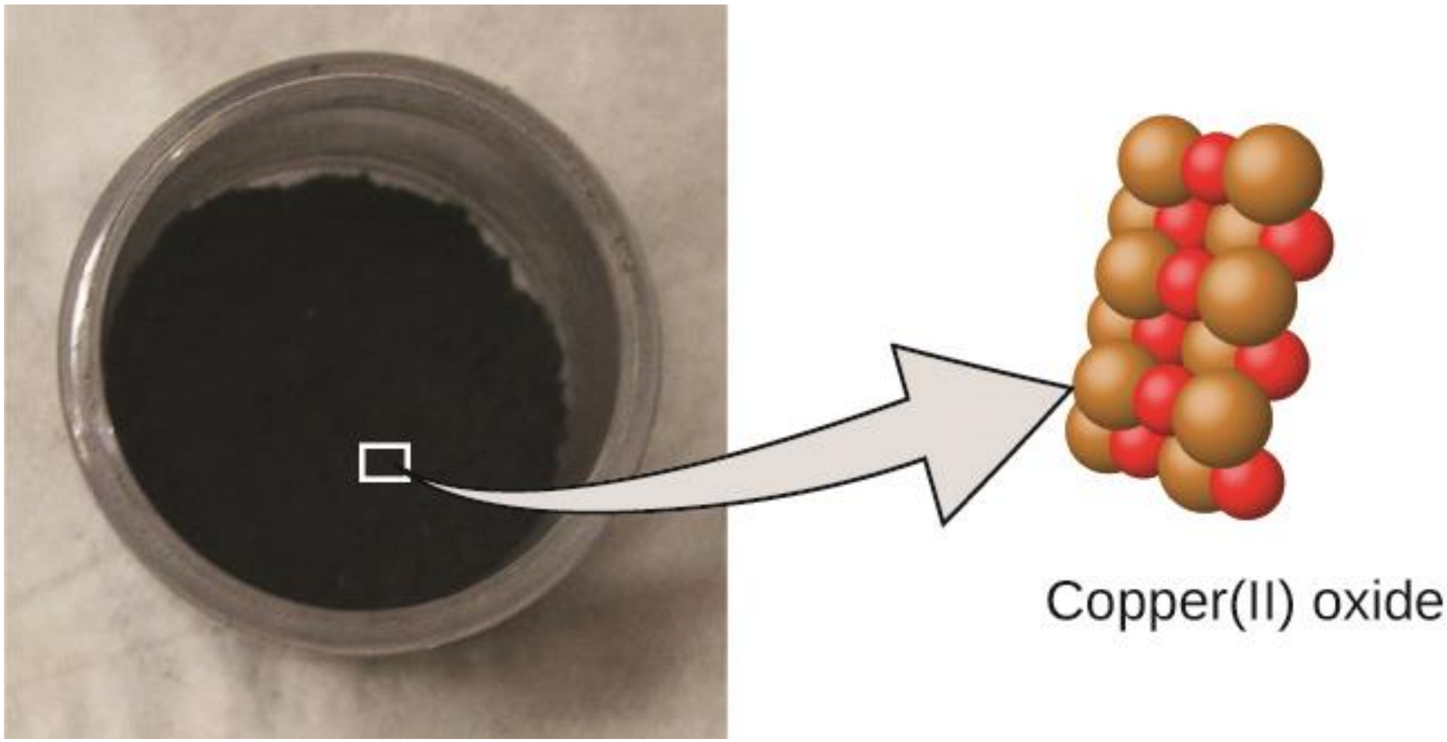
1) Matter is composed of exceedingly small particles called atoms. An atom is the smallest unit of an element that can participate in a chemical change.
2) An element consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element.
3) Atoms of one element differ in properties from atoms of all other elements.
4) A compound consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the number of atoms of each of its elements are always present in the same ratio.
5) Atoms are neither created nor destroyed during a chemical change, but instead rearrange to yield a different type(s) of matter.

Figure 2.2



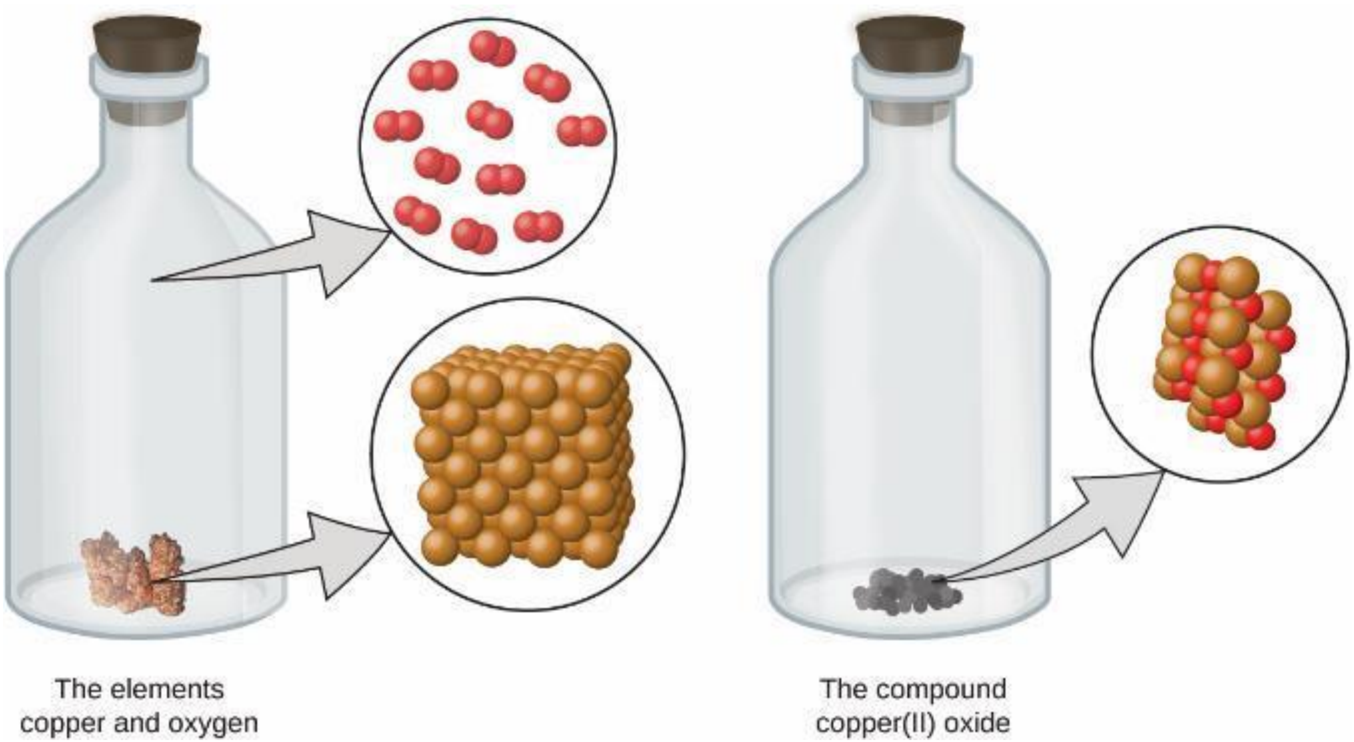
A pre-1982 copper penny (left) contains approximately 3×10^{22} copper atoms (several dozen are represented as brown spheres at the right), each of which has the same chemical properties. (credit: modification of work by “slgckgc”/Flickr)

Figure 2.3



Copper(II) oxide, a powdery, black compound, results from the combination of two types of atoms—copper (brown spheres) and oxygen (red spheres)—in a 1:1 ratio. (credit: modification of work by "Chemicalinterest"/Wikimedia Commons)

Figure 2.4



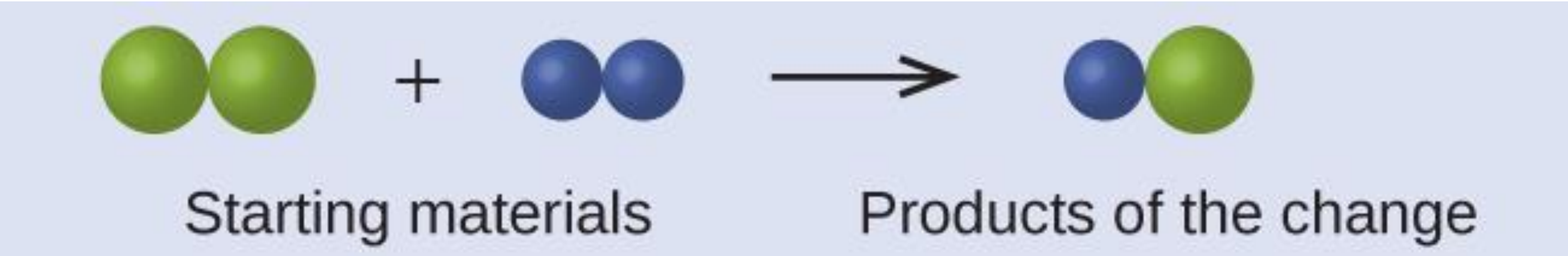
When the elements copper (a shiny, red-brown solid, shown here as brown spheres) and oxygen (a clear and colorless gas, shown here as red spheres) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid). (credit copper: modification of work by http://imagesof-elements.com/copper.php)

Law of Conservation of Matter

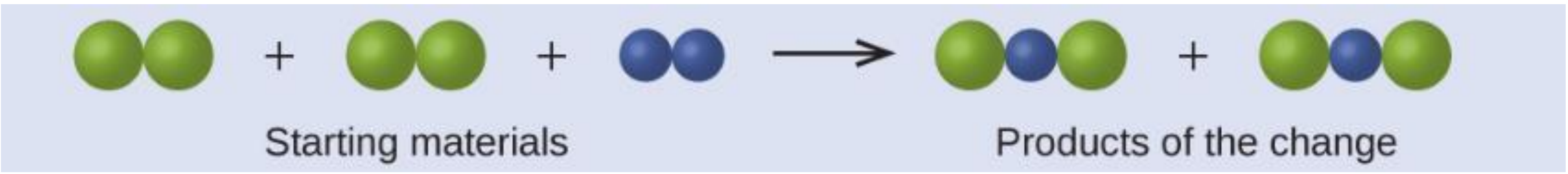
- Dalton’s atomic theory provides a microscopic explanation of the many macroscopic properties of matter that you’ve learned about.

- If atoms are neither created nor destroyed during a chemical change, then the total mass of matter present when matter changes from one type to another will remain constant (the law of conservation of matter).

Example 2.1



Example 2.1



Law of Definite Proportions

- **Law of definite proportions** or the **law of constant composition**: All samples of a pure compound contain the same elements in the same proportion by mass.
    - Illustrated by experiments performed by French chemist Joseph Proust.

Table 2.1 Constant Composition of Isooctane

Sample	Carbon	Hydrogen	Mass Ratio
A	14.82 g	2.78 g	$\frac{14.82 \text{ g carbon}}{2.78 \text{ g hydrogen}} = \frac{5.33 \text{ g carbon}}{1.00 \text{ g hydrogen}}$
B	22.33 g	4.19 g	$\frac{2.33 \text{ g carbon}}{4.19 \text{ g hydrogen}} = \frac{5.33 \text{ g carbon}}{1.00 \text{ g hydrogen}}$
C	19.40 g	3.64 g	$\frac{19.40 \text{ g carbon}}{3.63 \text{ g hydrogen}} = \frac{5.33 \text{ g carbon}}{1.00 \text{ g hydrogen}}$

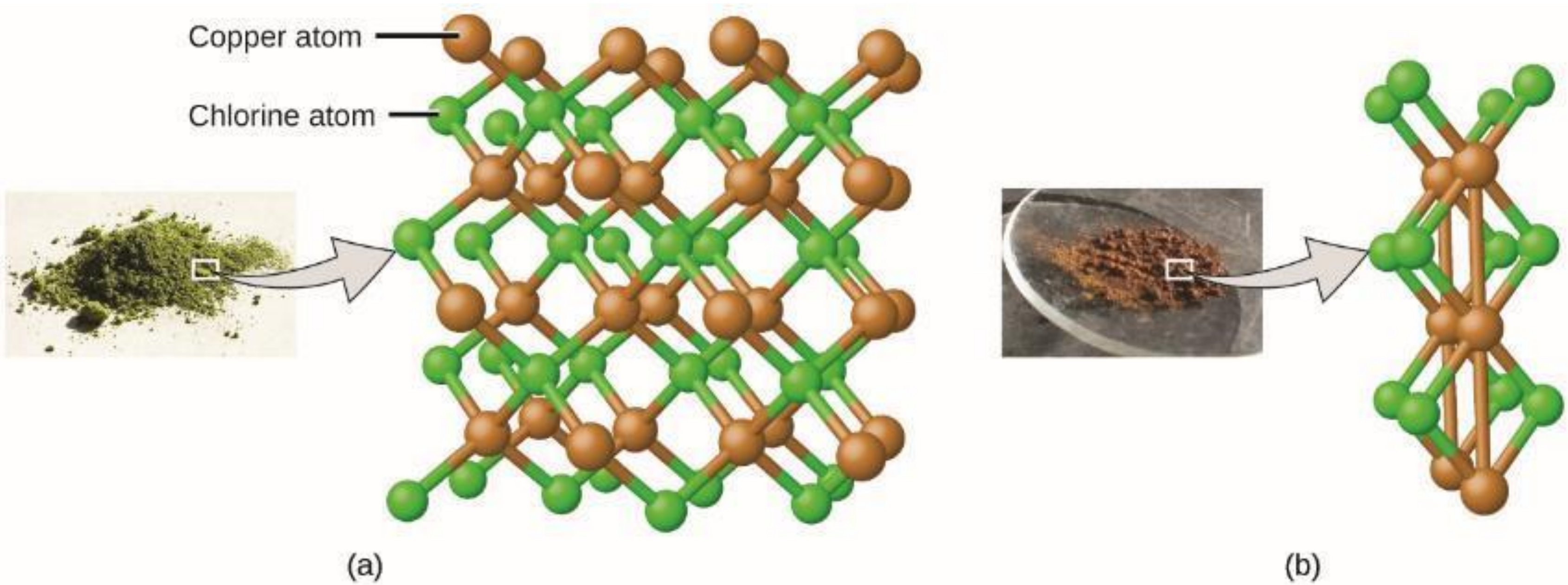
Law of Multiple Proportions

- The law of multiple proportions states that when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small, whole numbers.

- Example: compounds containing chlorine and copper
  - A green solid contains 0.558 g Cl to 1 g Cu.
  - A brown solid contains 1.116 g Cl to 1 g Cu.

$$\frac{\dfrac{1.116\ g\ Cl}{1\ g\ Cu}}{\dfrac{0.558\ g\ Cl}{1\ g\ Cu}} = \frac{2}{1}$$

Figure 2.5



Compared to the copper chlorine compound in (a), where copper is represented by brown spheres and chlorine by green spheres, the copper chlorine compound in (b) has twice as many chlorine atoms per copper atom. (credit a: modification of work by "Benjah-bmm27"/Wikimedia Commons; credit b: modification of work by "Walkerma"/Wikimedia Commons)

Learning Objectives

- 2.2 Evolution of Atomic Theory
  - Outline milestones in the development of modern atomic theory
  - Summarize and interpret the results of the experiments of Thomson, Millikan, and Rutherford
  - Describe the three subatomic particles that compose atoms
  - Define isotopes and give examples for several elements

Evolution of Atomic Theory

- What are atoms composed of?
- Is there something smaller than an atom?
- Here, we will discuss some of these key developments.

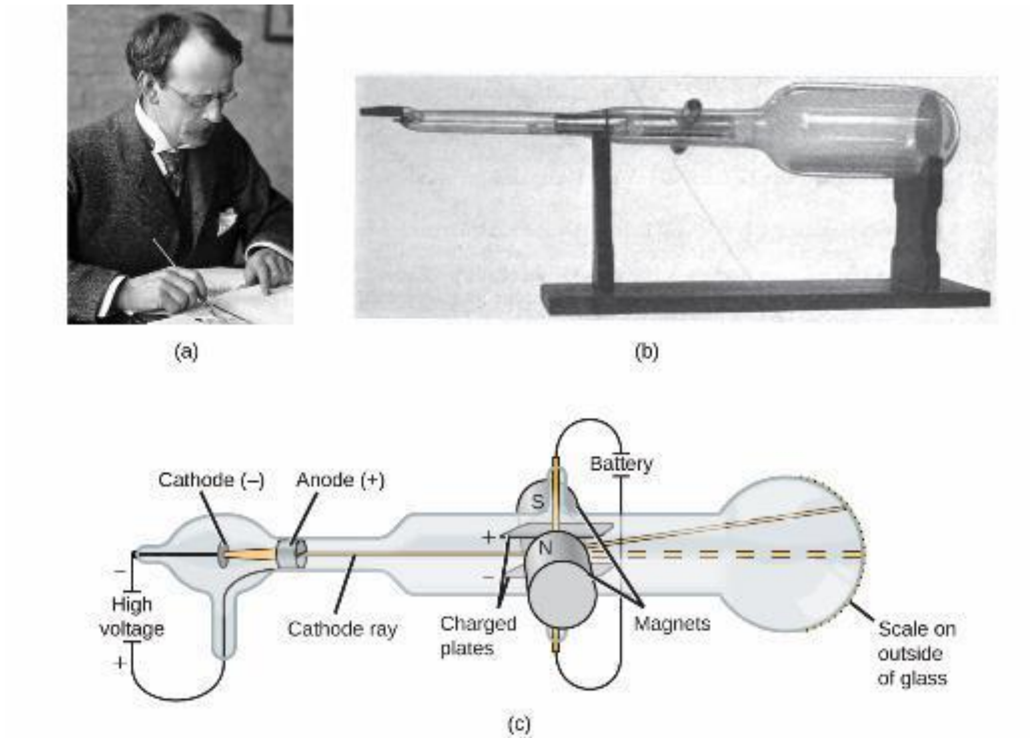
Discovery of the Electron: J.J. Thomson

- J.J. Thomson experimented with cathode ray tubes.
- Cathode ray tube:
  - A sealed glass tube from which almost all the air had been removed
  - Contained two metal electrodes
- When a high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them.
- Regardless of the metals used, this beam was always deflected toward the positive charge and away from the negative charge.
- Thompson was able to calculate the charge-to-mass ratio of the cathode ray particles.

Thompson's Results

- The cathode ray particles were much lighter than atoms.
- These particles are negatively charged.
- These particles are indistinguishable, regardless of the source material.
- This cathode ray particle is what we now call an electron—a negatively charged, subatomic particle with a mass more than one thousand times less than that of an atom.

Figure 2.6

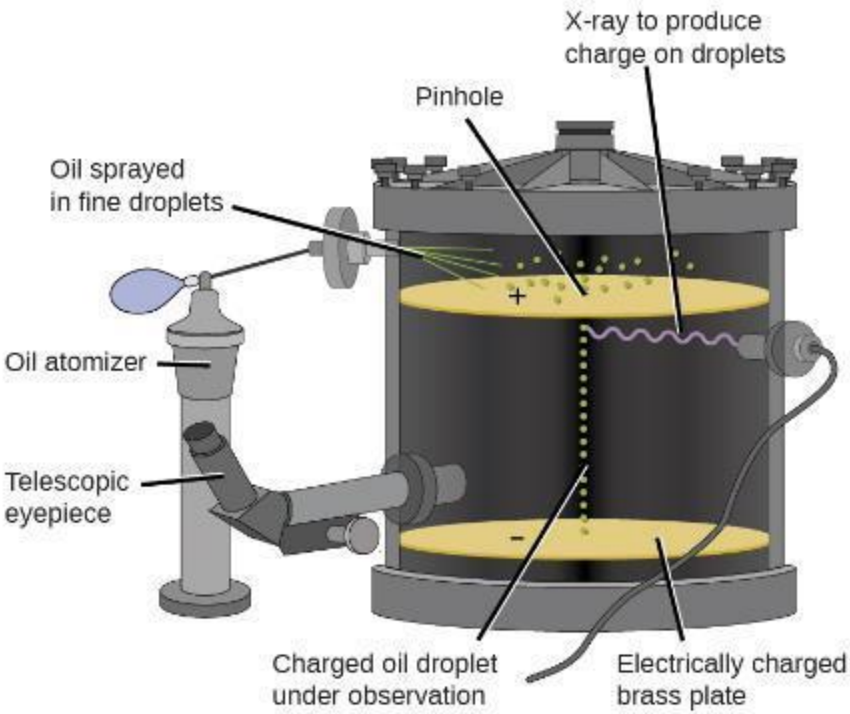


(a) J. J. Thomson produced a visible beam in a cathode ray tube. (b) This is an early cathode ray tube, invented in 1897 by Ferdinand Braun. (c) In the cathode ray, the beam (shown in yellow) comes from the cathode and is accelerated past the anode toward a fluorescent scale at the end of the tube. Simultaneous deflections by applied electric and magnetic fields permitted Thomson to calculate the mass-to-charge ratio of the particles composing the cathode ray. (credit a: modification of work by Nobel Foundation; credit b: modification of work by Eugen Nesper; credit c: modification of work by "Kurzon"/Wikimedia Commons)

Discovery of the Electron: Robert A Millikan

- Robert A. Millikan’s Oil Drop Experiment (1909)
- Millikan created microscopic oil droplets, which were electrically charged.
- These drops could be slowed or reversed by an electric field.
- Millikan was able to determine the charge on individual drops.

Figure 2.7



Oil drop	Charge in coulombs (C)
A	4.8×10^{19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{19} C

Millikan's experiment measured the charge of individual oil drops. The tabulated data are examples of a few possible values.

Millikan's Results

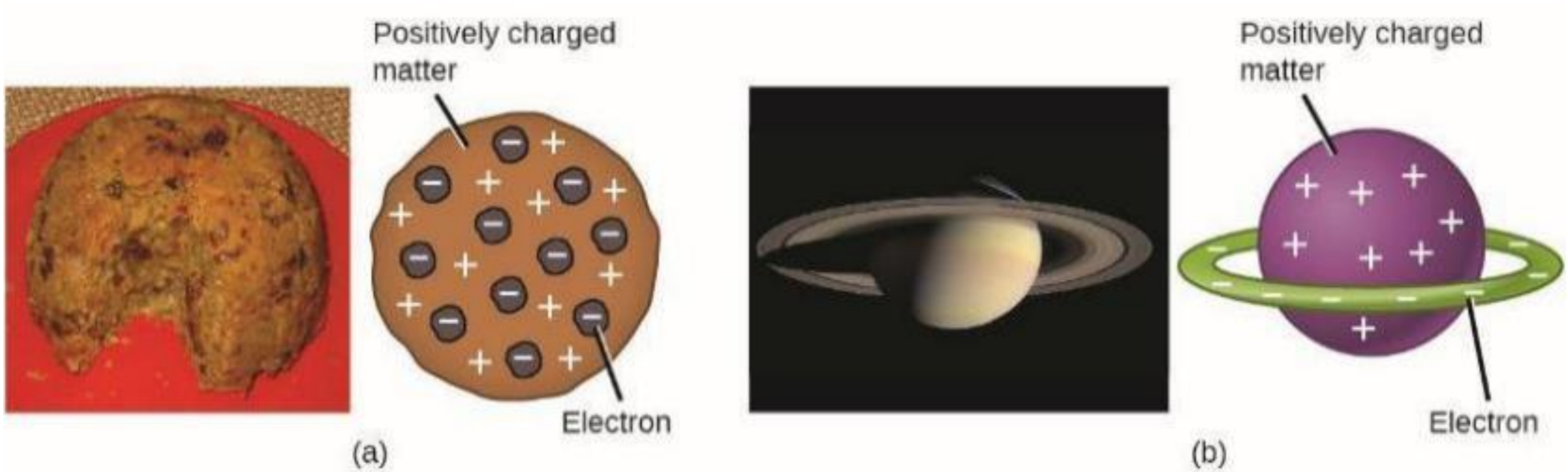
- The charge of an oil drop was always a multiple of a specific charge, 1.6×10^{-19} C.

- Millikan concluded that 1.6×10^{-19} C was the charge of a single electron.

- Thompson already showed the charge to mass ratio of an electron to be 1.759×10^{11} C/kg.

$$Mass\ of\ electron = 1.602 \times 10^{-19}\,C\left(\frac{1\,kg}{1.759 \times 10^{11}\,C}\right) = 9.107 \times 10^{-31}\,kg$$

Figure 2.8

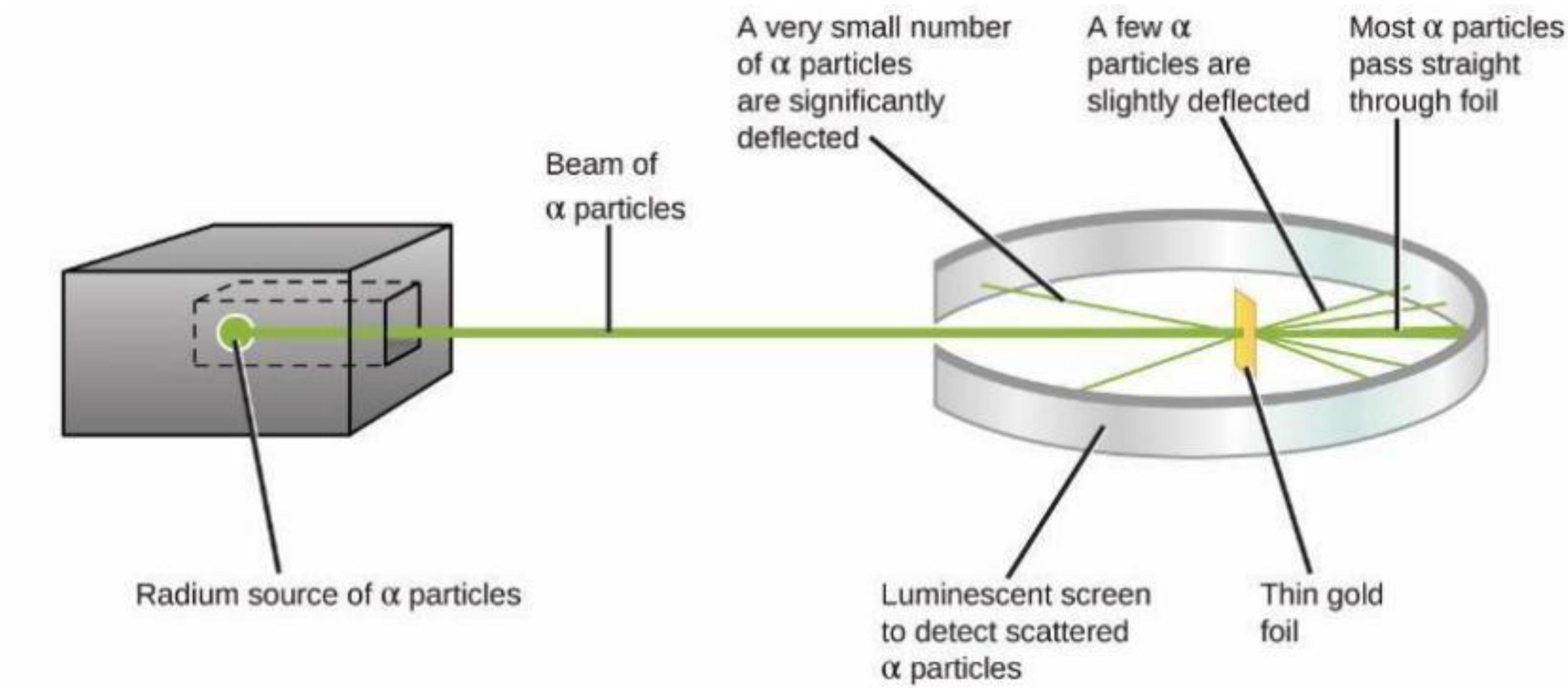


(a) Thomson suggested that atoms resembled plum pudding, an English dessert consisting of moist cake with embedded raisins ("plums"). (b) Nagaoka proposed that atoms resembled the planet Saturn, with a ring of electrons surrounding a positive "planet." (credit a: modification of work by "Man vyi"/Wikimedia Commons; credit b: modification of work by "NASA"/Wikimedia Commons)

Discovery of the Nucleus: Ernest Rutherford

- Ernest Rutherford's Gold Foil Scattering Experiment
- Aimed a beam of alpha particles (a particles) at a very thin piece of gold foil.
- α particles are positively charged.
- The scattering of these α particles was examined using a luminescent screen that would glow briefly when hit.

Figure 2.9

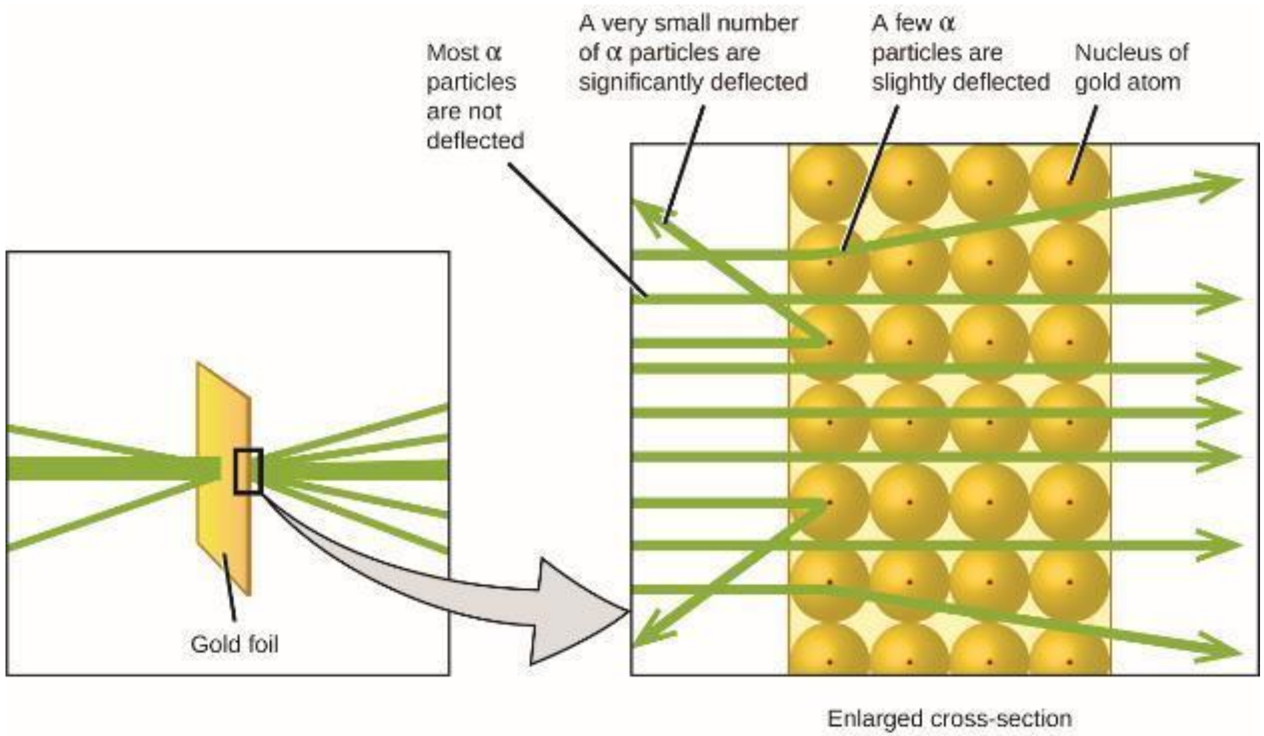


Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly deflected.

Rutherford’s Results

- The volume occupied by an atom must consist of a large amount of empty space.
- A small, relatively heavy, positively charged body, the **nucleus**, must be at the center of each atom.
- The nucleus contains most of the atom’s mass.
- Negatively charged electrons surround the nucleus.
- The **proton**, a positively charged, subatomic particle is located in the nucleus.

Figure 2.10



The α particles are deflected only when they collide with or pass close to the much heavier, positively charged gold nucleus. Because the nucleus is very small compared to the size of an atom, very few α particles are deflected. Most pass through the relatively large region occupied by electrons, which are too light to deflect the rapidly moving particles.

Other Important Discoveries of the 20^{th} Century

- **Isotopes:** Atoms of the same element that differ in mass
  - Frederick Soddy of England. Noble Prize in 1921.

- **Neutrons:** Uncharged, subatomic particles with a mass approximately the same as that of protons
  - Discovered by James Chadwick in 1932.
  - Neutrons are also found in the nucleus.

Learning Objectives

- 2.3 Atomic Structure and Symbolism
  - Write and interpret symbols that depict the atomic number, mass number, and charge of an atom or ion
  - Define the atomic mass unit and average atomic mass
  - Calculate average atomic mass and isotopic abundance

Atomic Structure and Symbolism

- The nucleus contains the majority of an atom’s mass.
- Protons and neutrons are much heavier than electrons.
- Electrons occupy almost all of an atom’s volume.
- Diameter of an atom ~ 10^{-10} m
- Diameter of a nucleus is 100,000 times smaller ~ 10^{-15} m

Figure 2.11

If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by “babyknight”/Wikimedia Commons; credit right: modification of work by Paxson Woelber)

Units

- Atoms and subatomic particles are very small.
  - Example: A carbon atom weighs less than 2×10^{-23} g.

- Electrons have a charge of less than 2×10^{-19} C.

- Small units are needed.
  - Atomic mass unit (amu).
    - 1 amu = 1.6605 x 10-24 g.
    - Mass of a carbon-12 atom = 12 amu

- Fundamental unit of charge (e).

  $e = 1.602 \times 10^{-19}$ C

Properties of Subatomic Particles

- Proton
  - Mass = 1.0073 amu
  - Charge = +1
- Neutron
  - Mass = 1.0087 amu (slightly heavier than a proton)
  - Charge = 0
- Electron
  - Mass = 0.00055 amu
  - Charge = –1

Atomic Number (Z)

- The number of protons in the nucleus of an atom is its **atomic number (Z)**.

- This is the defining trait of an element: Its value determines the identity of the atom.

- For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have.

Neutral Atoms

- A neutral atom must contain the same number of positive and negative charges.
- The number of protons equals the number of electrons.
- Therefore, the atomic number also indicates the number of electrons in a neutral atom.

Mass Number (A)

- The total number of protons and neutrons in an atom is called its **mass number (A)**.

- The number of neutrons is therefore the difference between the mass number and the atomic number.

atomic number (Z) = number of protons

mass number (A) = number of protons + number of neutrons

A – Z = number of neutrons

Ions

- When the number of protons and electrons are NOT equal, the atom is electrically charged and called an ion.

Charge of an atom = number of protons – number of electrons

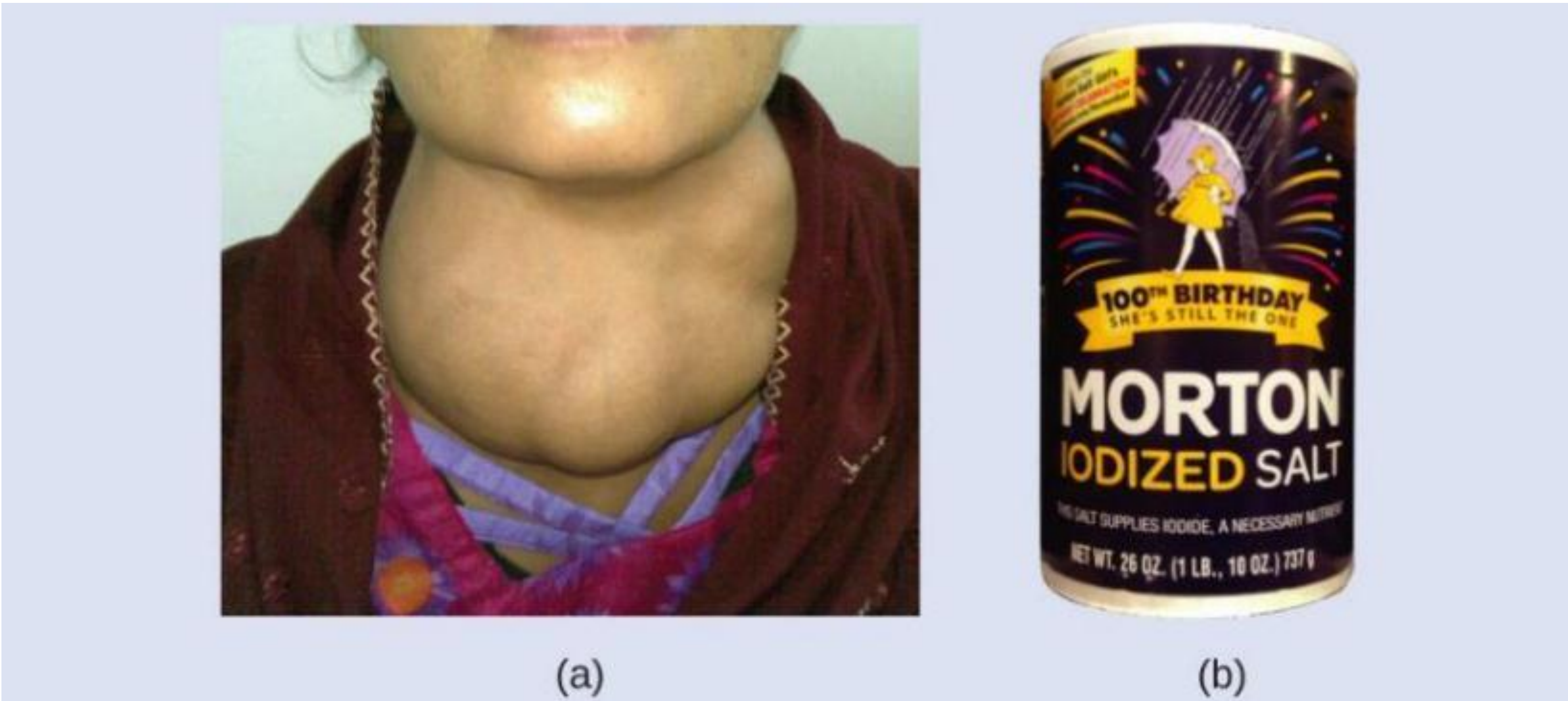
- Atoms (and molecules) acquire charge by losing or gaining electrons.

Cations and Anions

- An atom that gains one or more electrons will exhibit a *negative charge* and is called an **anion**.
  - Example: A neutral oxygen atom (Z = 8) has eight electrons, and if it gains two electrons it will become an anion with a 2− charge (8 − 10 = 2−).

- An atom that loses one or more electrons will exhibit a *positive charge* and is called an **cation**.
  - Example: A neutral sodium atom (Z = 11) has 11 electrons. If this atom loses one electron, it will become a cation with a 1+ charge (11 − 10 = 1+).

Figure 2.12



(a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification of work by "Almazi"/Wikimedia Commons; credit b: modification of work by Mike Mozart)

Chemical Symbols

- A chemical symbol is an abbreviation that we use to indicate an element or an atom of an element.
  - Example: The symbol for mercury is Hg.
- Some symbols are derived from the common name of the element; others are abbreviations of the name in another language.
- Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112.
- Only the first letter of a chemical symbol is capitalized.

Figure 2.13

The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

Table 2.3 Some Common Elements and Their Symbols

Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from *ferrum*)
bromine	Br	lead	Pb (from *plumbum*)
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from *hydrargyrum*)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from *kalium*)
copper	Cu (from *cuprum*)	silicon	Si
fluorine	F	silver	Ag (from *argentum*)
gold	Au (from *aurum*)	sodium	Na (from *natrium*)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from *stannum*)
iodine	I	zinc	Zn

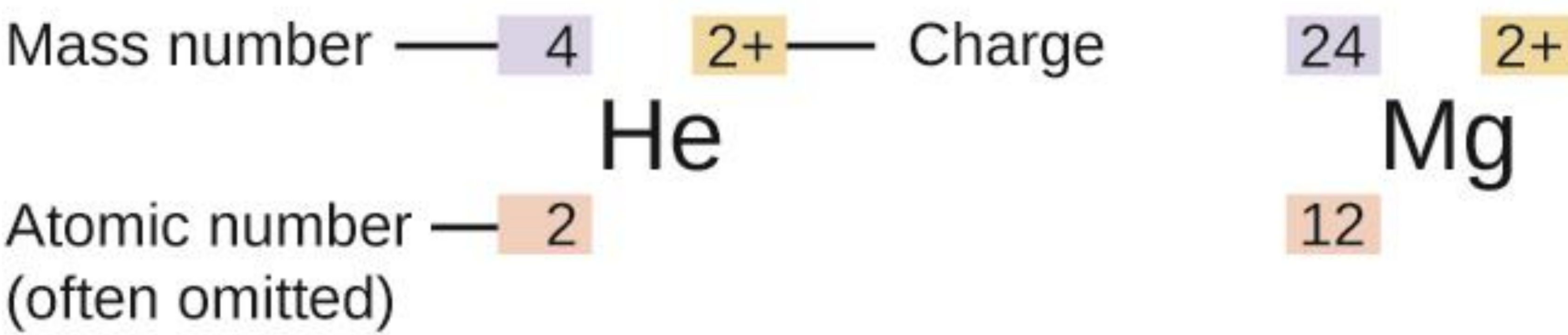
Isotopes

- The symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol.

- The atomic number is sometimes written as a subscript to the left of the element symbol.

- For example, magnesium exists as a mixture of three isotopes.
  - ^{24}Mg, ^{25}Mg, and ^{26}Mg
  - All isotopes have 12 protons, but the number of neutrons are different.

Figure 2.14



The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

Table 2.4 (Partial) Nuclear Compositions of Atoms of the Very Light Elements

- Hydrogen exists as a mixture of three isotopes

Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
hydrogen	$^{1}_{1}H$ (protium)	1	1	0	1.0078	99.989
	$^{2}_{1}H$ (deuterium)	1	1	1	2.0141	0.0115
	$^{3}_{1}H$ (tritium)	1	1	2	3.01605	trace

Atomic Mass

- Each proton and each neutron has a mass of ~ 1 amu.
- Each electron weighs far less.
- Therefore the **atomic mass** of a single atom in amu is *approximately* equal to its mass number.
- However, most elements exist naturally as a mixture of two or more isotopes.
- The periodic table lists the weighted, average mass of all the isotopes present in a naturally occurring sample of that element.

Atomic Mass (continued)

$$average\ mass = \sum_i (fractional\ abundance \times isotopic\ mass)_i$$

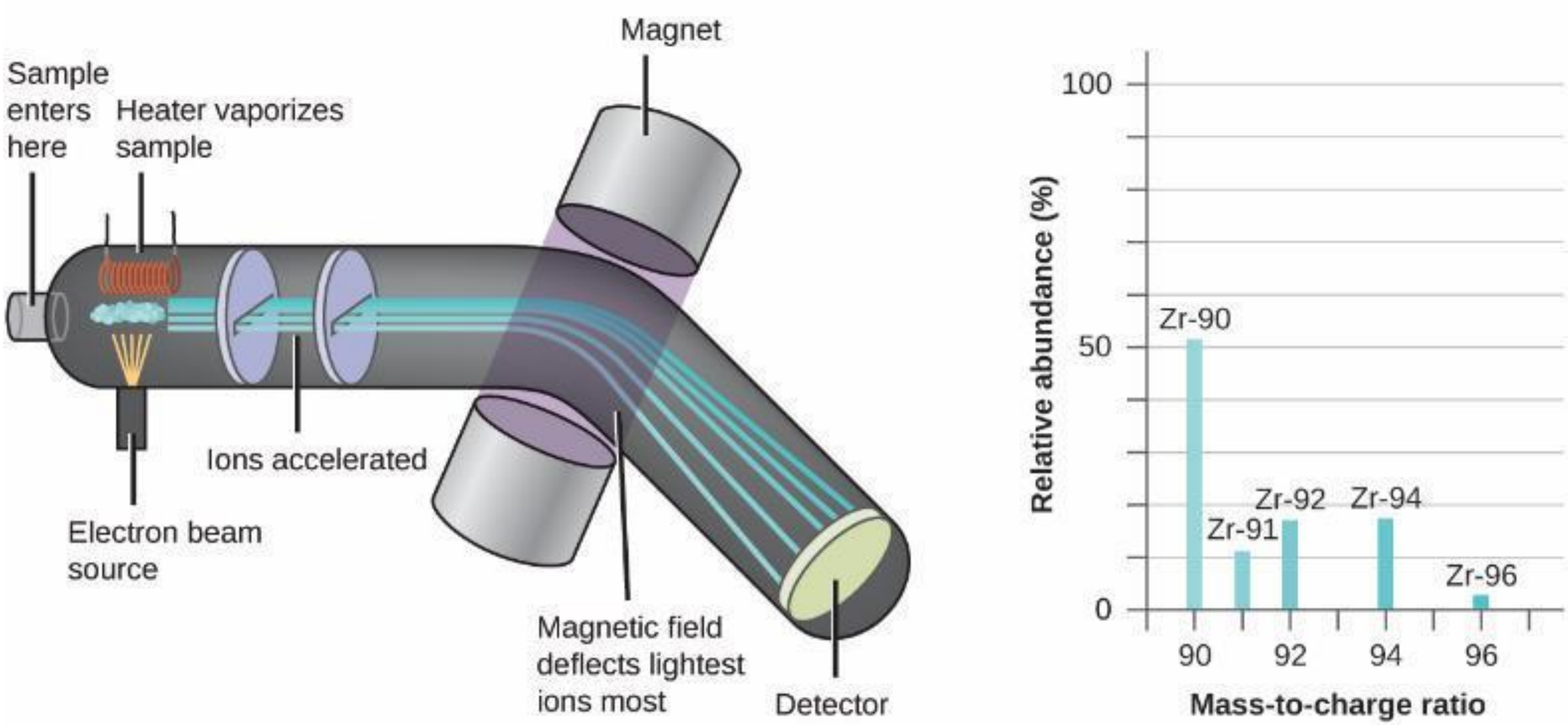
- For example, the element boron is composed of two isotopes:
  - 19.9% ^{10}B with a mass of 10.0129 amu
  - 80.1% ^{11}B with a mass of 11.0093 amu

$$boron\ average\ mass$$
$$= (0.199 \times 10.0129\ amu) + (0.801 \times 11.0093\ amu)$$
$$= 10.81\ amu$$

Mass Spectrometry (MS)

- The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer.
- The sample is vaporized and exposed to a high-energy electron beam that causes the sample's atoms (or molecules) to become electrically charged, typically by losing one or more electrons.
- These cations are then separated by their mass and charge.

Figure 2.15



Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.

Learning Objectives

- 2.4 Chemical Formulas
  - Symbolize the composition of molecules using molecular formulas and empirical formulas
  - Represent the bonding arrangement of atoms within molecules using structural formulas

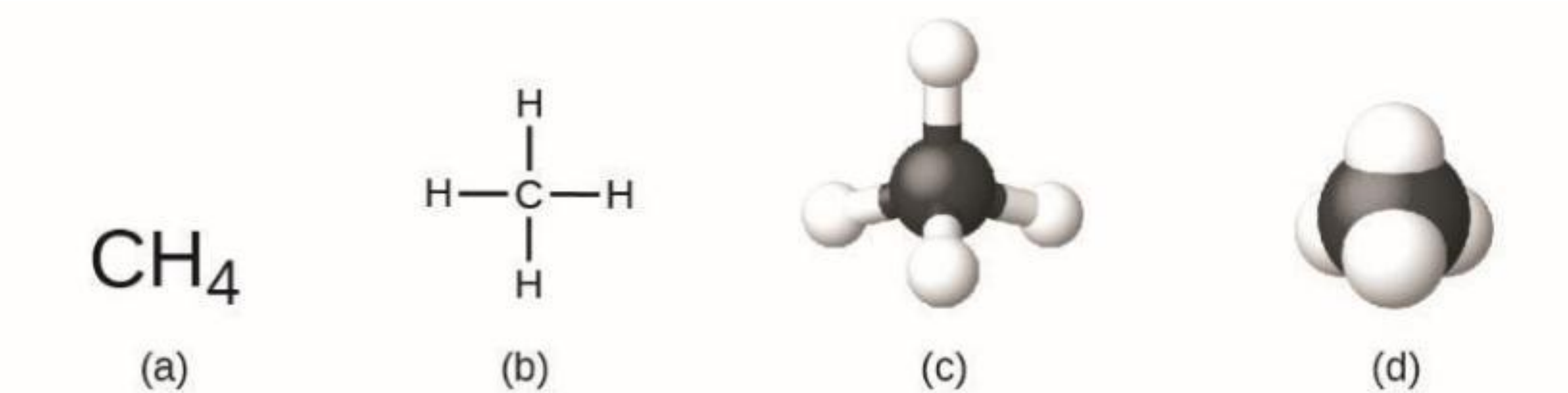
Chemical Formulas

- **Molecular formula**: A representation of a molecule or compound which consists of the following:
  1) Chemical symbols to indicate the types of atoms.
  2) Subscripts after the symbol to indicate the number of each type of atom in the molecule.

  - Subscripts are used only when more than one atom of a given type is present.

- A **structural formula** shows the same information as a molecular formula but also shows how the atoms are connected.

Figure 2.16

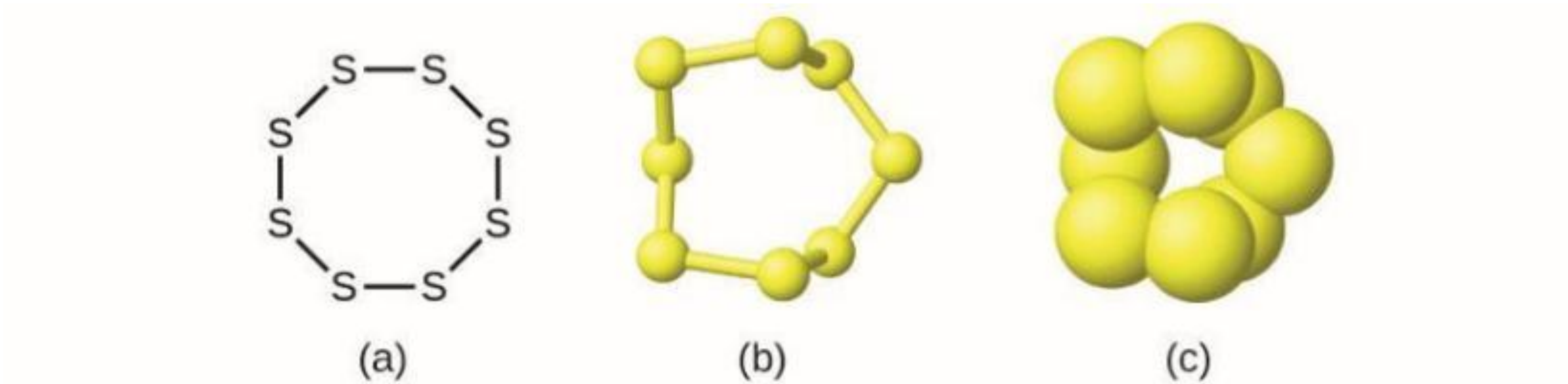


A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.

Elements That Exist as Molecules

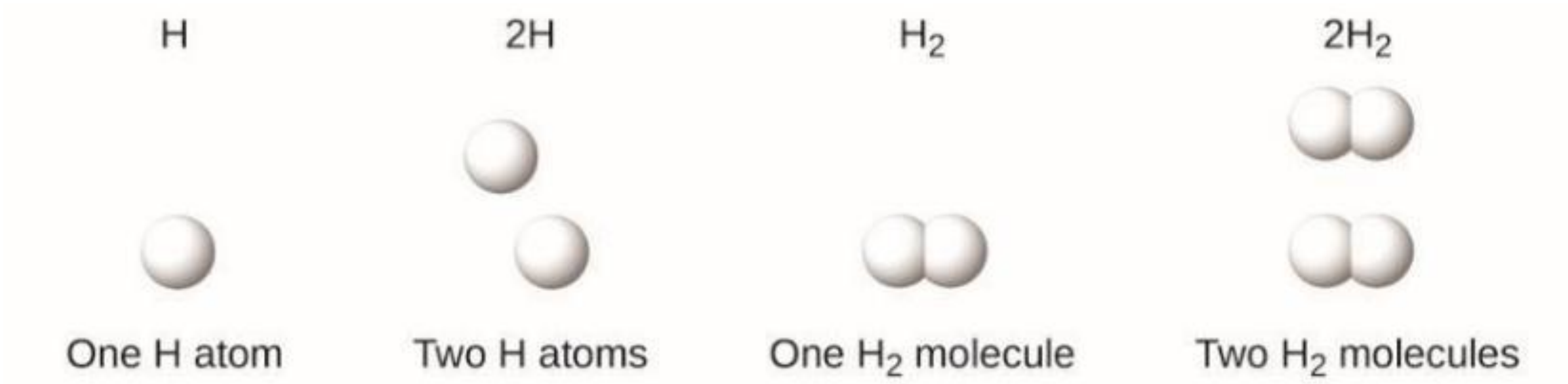
- Many elements consist of discrete, individual atoms.
- Some elements exist as molecules.
- Diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2
- The most common form of elemental sulfur exists as S_8.

Figure 2.17



A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

Figure 2.18



The symbols H, 2H, H_2, and $2H_2$ represent very different entities.

Empirical Formula

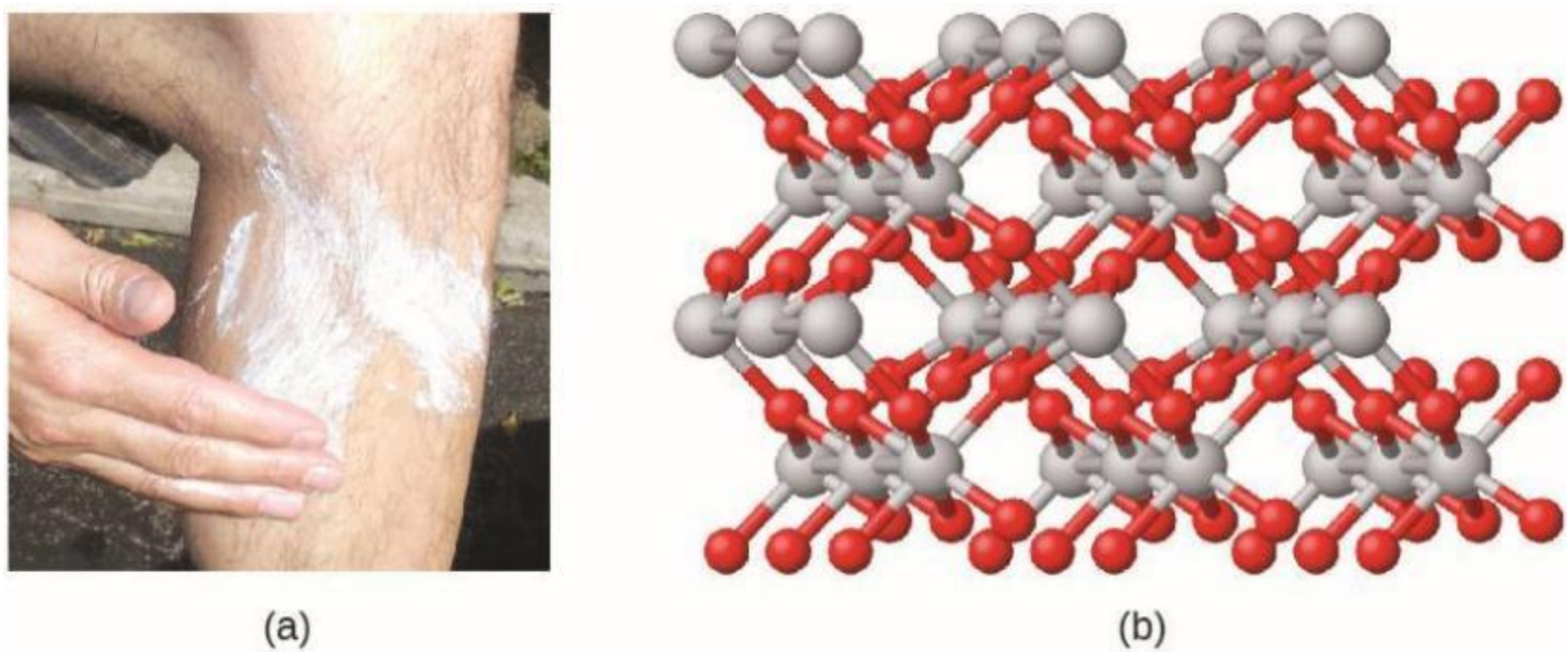
- An **empirical formula** indicates the simplest whole-number ratio of the number of atoms (or ions) in the compound.

- A **molecular formula** indicates the actual numbers of atoms of each element in a molecule of the compound.

- Example: benzene
  - Molecular formula = C_6H_6 Empirical formula = CH

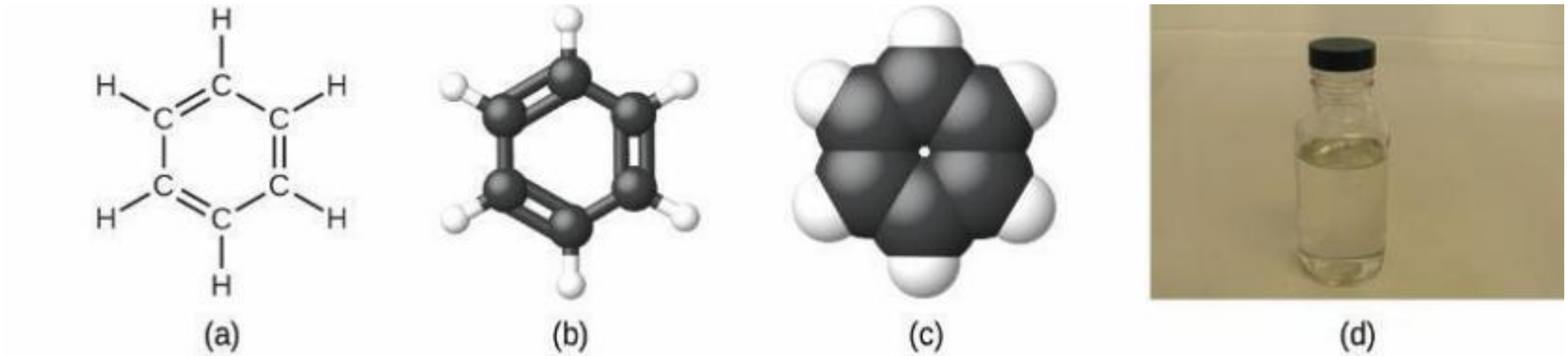
- Example: acetic acid
  - Molecular formula = $C_2H_4O_2$ Empirical formula = CH_2O

Figure 2.19



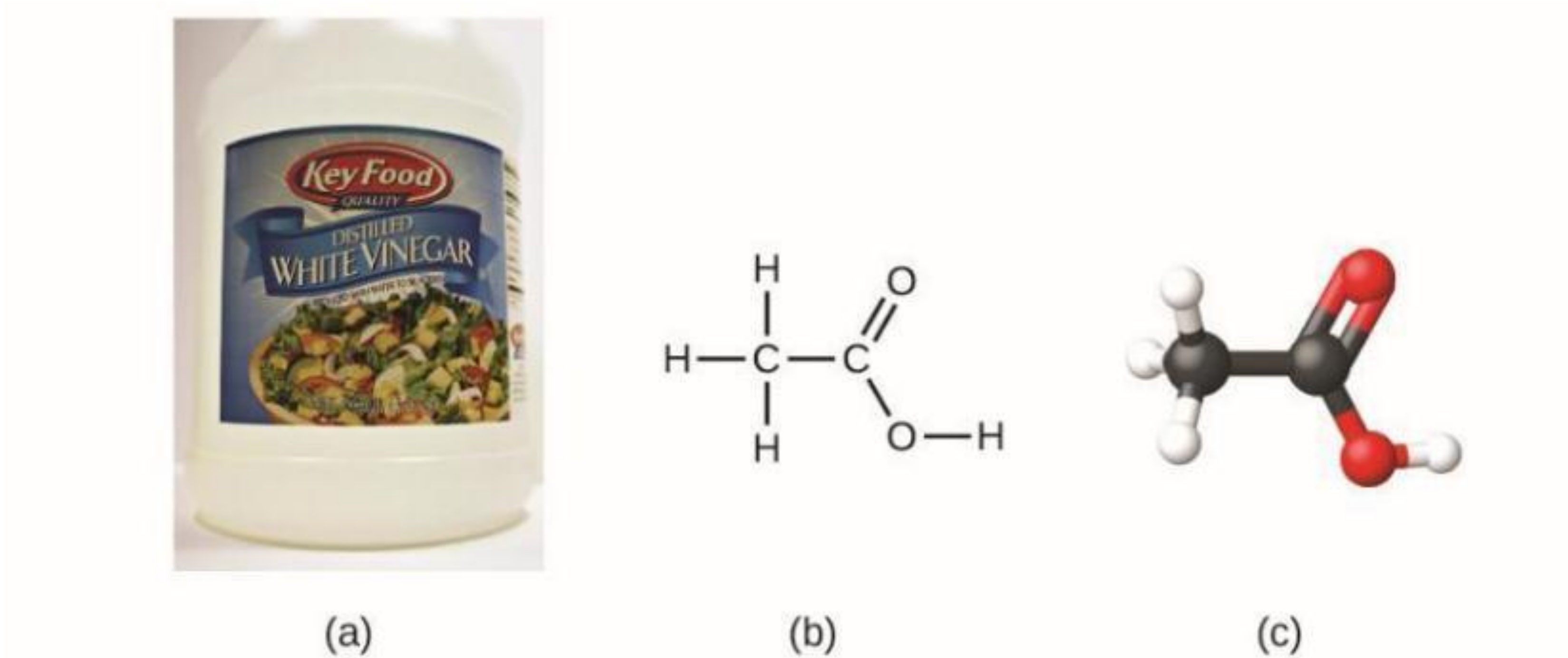
(a) The white compound titanium dioxide provides effective protection from the sun. (b) A crystal of titanium dioxide, TiO_2, contains titanium and oxygen in a ratio of 1 to 2. The titanium atoms are gray and the oxygen atoms are red. (credit a: modification of work by "osseous"/Flickr)

Figure 2.20



Benzene, C_6H_6, is produced during oil refining and has many industrial uses. A benzene molecule can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. (d) Benzene is a clear liquid. (credit d: modification of work by Sahar Atwa)

Figure 2.21



(a) Vinegar contains acetic acid, $C_2H_4O_2$, which has an empirical formula of CH_2O. It can be represented as (b) a structural formula and (c) as a ball-and-stick model. (credit a: modification of work by "HomeSpot HQ"/Flickr)

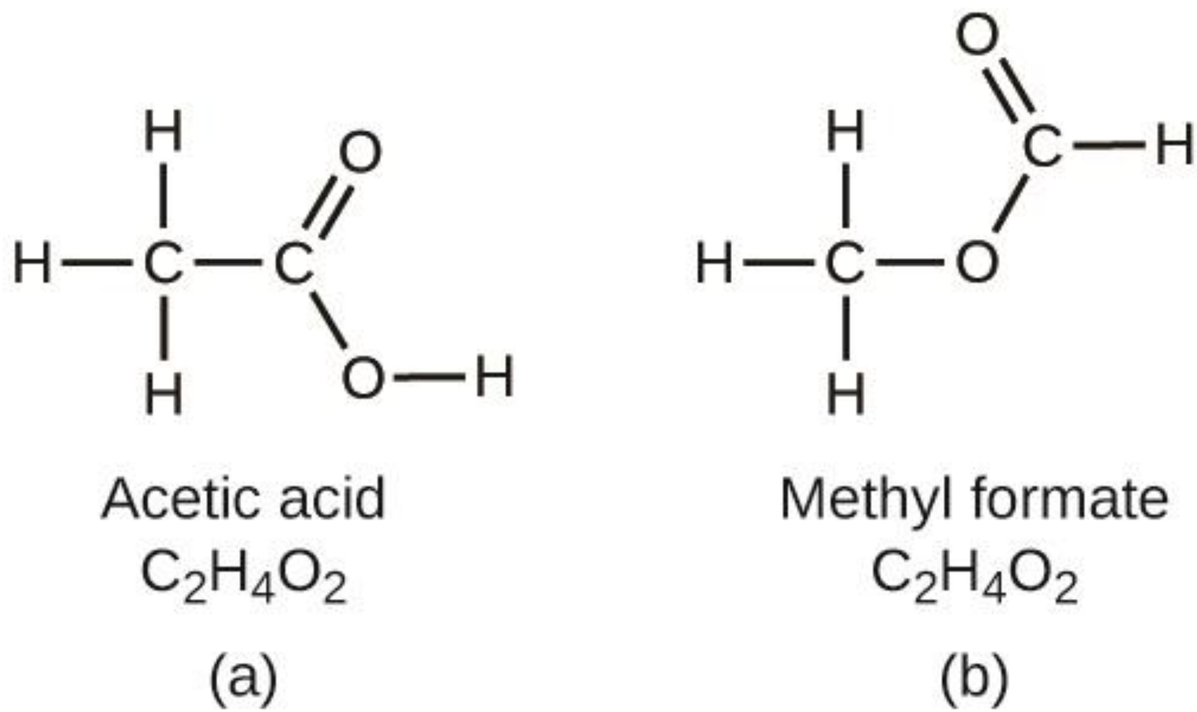
Figure 2.22

Chemist Lee Cronin has been named one of the UK's 10 most inspirational scientists. The youngest chair at the University of Glasgow, Lee runs a large research group, collaborates with many scientists worldwide, has published over 250 papers in top scientific journals, and has given more than 150 invited talks. His research focuses on complex chemical systems and their potential to transform technology, but also branches into nanoscience, solar fuels, synthetic biology, and even artificial life and evolution. (credit: image courtesy of Lee Cronin)

Isomers

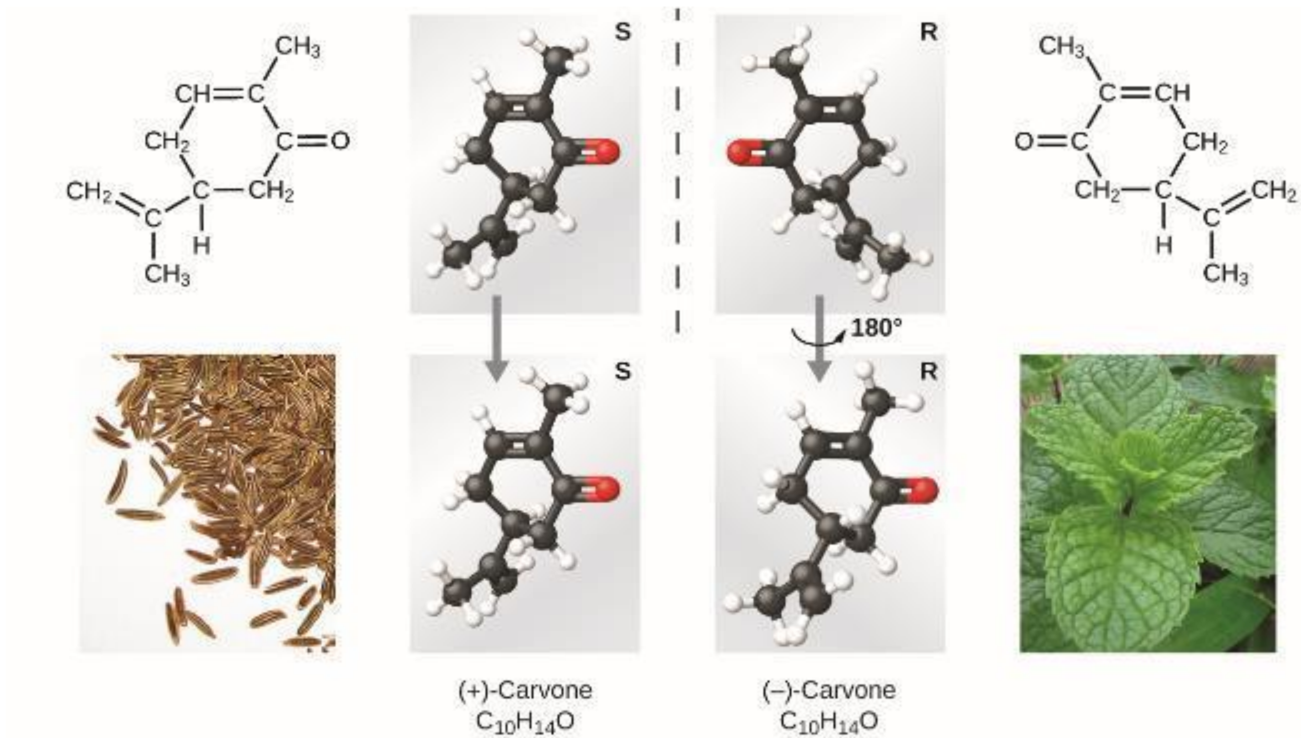
- It may be possible for the same atoms to be arranged in different ways.
- **Isomers:** Compounds with the same chemical formula but different molecular structures.
- Example: Acetic acid and methyl formate both have the molecular formula $C_2H_4O_2$, but they have different structures and properties.

Figure 2.23



Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula ($C_2H_4O_2$) but different structures (and therefore different chemical properties).

Figure 2.24



Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by “Miansari66”/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)

Learning Objectives

- 2.5 The Periodic Table
  - State the periodic law and explain the organization of elements in the periodic table
  - Predict the general properties of elements based on their location within the periodic table
  - Identify metals, nonmetals, and metalloids by their properties and/or location on the periodic table

The Periodic Table

- Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time.

- For example:
  - Lithium (Li), sodium (Na), and potassium (K) are all shiny, conduct heat and electricity well, and have similar chemical properties.
  - Calcium (Ca), strontium (Sr), and barium (Ba) are also shiny, conduct heat and electricity well, but are less reactive than Li, Na, and K.

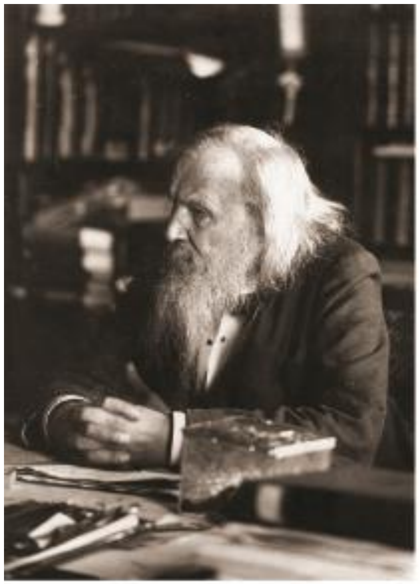
The First Periodic Table

- Both Mendeleev and Meyer published tables with the elements arranged according to increasing atomic mass.

- Mendeleev used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were not yet known.

- The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev’s work.

Figure 2.25



(a)

Reihen	Gruppo I. — R^2O	Gruppo II. — RO	Gruppo III. — R^2O^3	Gruppo IV. RH^4 RO^2	Gruppo V. RH^3 R^2O^5	Gruppo VI. RH^2 RO^3	Gruppo VII. RH R^2O^7	Gruppo VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

(b)

(a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by “Den fjättrade ankan”/Wikimedia Commons)

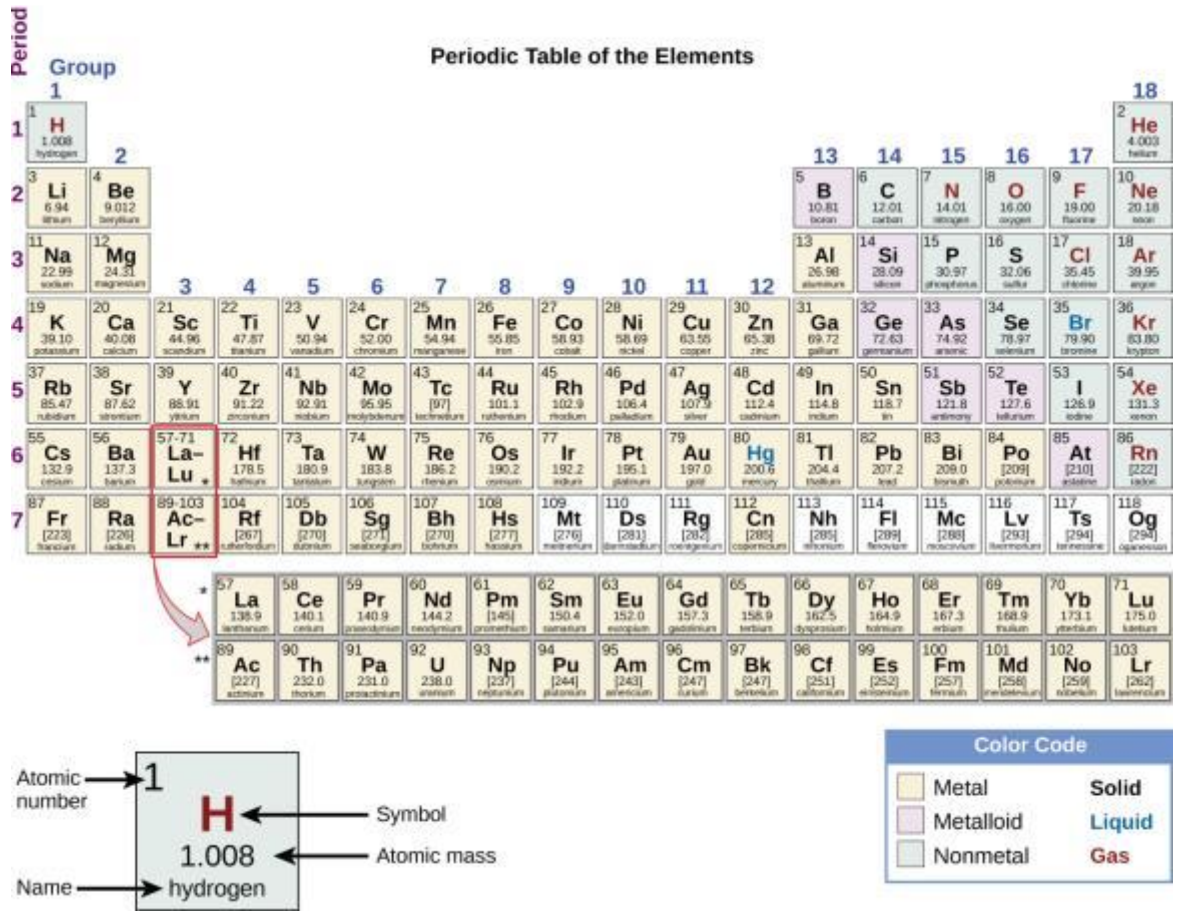
The Modern Periodic Table

- By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses.

- **Periodic Law:** The properties of the elements are periodic functions of their atomic numbers.

- A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column
  - Periods or series: horizontal rows
  - Groups: vertical columns (numbered 1–18)

Figure 2.26



Elements in the periodic table are organized according to their properties.

Classifications of Elements

- **Metals** are shiny, malleable, good conductors of heat and electricity.
- **Nonmetals** appear dull, poor conductors of heat and electricity.
- **Metalloids** conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals.

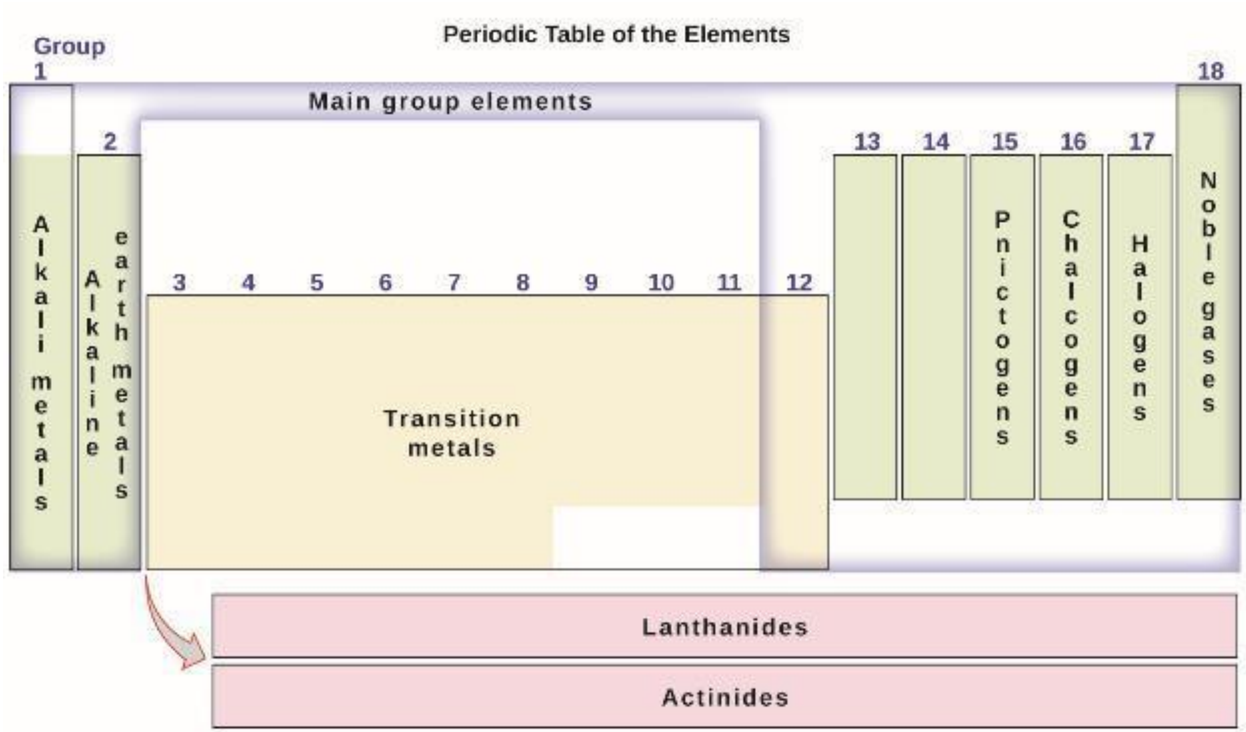
Classifications of Elements (continued)

- **Main group elements** (or representative elements)
  - Groups: 1, 2, 13–18
- **Transition metals**
  - Groups: 3–13
- **Inner transition metals**
  - Two rows at the bottom of the periodic table.
    - Lanthanides: top row
    - Actinides: bottom row

Classifications of Elements (continued)

- Alkali metals: group 1 (except hydrogen)
- Alkaline earth metals: group 2
- Pnictogens: group 15
- Chalcogens: group 16
- Halogens: group 17
- Noble Gases (or inert gases): group 18

Figure 2.27



The periodic table organizes elements with similar properties into groups.

Learning Objectives

- 2.6 Molecular and Ionic Compounds
  - Define ionic and molecular (covalent) compounds
  - Predict the type of compound formed from elements based on their location within the periodic table
  - Determine formulas for simple ionic compounds

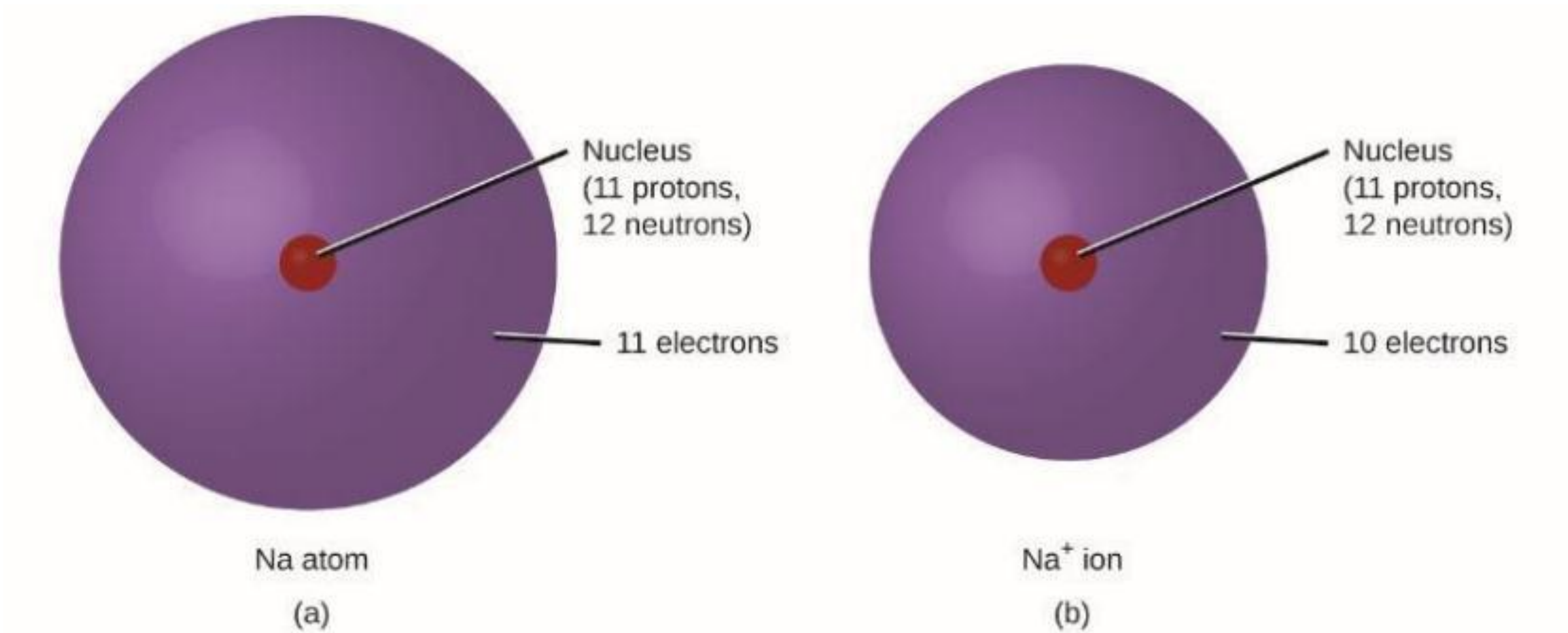
Molecular and Ionic Compounds

- In ordinary chemical reactions, the nucleus of each atom (and thus the identity of the element) remains unchanged.

- Electrons participate in chemical reactions by being gained, lost, or shared.

- The gain or lose of electrons, results in the formation of ions.

Figure 2.28



(a) A sodium atom (Na) has equal numbers of protons and electrons (11) and is uncharged. (b) A sodium cation (Na^+) has lost an electron, so it has one more proton (11) than electrons (10), giving it an overall positive charge, signified by a superscripted plus sign.

Predicting Ion Charge

- The periodic table can serve as a guide for predicting the ionic charge of main-group elements.

- Many **main-group metals** *lose* enough electrons to leave them with the same number of electrons as an atom of the preceding noble gas.
  - Group 1: lose one electron, form a cation with a 1+ charge
  - Group 2: lose two electrons, form a cation with a 2+ charge

- Many **nonmetals** *gain* enough electrons to give them the same number of electrons as an atom of the next noble gas.
  - Group 17: gain one electron, form an anion with a 1– charge.
  - Group 16: gain two electrons, form an anion with a 2– charge.

Predicting Ion Charge (continued)

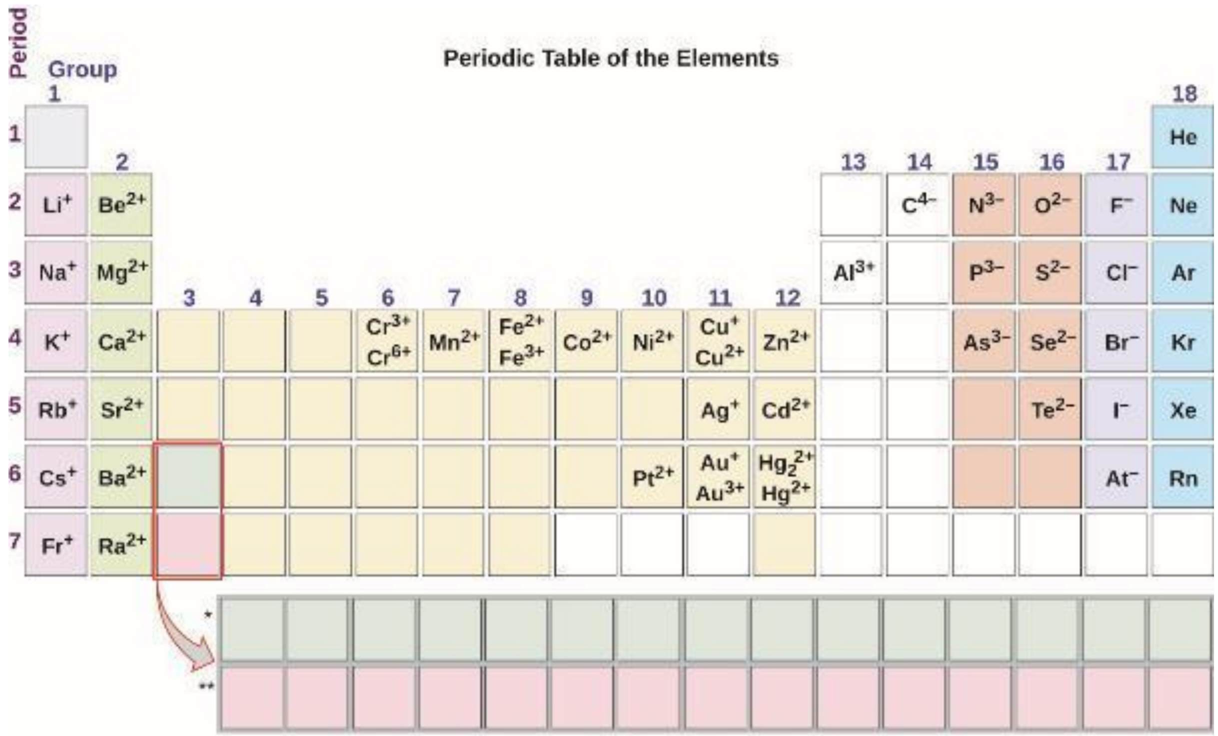
- Example: **Ca (group 2)**
  - Ca atom (20 protons, 20 electrons)
  - Loses 2 electrons
  - Now a Ca^{2+} ion (20 protons, 18 electrons)
  - Same number of electrons as the preceding noble gas, Ar.

- Example: **Br (group 17)**
  - Br atom (35 protons, 35 electrons)
  - Gains 1 electron
  - Now a Br^{-} ion (35 protons, 36 electrons)
  - Same number of electrons as the next noble gas, Kr.

Predicting Ion Charge (continued)

- Moving from far left to far right in the periodic table:
  - Positive charges of cations are equal to the group number.
- Moving from the far right to the far left in the periodic table:
  - Negative charges of anions are equal to the number of groups moved left from the noble gas.
- This method is less reliable for transition metals.
  - Cu forms ions of 1+ and 2+ charge.
  - Fe forms ions of 2+ and 3+ charge.

Figure 2.29



Some elements exhibit a regular pattern of ionic charge when they form ions.

Polyatomic Ions

- The ions that we have discussed so far are called **monatomic ions**, that is, they are ions formed from only one atom.
- **Polyatomic ions** are electrically charged molecules (a group of bonded atoms with an overall charge).
- **Oxyanions** are polyatomic ions that contain one or more oxygen atoms.

Table 2.5 (Partial) Common Polyatomic Ions

Name	Formula	Related Acid	Formula
ammonium	NH_4^+	nitrate	NO_3^-
hydronium	H_3O^+	nitrate	NO_2^-
peroxide	O_2^{2-}	sulfate	SO_4^{2-}
hydroxide	OH^-	hydrogen sulfate	HSO_4^-
acetate	CH_3COO^-	sulfite	SO_3^{2-}
cyanide	CN^-	hydrogen sulfite	HSO_3^-
azide	N_3^-	phosphate	PO_4^{3-}
carbonate	CO_3^{2-}	hydrogen phosphate	HPO_4^{2-}
bucarbonate	HCO_3^-		
dihydrogen phosphate	$H_2PO_4^-$		
perchlorate	ClO_4^-		
chlorate	ClO_3^-		
chlorite	ClO_2^-		
hypochlorite	ClO^-		
chromate	CrO_4^{2-}		
dichromate	$Cr_2O_7^{2-}$		
permanganate	MnO_4^-		

Naming Oxyanions

- There is a system for naming oxyanions.

- When a nonmetal forms two oxyanions:
  - *–ate* is the suffix used for the ion with the larger number of oxygen atoms
  - *–ite* is the suffix used for the ion with the smaller number of oxygen atoms

- When a nonmetal forms more than two oxyanions, prefixes are used in addition to *–ate* and *–ite*.
  - *per–* (largest number of oxygens)
  - *hypo* – (smallest number of oxygens)

Types of Chemical Bonds

- When electrons are transferred, ions form, and an **ionic bond** results.
- Ionic bonds are electrostatic forces of attraction.
- When electrons are shared and molecules form, a **covalent bond** results.
- Compounds are classified as ionic or molecular (covalent) on the basis of the bonds present in them.

Ionic Compounds

- Metals readily lose electrons—form cations.
- Nonmetals readily gain electrons—form anions.
- When a metal and nonmetal react, a transfer of electrons usually takes place.
- Metals and nonmetals generally form ionic compounds.
- A compound that contains ions and is held together by ionic bonds is called an **ionic compound**.

Ionic Compound Examples

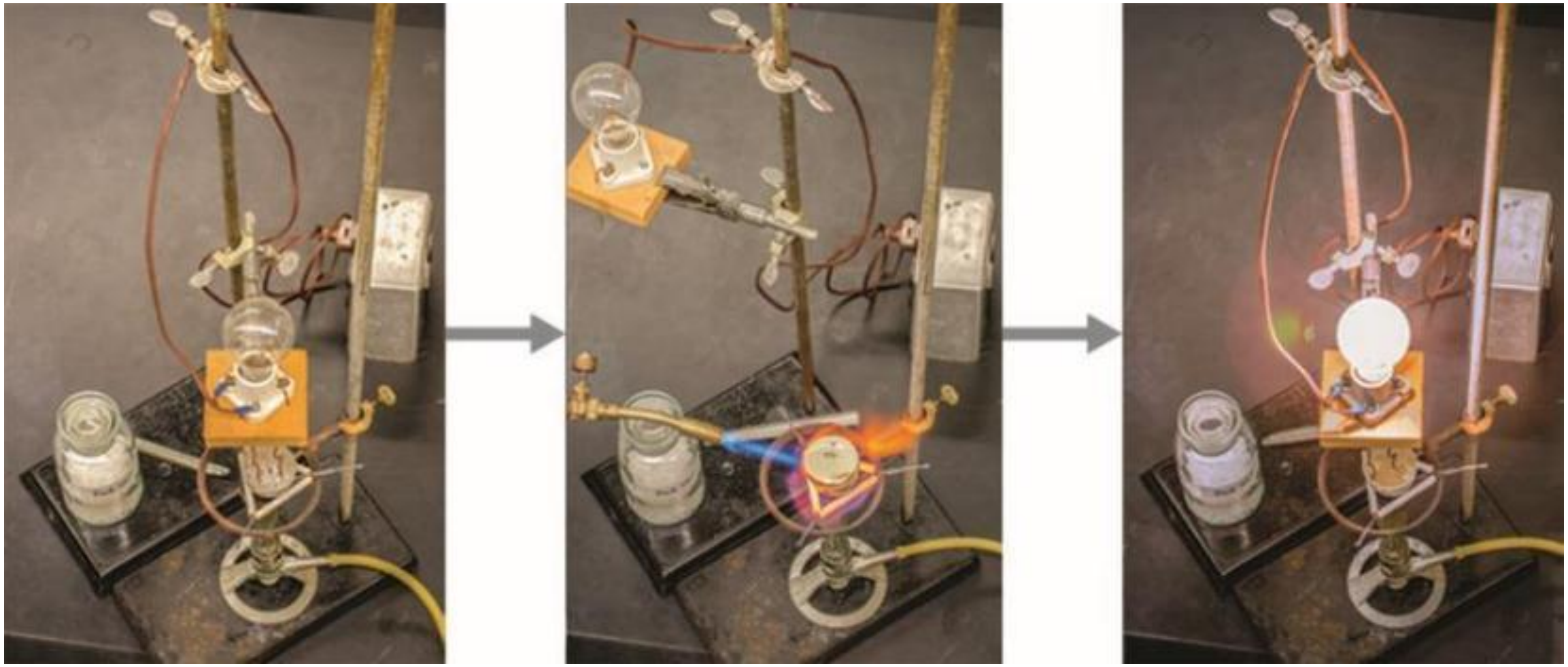
- Na and Cl
  - One Na atom gives up one electron forming a Na^{+} ion.
  - One Cl atom accepts that electron forming a Cl^{-} ion.
  - The ionic compound, NaCl forms.

- Ca and Cl
  - One Ca atom gives up two electrons forming a Ca^{2+} ion.
  - Two Cl atoms each accept one electron forming two Cl^{-} ions.
  - The ionic compound, $CaCl_2$ forms.

Properties of Ionic Compounds

- Typically solids with high melting and boiling points.
- Nonconductive in solid form.
- Conductive in molten form.

Figure 2.30



Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)

Formulas of Ionic Compounds

- Ionic compounds are electrically neutral overall.
- The formula of an ionic compound must have a ratio of ions such that the numbers of positive and negative charges are equal.
- These formulas are not molecular formulas.
- Example: Al^{3+} and O^{2-} forms Al_2O_3
  - Two Al^{3+} ions gives six positive charges.
  - Three O^{2-} ions gives six negative charges.

Figure 2.31

Although pure aluminum oxide is colorless, trace amounts of iron and titanium give blue sapphire its characteristic color. (credit: modification of work by Stanislav Doronenko)

Formulas of Ionic Compounds

- Many ionic compounds contain polyatomic ions as the cation, the anion, or both.
- Treat polyatomic ions as discrete units.
- Parentheses in a formula are used to indicate a group of atoms that behave as a unit.
- Example: Ca^{2+} and PO_4^{3-} forms $Ca_3(PO_4)_2$
  - Three Ca^{2+} ions gives six positive charges.
  - Two PO_4^{3-} ions gives six negative charges.

Molecular Compounds

- Molecular compounds (covalent compounds) result when atoms share electrons.
- Exist as discrete, neutral molecules.
- Usually formed by a combination of nonmetals.
- Often exist as gases, low-boiling liquids, and low-melting solids.

Learning Objectives

- 2.7 Chemical Nomenclature
  - Derive names for common types of inorganic compounds using a systematic approach

Chemical Nomenclature

- **Nomenclature:** A collection of rules for naming things.

- Compounds are identified by both their formula and name.

- We will learn how to name the following types of inorganic compounds:
  - **Ionic and molecular binary compounds**: composed of two elements.
  - Ionic compounds containing polyatomic ions.
  - Acids

Naming Ionic Compounds

- Name the cation first, followed by the name of the anion.
- A monoatomic cation is just given the name of the element.
- A monoatomic anion is given the name of the element with its ending replaced by the suffix –*ide*.
- A polyatomic ion is just given the name of the ion.

Table 2.6 Names of Some Ionic Compounds

NaCl, sodium chloride	Na_2O, sodium oxide
KBr, potassium bromide	CdS, cadmium sulfide
CaI_2, calcium iodide	Mg_3N_2, magnesium nitride
CsF, cesium fluoride	Ca_3P_2, calcium phosphide
LiCl, lithium chloride	Al_4C_3, aluminum carbide

Table 2.7 Names of Some Polyatomic Ionic Compounds

$KC_2H_3O_2$, potassium acetate	$CaSO_4$, calcium sulfate
NH_4Cl, ammonium chloride	$Al_2(CO_3)_3$, aluminum carbonate
$NaHCO_3$, sodium bicarbonate	$Mg_3(PO_4)_2$, magnesium phosphate

Naming Ionic Compounds Containing a Metal Ion with a Variable Charge

- Most of the transition metals and some main group metals can form two or more cations with different charges.
- The charge of the metal ion is specified by a Roman numeral in parentheses after the name of the metal.

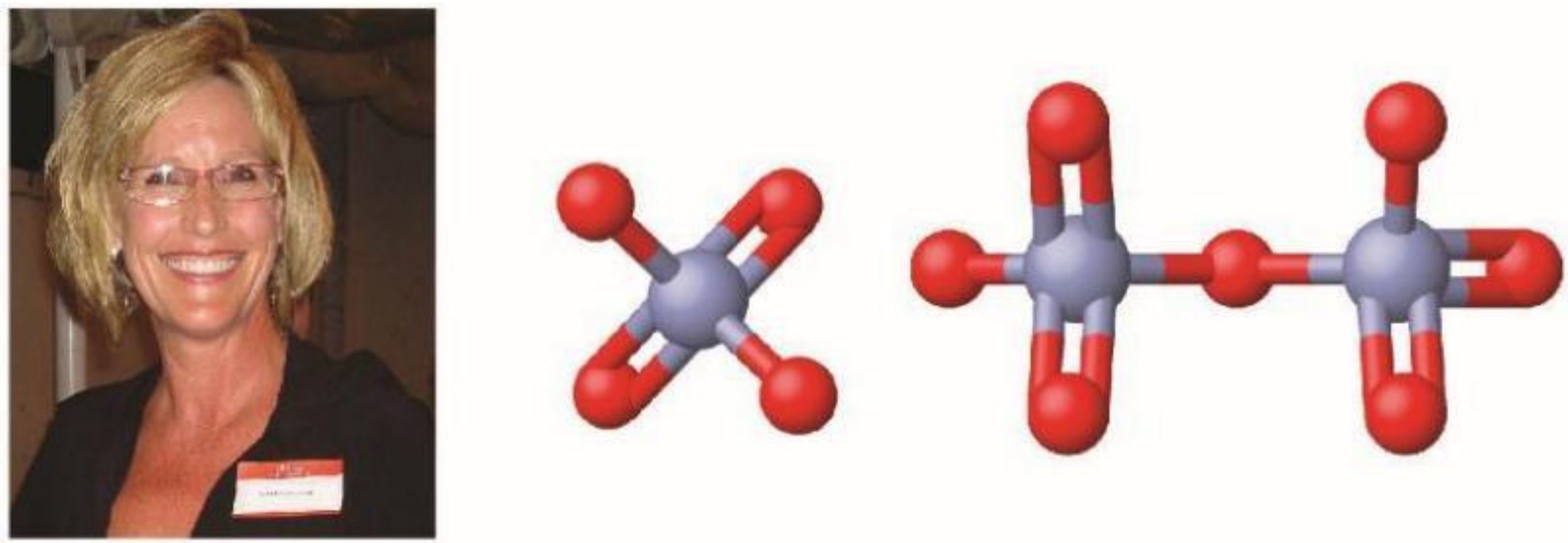
Table 2.9 Some Ionic Compounds with Variably Charged Metal Ions

Compound	Name
$FeCl_2$	iron(II) chloride
$FeCl_3$	iron(III) chloride
Hg_2O	mercury(I) oxide
HgO	mercury(II) oxide
SnF_2	tin(II) fluoride
SnF_4	tin(IV) fluoride

Table 2.10 Nomenclature Prefixes

Number	Prefix		Number	Prefix
1 (sometimes omitted)	mono-		6	hexa-
2	di-		7	hepta-
3	tri-		8	octa-
4	tetra-		9	nona-
5	penta-		10	deca-

Figure 2.32



(a) (b)

(a) Erin Brockovich found that Cr(IV), used by PG&E, had contaminated the Hinckley, California, water supply. (b) The Cr(VI) ion is often present in water as the polyatomic ions chromate, CrO_4^{2-} (left), and dichromate, $Cr_2O_7^{2-}$ (right).

Naming Ionic Hydrates

- **Hydrate:** Compound, often ionic, that contains one or more water molecules bound within its crystals.
- Hydrates may typically be dehydrated by heating to remove the bound water molecules, yielding the anhydrous compound.
- To name hydrates:
  1) Name the anhydrous compound (per usual rules)
  2) Add the word hydrate with a Greek prefix denoting the number of water molecules
- Formulas for hydrates are written by appending the formula for water to the formula for the anhydrous compound, including a stoichiometric coefficient denoting the number of water molecules, and separated by a vertically centered dot
- Examples:
  copper(II) sulfate pentahydrate $CuSO_4{\cdot}5H_2O$
  calcium chloride monohydrate $CaCl_2{\cdot}H_2O$

Naming Binary Molecular (Covalent) Compounds

- Molecular compounds are name using a different set of rules.
- Covalent bonding allows for significant variation in the ratios of the atoms in a molecule.
- The names for molecular compounds must explicitly identify these ratios.
- The name of the more metallic element (the one farther to the left and/or bottom of the periodic table) is named first.
- Followed by the name of the more nonmetallic element (the one farther to the right and/or top) with its ending changed to the suffix –*ide*.
- The numbers of atoms of each element are designated by Greek prefixes.

Nomenclature Prefixes

- When only one atom of the first element is present, the prefix *mono-* is usually not used.

- When two vowels are adjacent, the *a* in the Greek prefix is usually dropped.

Table 2.11 Names of Some Molecular Compounds Composed of Two Elements

Compound	Name		Compound	Name
SO_2	sulfur dioxide		BCl_3	boron trichloride
SO_3	sulfur trioxide		SF_6	sulfur hexafluoride
NO_2	nitrogen dioxide		PF_5	phosphorus pentafluoride
N_2O_4	dinitrogen tetroxide		P_4O_{10}	tetraphosphorus decaoxide
N_2O_5	dinitrogen pentoxide		P_4O_{10}	iodine heptafluoride

Naming Acids

- Some compounds containing hydrogen are members of an important class of substances known as acids.
- Many acids release hydrogen ions, H^+, when dissolved in water.
- A mixture of an acid with water is given a special name to denote this property.

Naming Binary Acids

1) The word “hydrogen” is changed to the prefix *hydro*–.
2) The other nonmetallic element name is modified by adding the suffix –*ic*.
3) The word “acid” is added as a second word.

Table 2.12 Names of Some Simple Acids

Name of Gas	Name of Acid
HF(*g*), hydrogen fluoride	HF(*aq*), hydrofluoric acid
HCl(*g*), hydrogen chloride	HCl(*aq*), hydrochloric acid
HBr(*g*), hydrogen bromide	HBr(*aq*), hydrobromic acid
HI(*g*), hydrogen iodide	HI(*aq*), hydroiodic acid
H_2S(*g*), hydrogen sulfide	H_2S(*aq*), hydrosulfuric acid

Naming Oxyacids

- **Oxyacids:** Compounds that contain hydrogen, oxygen, and at least one other element, and are bonded in such a way as to impart acidic properties to the compound.

- Typical oxyacids consist of hydrogen combined with a polyatomic, oxygen-containing ion.

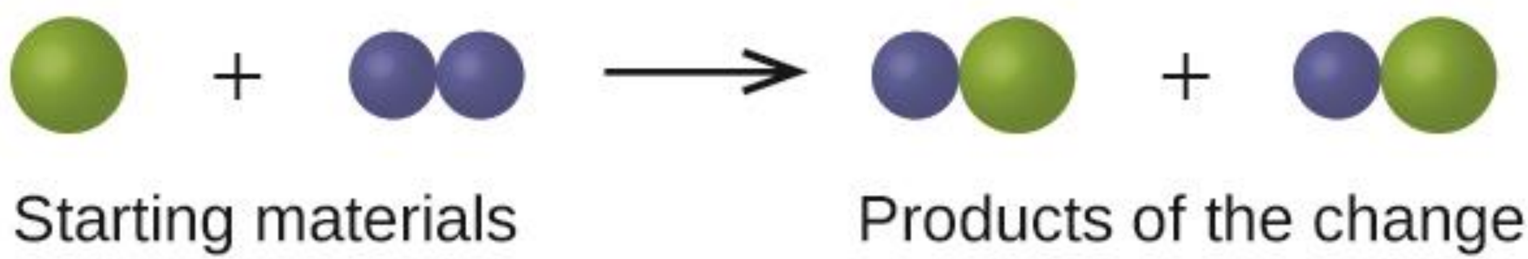
- To name oxyacids:

1) Omit "hydrogen"
2) Start with the root name of the anion
3) Replace –ate with –ic, or –ite with –ous
4) Add "acid"

Table 2.13 Names of Common Oxyacids

Formula	Anion Name	Acid Name
$HC_2H_3O_2$	acetate	acetic acid
HNO_3	nitrate	nitric acid
HNO_2	nitrite	nitrous acid
$HClO_4$	perchlorate	perchloric acid
H_2CO_3	carbonate	carbonic acid
H_2SO_4	sulfate	sulfuric acid
H_2SO_3	sulfite	sulfurous acid
H_3PO_4	phosphate	phosphoric acid

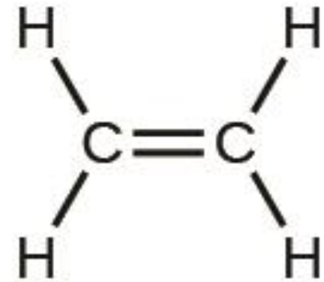
Exercise 1

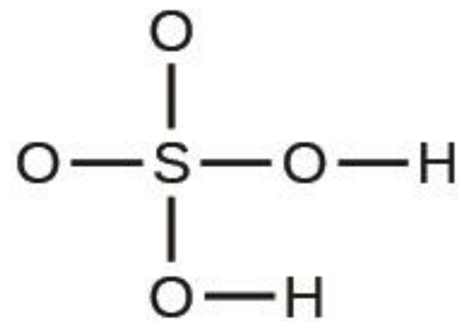


Exercise 29

(a) 

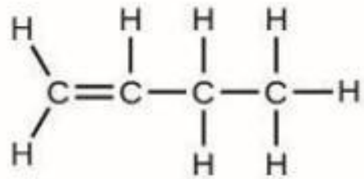
(b) 

(c) 

(d) 

Exercise 30

(a)



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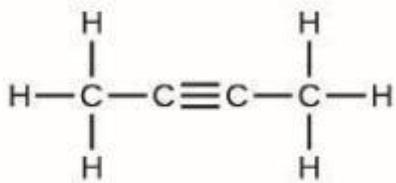

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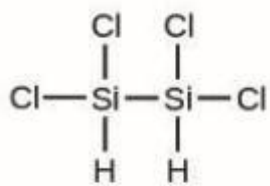

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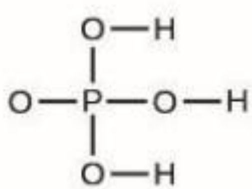
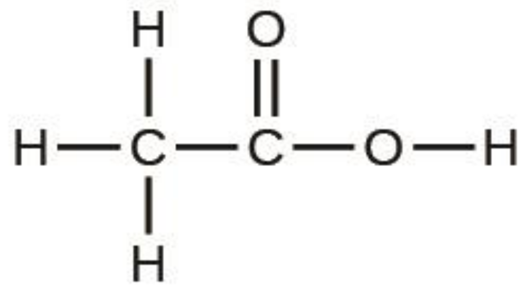


Exercise 33

(a)



(b)