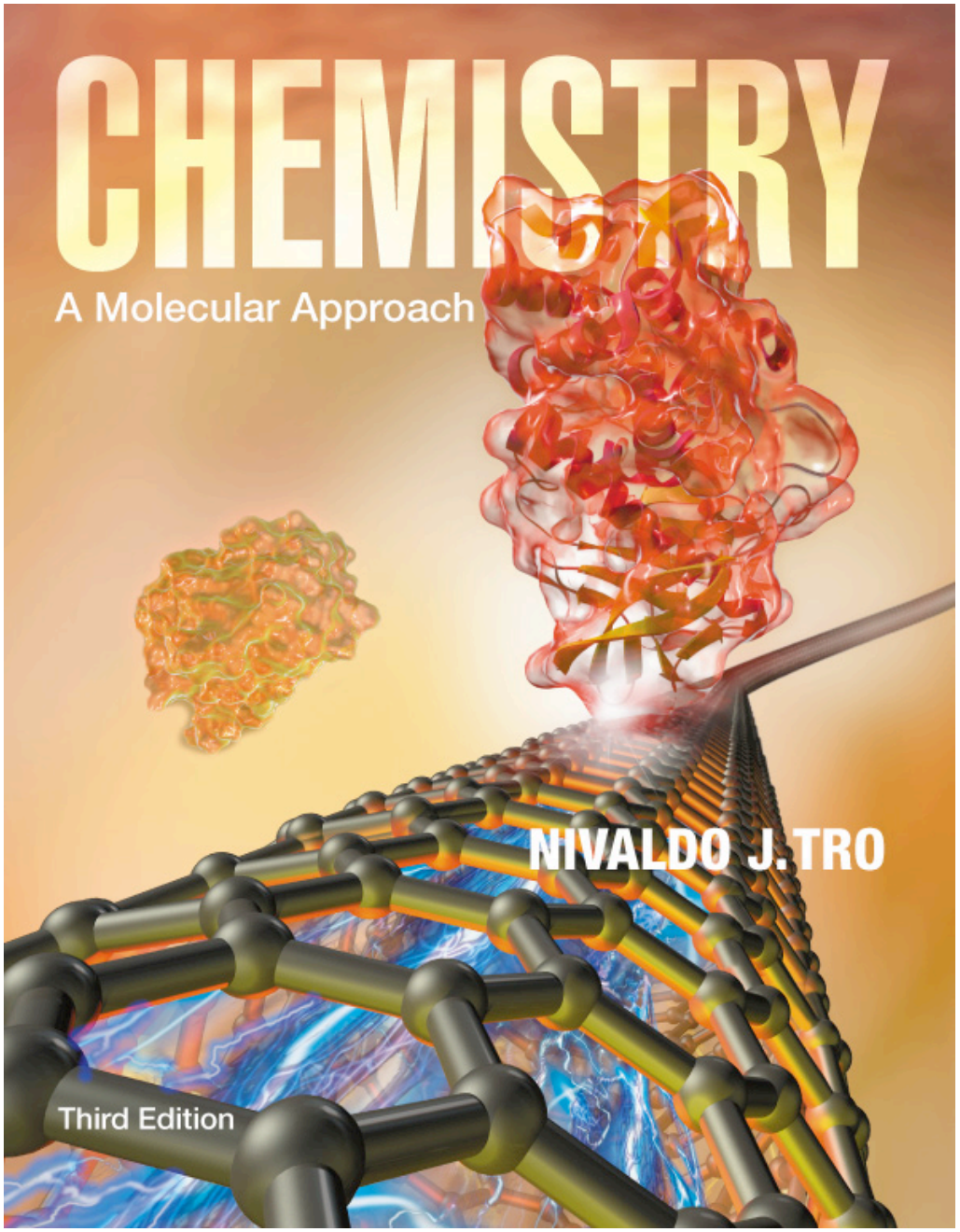


Lecture Presentation

Chapter 1

Matter, Measurement, and Problem Solving

Christian Madu, Ph.D.
Collin College

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What Do You Think?

- **What do you think** is the most important idea in all of human knowledge?

- If we limit ourselves only to scientific answers, it would be this:

  - **The properties of matter are determined by the properties of molecules and atoms**.

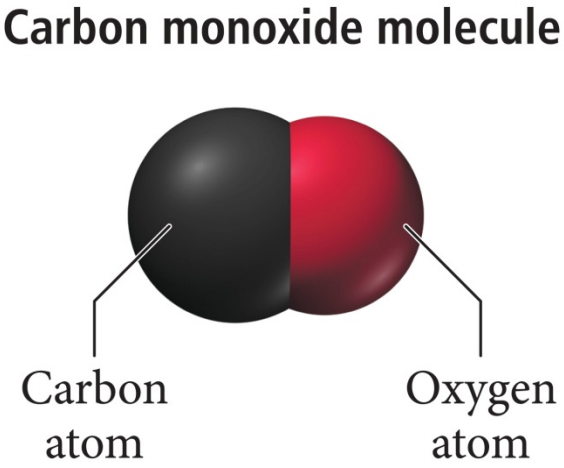
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What Do You Think?

- Atoms and molecules determine how matter behaves; if they were different, matter would be different.
  - The properties of water molecules determine how water behaves; the properties of sugar molecules determine how sugar behaves.
- The understanding of matter at the molecular level gives us unprecedented control over that matter.

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Atoms and Molecules



- The air contains carbon monoxide pollutant.
- Carbon monoxide gas is composed of carbon monoxide **molecules**.
- Each molecule contains a carbon **atom** and an oxygen **atom** held together by a chemical bond.

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Atoms and Molecules

- **Atoms** are the submicroscopic particles that constitute the fundamental building blocks of ordinary matter.

- Free atoms are rare in nature; instead they bind together in specific geometrical arrangements to form **molecules**.

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Atoms and Molecules

- If we want to understand the substances around us, we must understand the atoms and molecules that compose them—this is the central goal of chemistry.

  – **Chemistry is the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules.**

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The Scientific Approach to Knowledge

- The approach to scientific knowledge is empirical—it is based on ***observation*** and ***experiment***.

- The scientific method is a process for understanding nature by observing nature and its behavior, and by conducting experiments to test our ideas.

- Key characteristics of the scientific method include **observation**, formulation of **hypotheses**, **experimentation**, and formulation of **laws and theories**.

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Observations

- Observations are also known as **data.**
- They are the descriptions about the characteristics or behavior of nature.
    - Antoine Lavoisier (1743–1794) noticed that there was no change in the total mass of material within the container during combustion.
- Observations often lead scientists to formulate a **hypothesis.**

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Hypothesis

- A hypothesis is a tentative interpretation or explanation of the observations.
    - For example, Lavoisier explained his observations on combustion by hypothesizing that when a substance burns, it combines with a component of air.
- A good hypothesis is *falsifiable.*
    - The results of an experiment may support a hypothesis or—prove it wrong—in which case the scientist must modify or discard the hypothesis.

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A Scientific Law

- A brief statement that summarizes past observations and predicts future ones
  - **Law of conservation of mass**— "In a chemical reaction matter is neither created nor destroyed."
- Allows you to predict future observations
  - So you can test the law with experiments
- Unlike state laws, you cannot choose to violate a scientific law.

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Theory

- One or more well-established hypotheses may form the basis for a scientific **theory**.
- A scientific theory is a model for the way nature is and tries to explain not merely what nature does, but why.
- Theories are validated by experiments.
- Theories can never be conclusively proven because some new observation or experiment always has the potential to reveal a flaw.

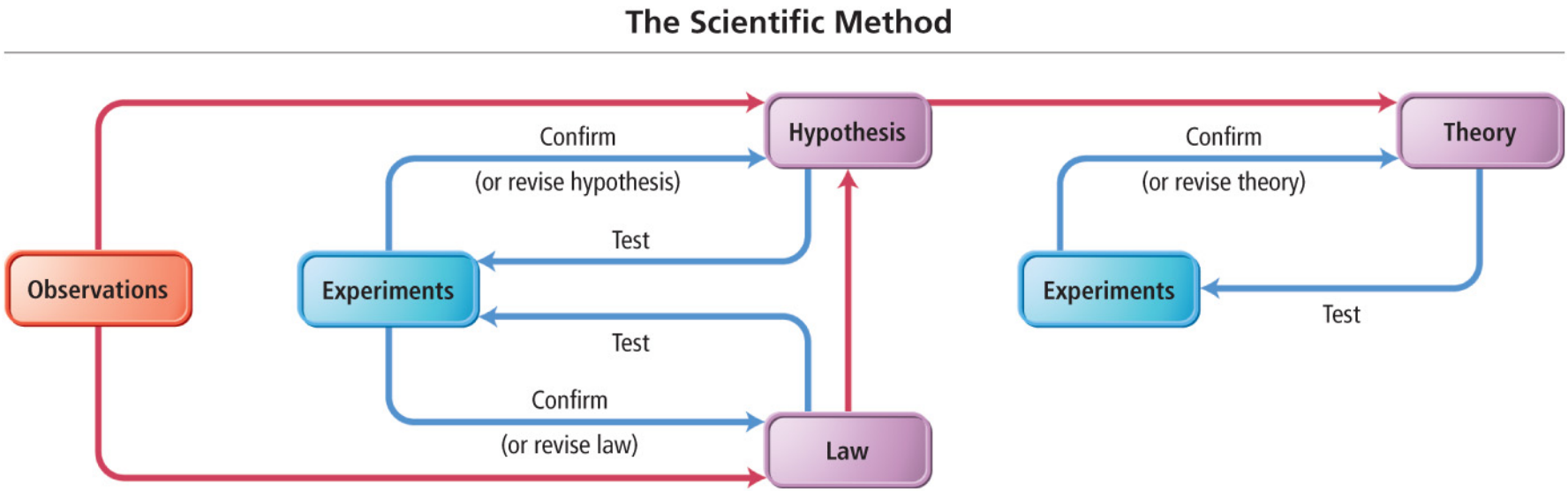
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Theory

- General explanation for the characteristics and behavior of nature
- Models of nature
    - Dalton's atomic theory
- Can be used to predict future observations
    - So they can be tested by experiments

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The Scientific Approach to Knowledge



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Conceptual Connection 1.1

Which statement best explains the difference between a law and a theory?

(a) A law is truth whereas a theory is a mere speculation.

(b) A law summarizes a series of related observations, while a theory gives the underlying reasons for them.

(c) A theory describes *what* nature does; a law describes *why* nature does it.

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Conceptual Connection 1.1

Which statement best explains the difference between a law and a theory?

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(c) A theory describes *what* nature does; a law describes *why* nature does it.

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The Classification of Matter

- **Matter** is anything that occupies space and has mass.
    - Your textbook, your desk, your chair, and even your body are all composed of matter.
- We can classify matter according to its **state** (its physical form) and its **composition** (the basic components that make it up).

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The States of Matter

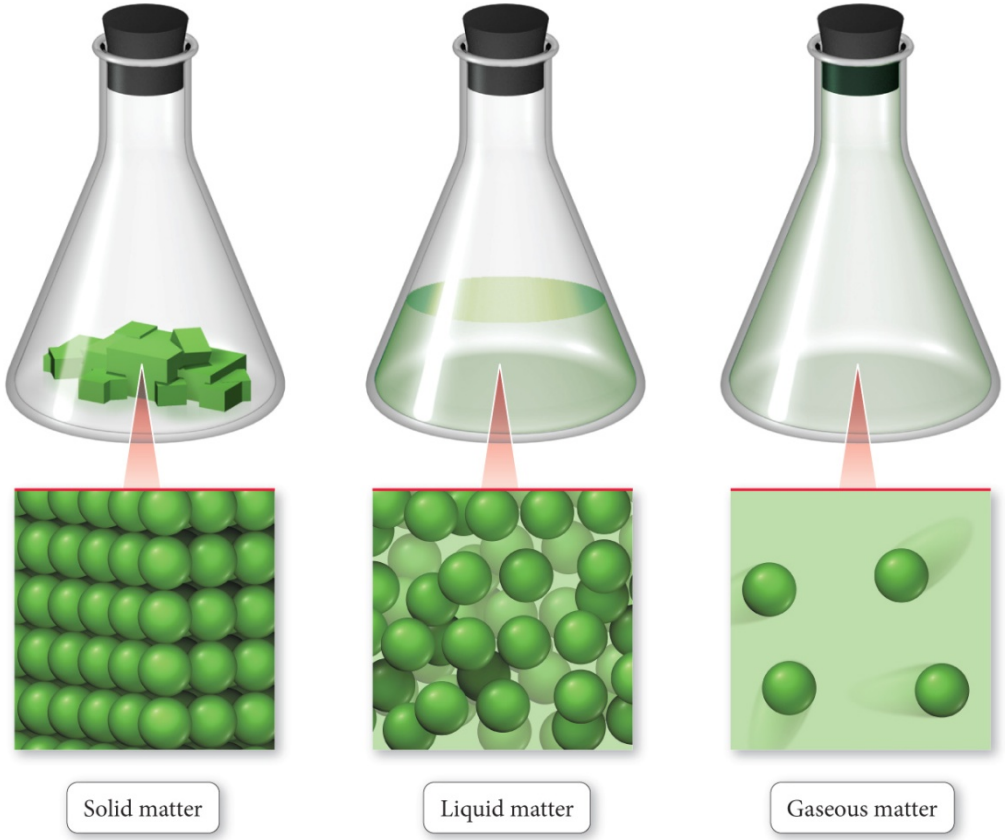
- **Matter** can be classified as solid, liquid, or gas based on what properties it exhibits.

- The state of matter changes from solid to liquid to gas with increasing temperature.

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Structure Determines Properties

- The atoms or molecules have different structures in solids, liquids, and gases—leading to different properties.



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Solid Matter

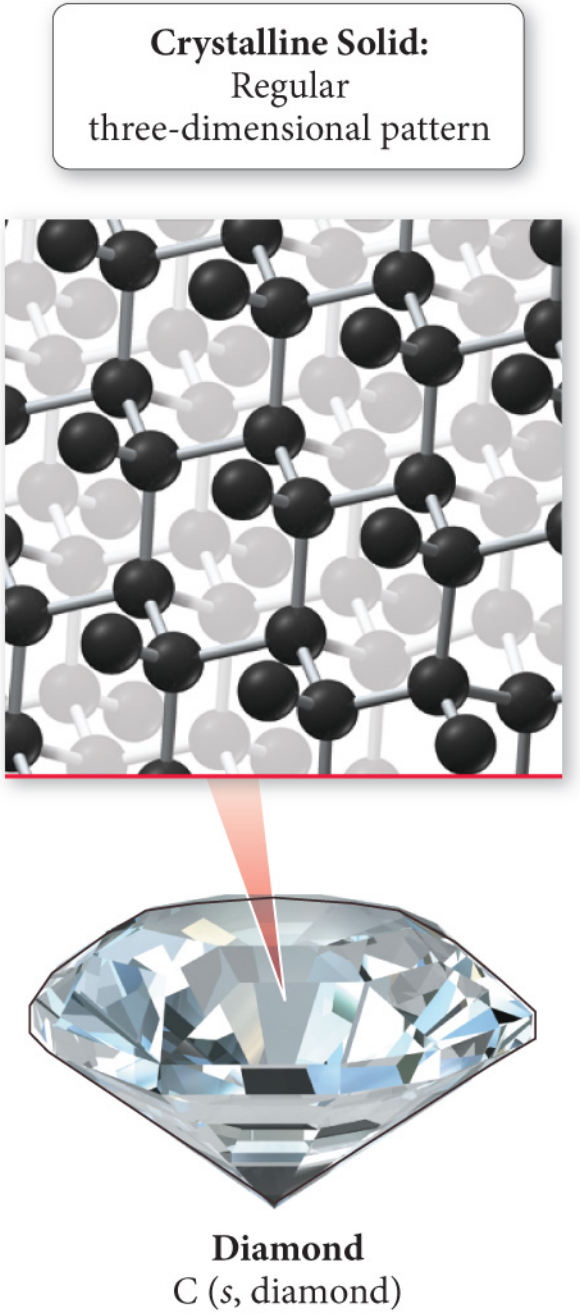
- In *solid matter*, atoms or molecules pack close to each other in fixed locations.
- Although the atoms and molecules in a solid vibrate, they do not move around or past each other.
- Consequently, a solid has a fixed volume and rigid shape.
    - Ice, aluminum, and diamond are good examples of solids.

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Solid Matter

- Solid matter may be **crystalline**—in which case its atoms or molecules are in patterns with long-range, repeating order.
  - Table salt and diamond are examples of solid matter.

- Others may be **amorphous**, in which case its atoms or molecules do not have any long-range order.
  - Examples of *amorphous* solids include glass and plastic.



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Liquid Matter

- In *liquid matter*, atoms or molecules pack about as closely as they do in solid matter, but they are free to move relative to each other.

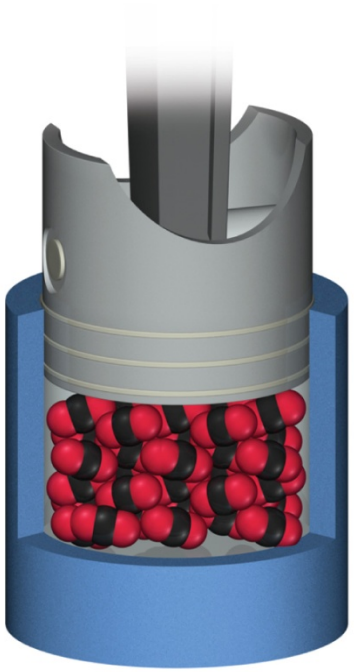
- Liquids have fixed volume but not a fixed shape.

- Liquids' ability to flow makes them assume the shape of their container.
    - Water, alcohol, and gasoline are all substances that are liquids at room temperature.

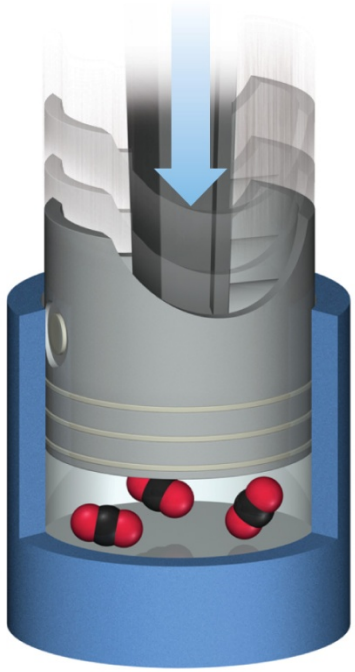
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Gaseous Matter

- In *gaseous matter*, atoms or molecules have a lot of space between them.
- They are free to move relative to one another.
- These qualities make gases *compressible*.



Solid–not compressible

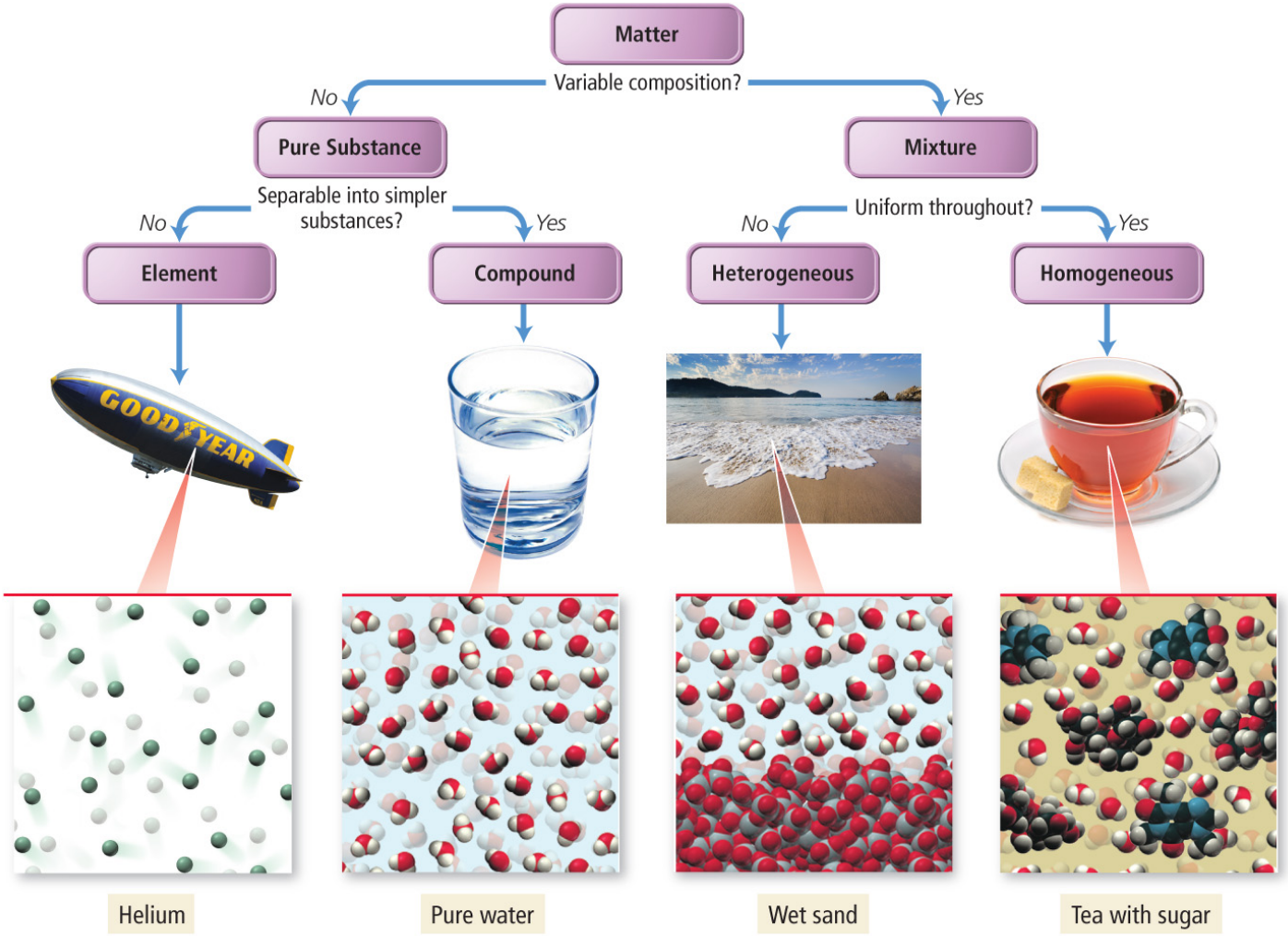


Gas–compressible

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The Classification of Matter by Components

- Matter can also be classified according to its composition: elements, compounds, and mixtures.



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Classification of Matter by Components

- The first division in the classification of matter is between a *pure substance* and a *mixture*.
- A **pure substance** is made up of only one component and its composition is invariant.
- A **mixture**, by contrast, is a substance composed of two or more components in proportions that can vary from one sample to another.

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Classification of Pure Substances

- Pure substances categorize into two types:
  - Elements
  - Compounds
- This categorization depends on whether or not they can be broken down (or decomposed) into simpler substances.

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Classification of Pure Substances

- An **element** is a substance that cannot be chemically broken down into simpler substances.
    - Basic building blocks of matter
    - Composed of single type of atom, like helium
- A **compound** is a substance composed of two or more elements in fixed definite proportions.
- Most elements are chemically reactive and combine with other elements to form compounds like water, sugar, etc.

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Classification of Mixtures

- **Mixtures** can be categorized into two types:
  - Heterogeneous mixtures
  - Homogeneous mixtures
- This categorization of mixture depends on how *uniformly* the substances within them mix.

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Heterogeneous Mixture

- A **heterogeneous mixture** is one in which the composition varies from one region of the mixture to another.
    - Made of multiple substances, whose presence can be seen (Example: a salt and sand mixture)
        - Portions of a sample of heterogeneous mixture have different composition and properties.

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Homogeneous Mixture

- A **homogeneous mixture** is one made of multiple substances, but appears to be one substance.

- All portions of a sample have the same composition and properties (like sweetened tea).

- Homogeneous mixtures have uniform compositions because the atoms or molecules that compose them mix uniformly.

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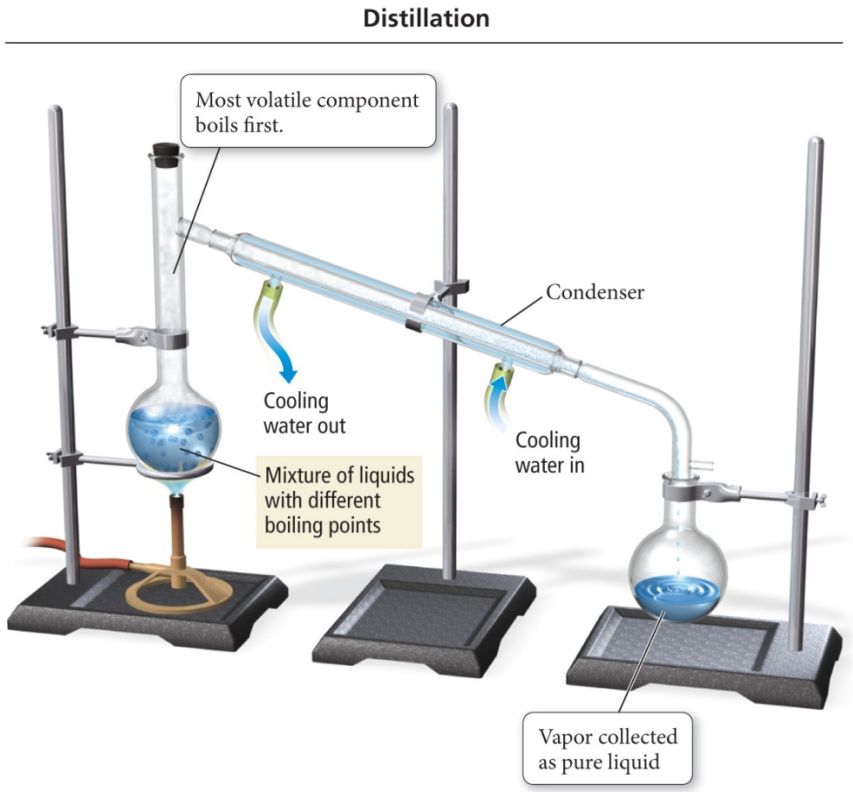
Separating Mixtures

- Mixtures are separable because the different components have different physical or chemical properties.
- Various techniques that exploit these differences are used to achieve separation.
- A mixture of sand and water can be separated by **decanting**—carefully pouring off the water into another container.

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Separating Mixtures

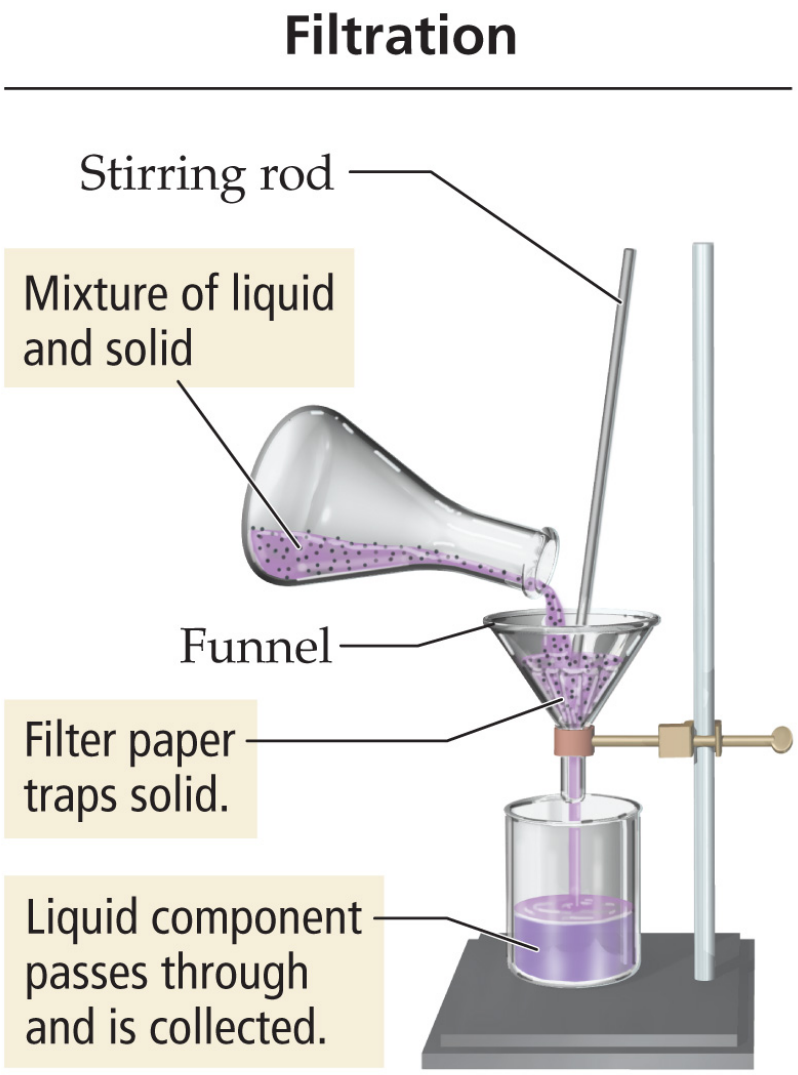
- A homogeneous mixture of liquids can usually be separated by **distillation**, a process in which the mixture is heated to boil off the more **volatile** (easily vaporizable) liquid. The volatile liquid is then re-condensed in a condenser and collected in a separate flask.



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Separating Mixtures

- A mixture of an insoluble solid and a liquid can be separated by **filtration**—process in which the mixture is poured through filter paper in a funnel.



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Physical and Chemical Changes

Physical Change:

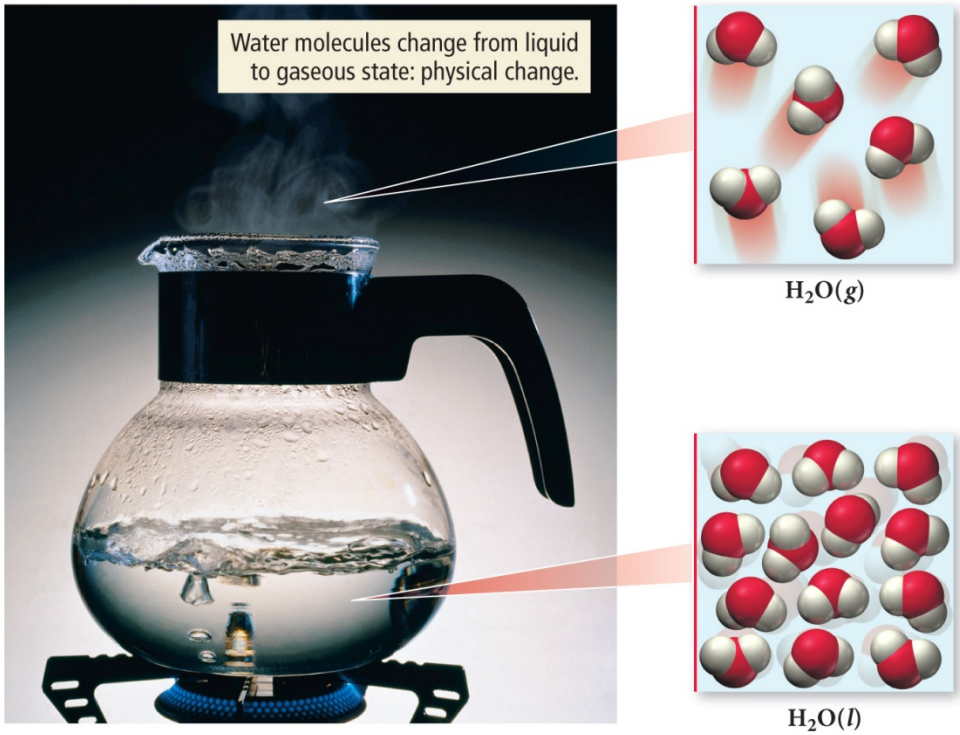
- Changes that alter only the state or appearance, but not composition, are **physical changes**.

- The atoms or molecules that compose a substance *do not change* their identity during a physical change.

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Physical Change

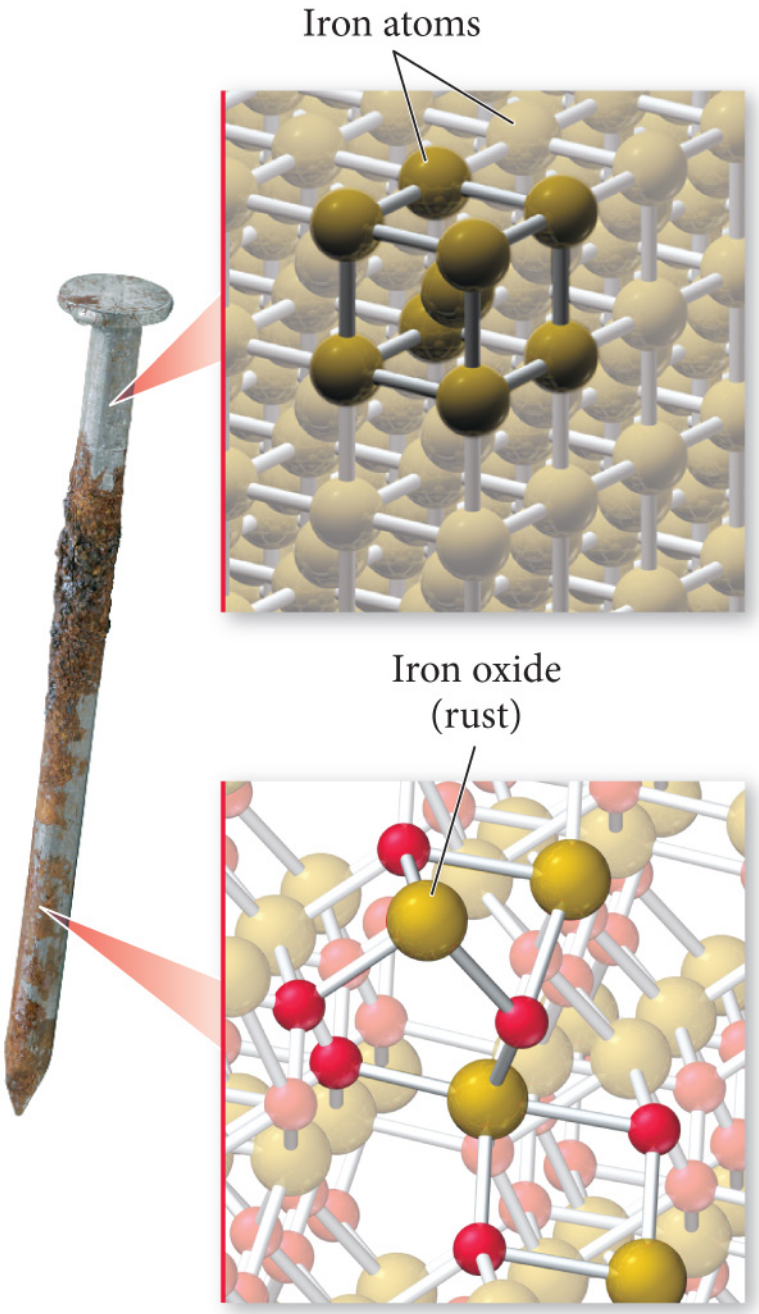
- When water boils, it changes its state from a liquid to a gas.
- The gas remains composed of water molecules, so this is a physical change.



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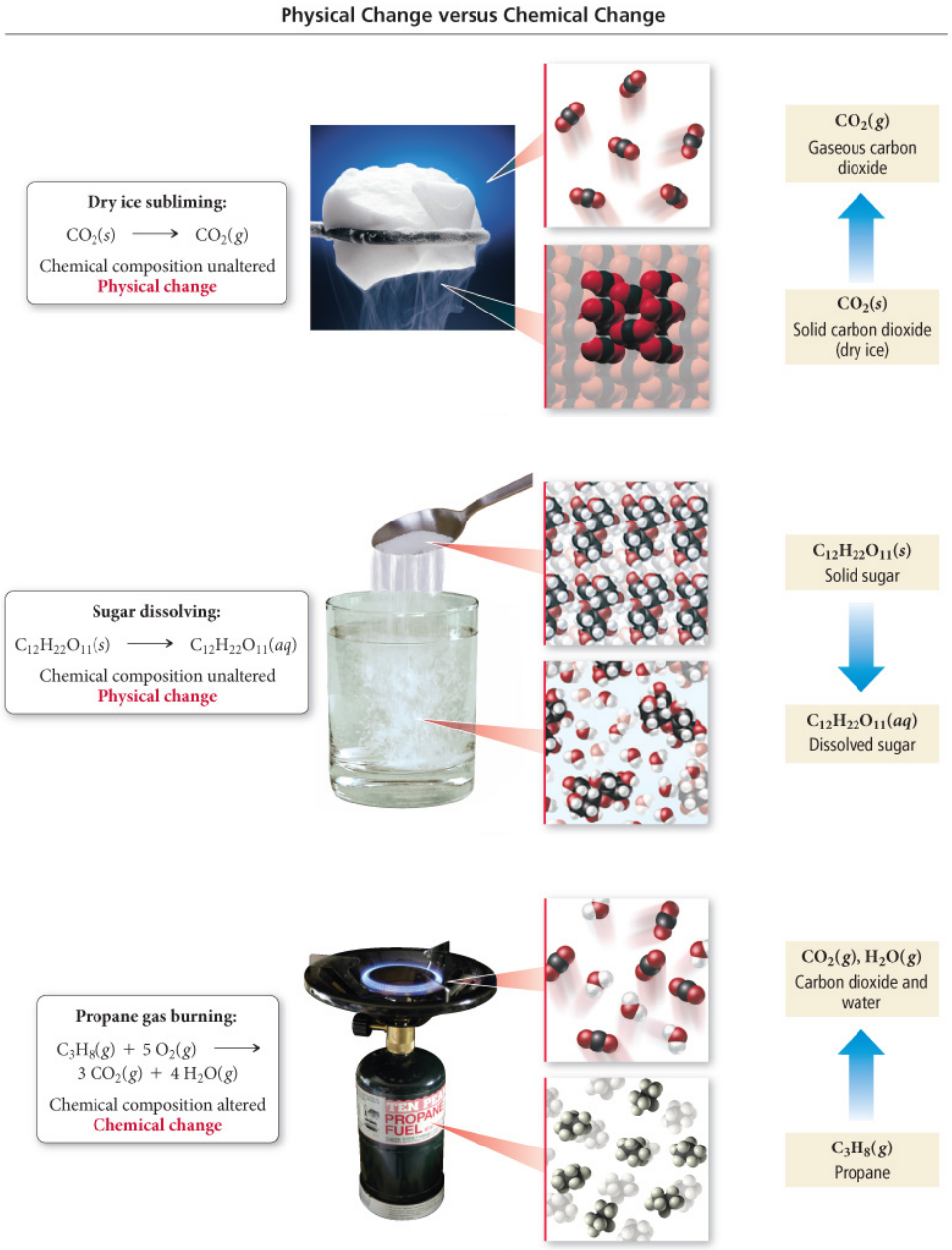
Chemical Change

- Changes that alter the composition of matter are **chemical changes**.
- During a chemical change, atoms rearrange, transforming the original substances into different substances.
- Rusting of iron is a chemical change.



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Physical and Chemical Changes



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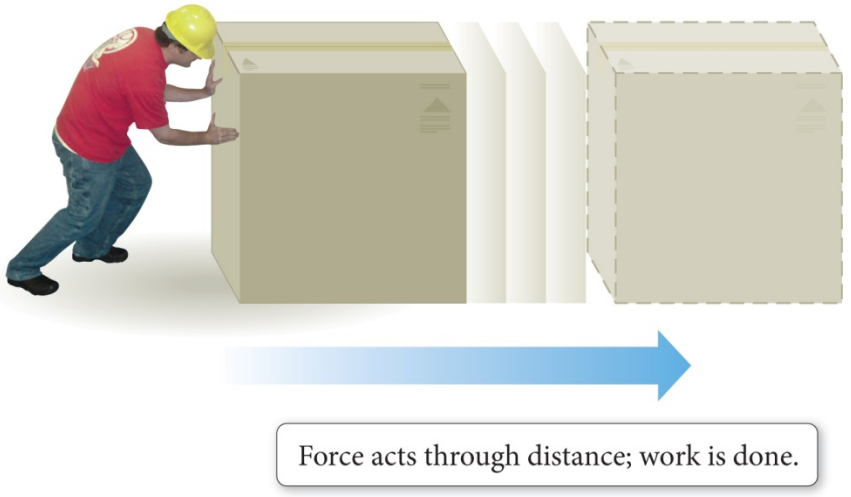
Physical and Chemical Properties

- A **physical property** is a property that a substance displays without changing its composition.
  - The smell of gasoline is a physical property.
  - Odor, taste, color, appearance, melting point, boiling point, and density are all physical properties.
- A **chemical property** is a property that a substance displays only by changing its composition via a chemical change (or chemical reaction).
  - The flammability of gasoline, in contrast, is a chemical property.
  - Chemical properties include corrosiveness, acidity, and toxicity.

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Energy: A Fundamental Part of Physical and Chemical Change

- **Energy** is the *capacity to do work.*
- **Work** is defined as the action of a force through a distance.



- When you push a box across the floor or pedal your bicycle across the street, you have done work.

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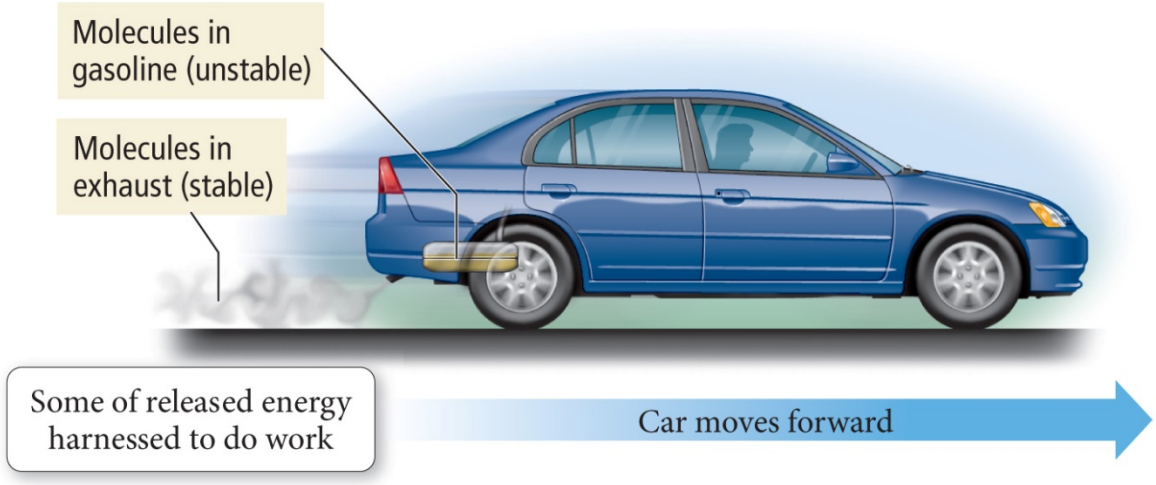
Energy

- **Kinetic energy** is the energy associated with the motion of an object.
- **Potential energy** is the energy associated with the position or composition of an object.
- **Thermal energy** is the energy associated with the temperature of an object.
  - Thermal energy is actually a type of kinetic energy because it arises from the motion of the individual atoms or molecules that make up an object.



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Summarizing Energy



- Energy is always conserved in a physical or chemical change; it is neither created nor destroyed (law of conservation of energy).

- Systems with high potential energy tend to change in a direction that lowers their potential energy, releasing energy into the surroundings.

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The Units of Measurement

- In chemistry, **units**—standard quantities used to specify measurements—are critical.
- **The two most common unit systems are as follows:**
  - **Metric system**, used in most of the world
  - **English system**, used in the United States
- Scientists use the **International System of Units (SI)**, which is based on the metric system.
  - The abbreviation *SI* comes from the French, phrase *Système International d'Unités.*

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The Standard Units

TABLE 1.1 SI Base Units

Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

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The Meter: A Measure of Length

- The **meter (m)** is slightly longer than a yard (1 yard is 36 inches, while 1 meter is 39.37 inches).



Yardstick



Meterstick

- 1 meter = 1/10,000,000 of the distance from the equator to the North Pole (through Paris).
  - The International Bureau of Weights and Measures now defines it more precisely as the distance light travels through a vacuum in a certain period of time, 1/299,792,458 second.

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The Kilogram: A Measure of Mass

- The **mass** of an object is a measure of the quantity of matter within it.
- The SI unit of mass = kilogram (kg).
  - 1 kg = 2 lb 3 oz
- A second common unit of mass is the gram (g).
  - One gram is 1/1000 kg.
- **Weight** of an object is a measure of the *gravitational pull* on its matter.



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The Second: A Measure of Time

- Measure of the duration of an event
- SI units = second (s)
- 1 s is defined as the period of time it takes for a specific number of radiation events of a specific transition from cesium-133.

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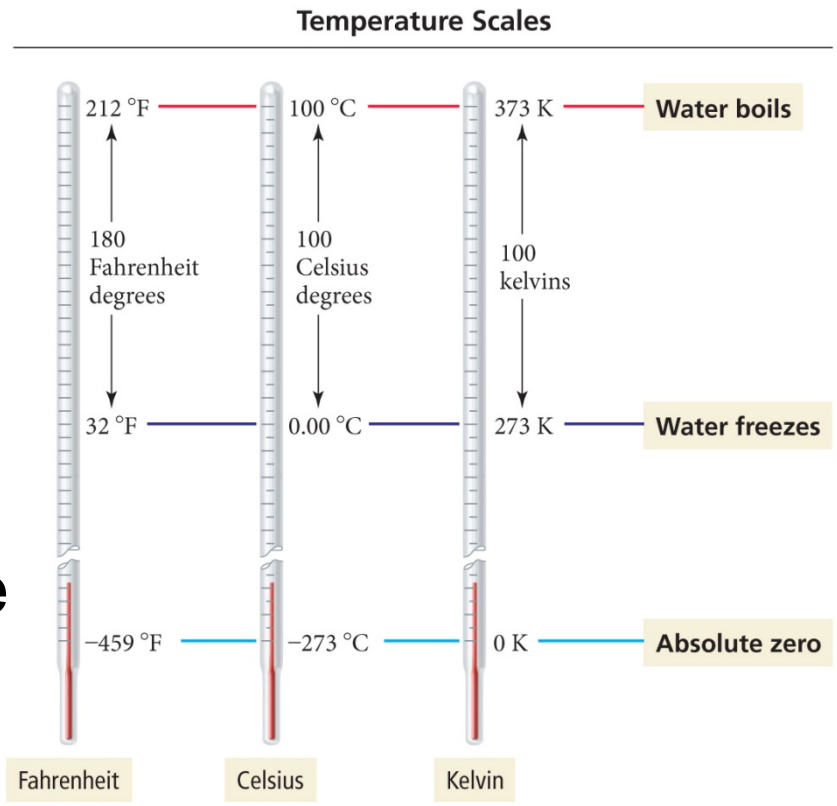
The Kelvin: A Measure of Temperature

- The **Kelvin (K)** is the SI unit of **temperature**.
- The temperature is a measure of the average amount of kinetic energy of the atoms or molecules that compose the matter.
- Temperature also determines the direction of thermal energy transfer, or what we commonly call heat.
- Thermal energy transfers from hot to cold objects.

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The Kelvin: A Measure of Temperature

- **Kelvin scale** (*absolute scale*) assigns 0 K (absolute zero) to the coldest temperature possible.
- Absolute zero (–273 °C or –459 °F) is the temperature at which molecular motion virtually stops. Lower temperatures do not exist.



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A Measure of Temperature

- The Fahrenheit degree is five-ninths the size of a Celsius degree.
- The Celsius degree and the Kelvin degree are the same size.
- Temperature scale conversion is done with these formulas:

$$°C = \frac{(°F - 32)}{1.8}$$

$$K = °C + 273.15$$

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Prefix Multipliers

- The International System of Units uses the **prefix multipliers** shown in Table 1.2 with the standard units.
- These multipliers change the value of the unit by the powers of 10 (just like an exponent does in scientific notation).
- For example, the kilometer has the prefix *kilo* meaning 1000 or 10^3.

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Counting Significant Figures

- Significant figures deal with writing numbers to reflect precision.
- The precision of a measurement depends on the instrument used to make the measurement.
- The preservation of this precision during calculations can be accomplished by using *significant figures.*

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Counting Significant Figures

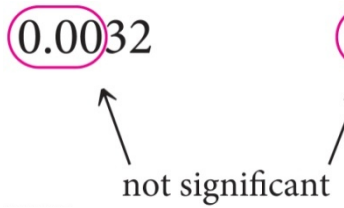
- *The greater the number of significant figures, the greater the certainty of the measurement.*

- **To determine the number of significant figures in a number, follow these rules (examples are on the right).**

Significant Figure Rules		Examples
1. All nonzero digits are significant	28.03	0.0540
2. Interior zeroes (zeroes between two nonzero digits) are significant.	408	7.0301

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Counting Significant Figures

Significant Figure Rules		Examples
3. Leading zeroes (zeroes to the left of the first nonzero digit) are not significant. They only serve to locate the decimal point.		
4. Trailing zeroes (zeroes at the end of a number) are categorized as follows:	45.000	3.5600
▪ Trailing zeroes after a decimal point are always significant.		
▪ Trailing zeroes before a decimal point (and after a nonzero number) are always significant.	140.00	2500.55
▪ Trailing zeroes before an *implied* decimal point are ambiguous and should be avoided by using scientific notation.	1200 1.2×10^3 1.20×10^3 1.200×10^3	Ambiguous 2 significant figures 3 significant figures 4 significant figures
▪ Decimal points are placed after one or more trailing zeroes if the zeroes are to be considered significant.	1200.	4 significant figures

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Exact Numbers

- *Exact numbers have an unlimited number of significant figures.*

- Exact counting of discrete objects
- Integral numbers that are part of an equation
- Defined quantities

- *Some conversion factors are defined quantities, while others are not.*

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Significant Figures in Calculations

- In calculations using measured quantities, the results of the calculation must reflect the precision of the measured quantities.

- We should not lose or gain precision during mathematical operations.

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Significant Figure: Rules for Calculations

Multiplication and Division Rule:

- In multiplication or division, the result carries the same number of significant figures as the factor with the fewest significant figures.

$$\underset{\text{(4 sig. figures)}}{1.052} \times \underset{\text{(5 sig. figures)}}{12.504} \times \underset{\text{(2 sig. figures)}}{0.53} = 6.7208 = \underset{\text{(2 sig. figures)}}{6.7}$$

$$\underset{\text{(5 sig. figures)}}{2.0035} \div \underset{\text{(3 sig. figures)}}{3.20} = 0.626094 = \underset{\text{(3 sig. figures)}}{0.626}$$

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Rules for Calculations

Addition and Subtraction Rule:

- In addition or subtraction the result carries the same number of decimal places as the quantity with the fewest decimal places.

$$\begin{array}{r} 2.345 \\ 0.07 \\ 2.9975 \\ \hline 5.4125 \end{array} = 5.41 \qquad \begin{array}{r} 5.9 \\ -0.221 \\ \hline 5.679 \end{array} = 5.7$$

It is helpful to draw a line next to the number with the fewest decimal places. This line determines the number of decimal places in the answer.

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Rules for Calculations

Rules for Rounding:

- When rounding to the correct number of significant figures,
    - round down if the last (or leftmost) digit dropped is four or less;
    - round up if the last (or leftmost) digit dropped is five or more.

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Rules for Rounding

- Round to two significant figures:

  5.37 rounds to 5.4

  5.34 rounds to 5.3

  5.35 rounds to 5.4

  5.349 rounds to 5.3

- Notice in the last example that only the *last (or leftmost) digit being dropped* determines in which direction to round—ignore all digits to the right of it.

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Rounding in Multistep Calculations

- To avoid rounding errors in multistep calculations round only the final answer.

- Do not round intermediate steps. If you write down intermediate answers, keep track of significant figures by underlining the least significant digit.

$$
\begin{aligned}
6.78 \times 5.903 \times (5.489 - 5.01) \\
&= 6.78 \times 5.903 \times 0.\underline{4}79 \\
&= 19.1707 \\
&= 19
\end{aligned}
$$

underline least significant digit

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Precision and Accuracy

- **Accuracy** refers to how close the measured value is to the actual value.

- **Precision** refers to how close a series of measurements are to one another or how reproducible they are.

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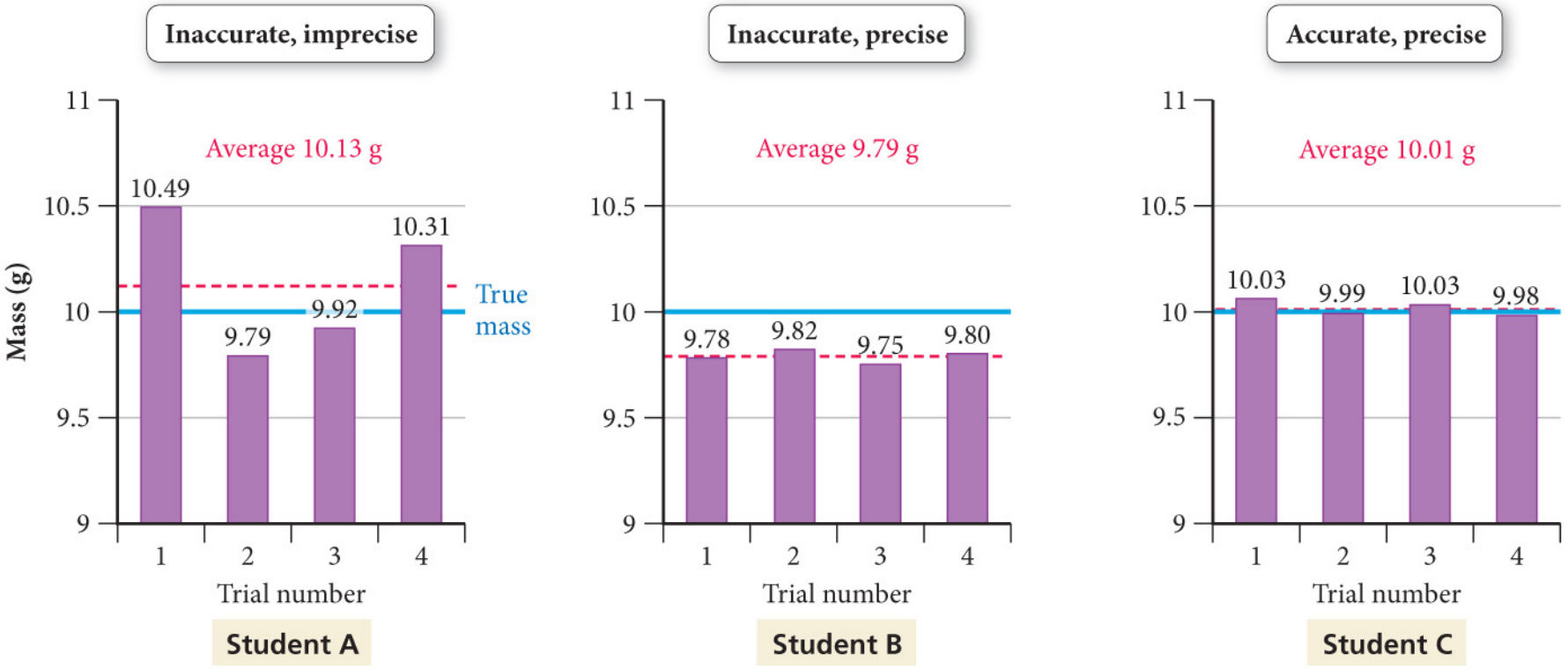
Precision and Accuracy

- Consider the results of three students who repeatedly weighed a lead block known to have a true mass of 10.00 g (indicated by the solid horizontal blue line on the graphs).

	Student A	Student B	Student C
Trial 1	10.49 g	9.78 g	10.03 g
Trial 2	9.79 g	9.82 g	9.99 g
Trial 3	9.92 g	9.75 g	10.03 g
Trial 4	10.31 g	9.80 g	9.98 g
Average	10.13 g	9.79 g	10.01 g

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Precision and Accuracy



- Measurements are said to be
  - precise if they are consistent with one another.
  - accurate only if they are close to the actual value.

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Precision and Accuracy

- The results of student A are both inaccurate (not close to the true value) and imprecise (not consistent with one another).
  - ***Random error*** *is an error that has the equal probability of being too high or too low.*
- The results of student B are precise (close to one another in value), but inaccurate.
  - ***Systematic error*** *is an error that tends toward being either too high or too low*.
- The results of student C display little systematic error or random error—they are both accurate and precise.

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Solving Chemical Problems

- Most chemistry problems you will solve in this course are **unit conversion problems.**

- Using units as a guide to solving problems is called **dimensional analysis.**

- Units should always be included in calculations; they are multiplied, divided, and canceled like any other algebraic quantity.

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Dimensional Analysis

- A unit equation is a statement of two equivalent quantities, such as

2.54 cm = 1 in.

- A **conversion factor** is a fractional quantity of a unit equation with the units we are ***converting from*** on the bottom and the units we are ***converting to*** on the top.

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Dimensional Analysis

- Most unit conversion problems take the following form:

$$\text{Information given} \times \text{conversion factor(s)} = \text{information sought}$$

$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

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Dimensional Analysis

Units Raised to a Power:

- When building conversion factors for units raised to a power, remember to raise both the number and the unit to the power. For example, to convert from in^2 to cm^2, we construct the conversion factor as follows:

$$2.54\,\text{cm} = 1\ \text{in}$$

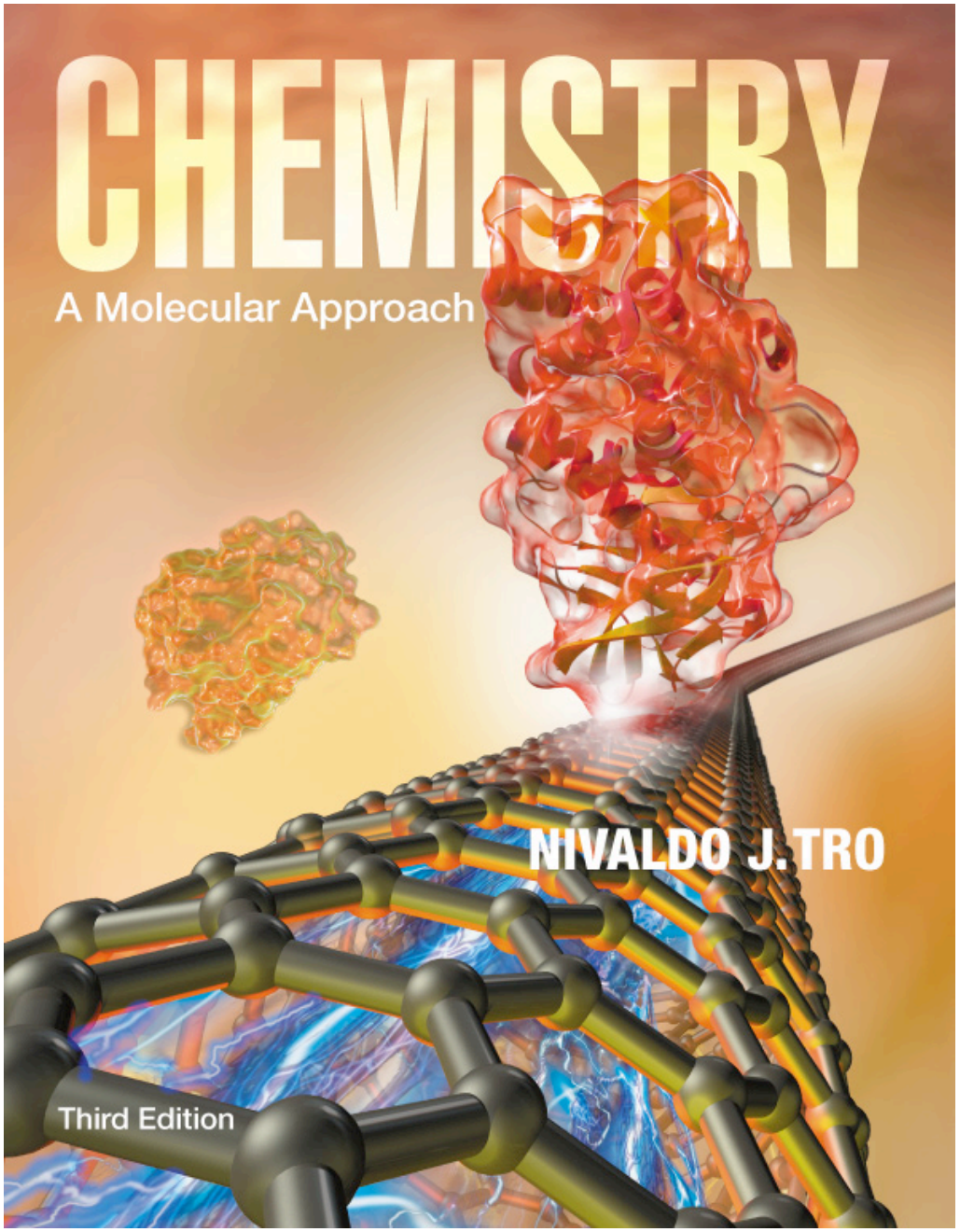
$$(2.54\,\text{cm})^2 = (1\ \text{in})^2$$

$$(2.54)^2\,\text{cm}^2 = 1^2\,\text{in}^2$$

$$6.45\,\text{cm}^2 = 1\ \text{in}^2$$

$$\frac{6.45\,\text{cm}^2}{1\ \text{in}^2} = 1$$

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Lecture Presentation

Chapter 2

Atoms and Elements

Christian Madu, Ph.D.
Collin College

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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.

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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.

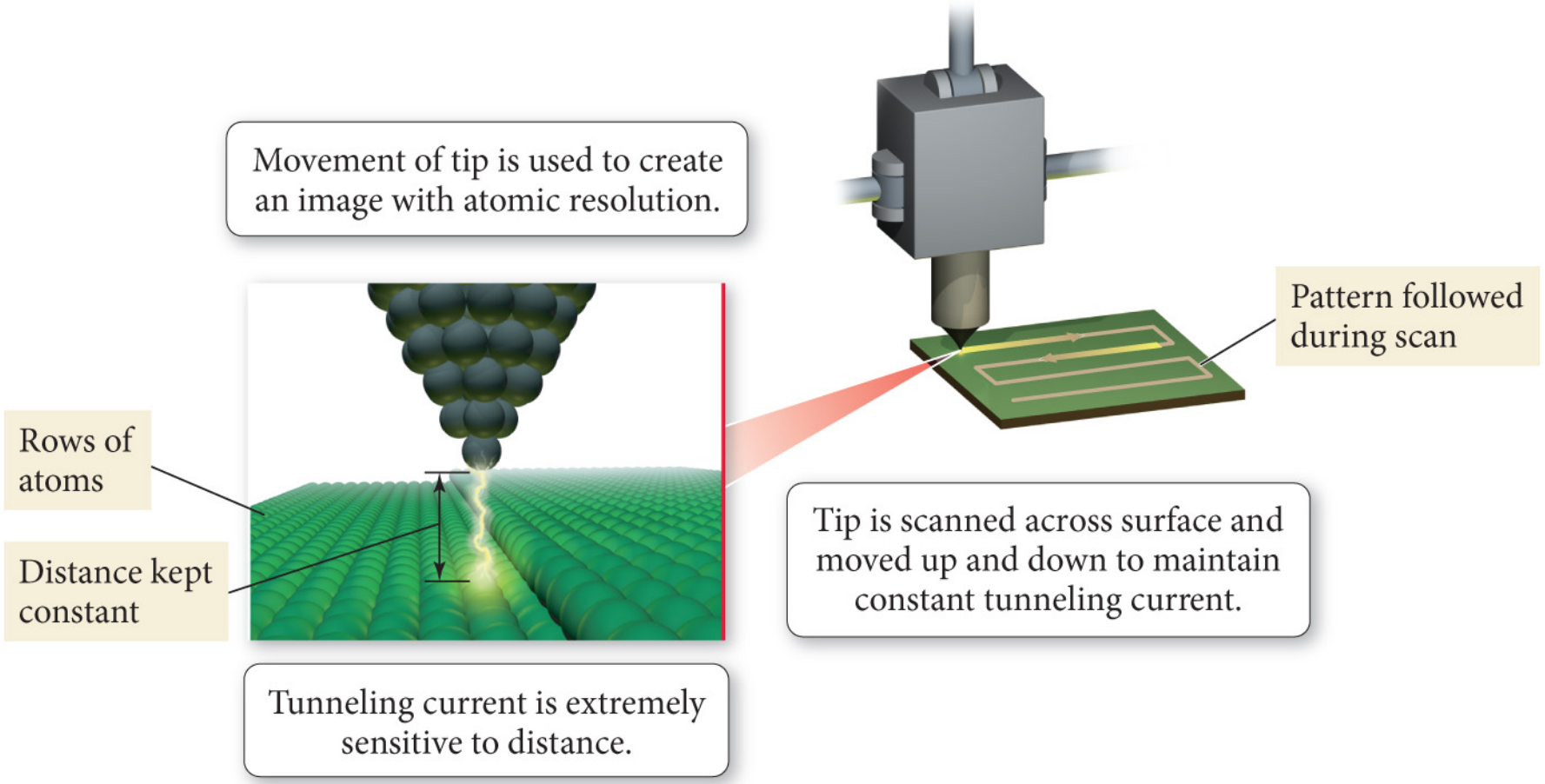
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Imaging and Moving Individual Atoms

- 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.

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Scanning Tunneling Microscopy



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

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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).

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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.

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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.

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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.

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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

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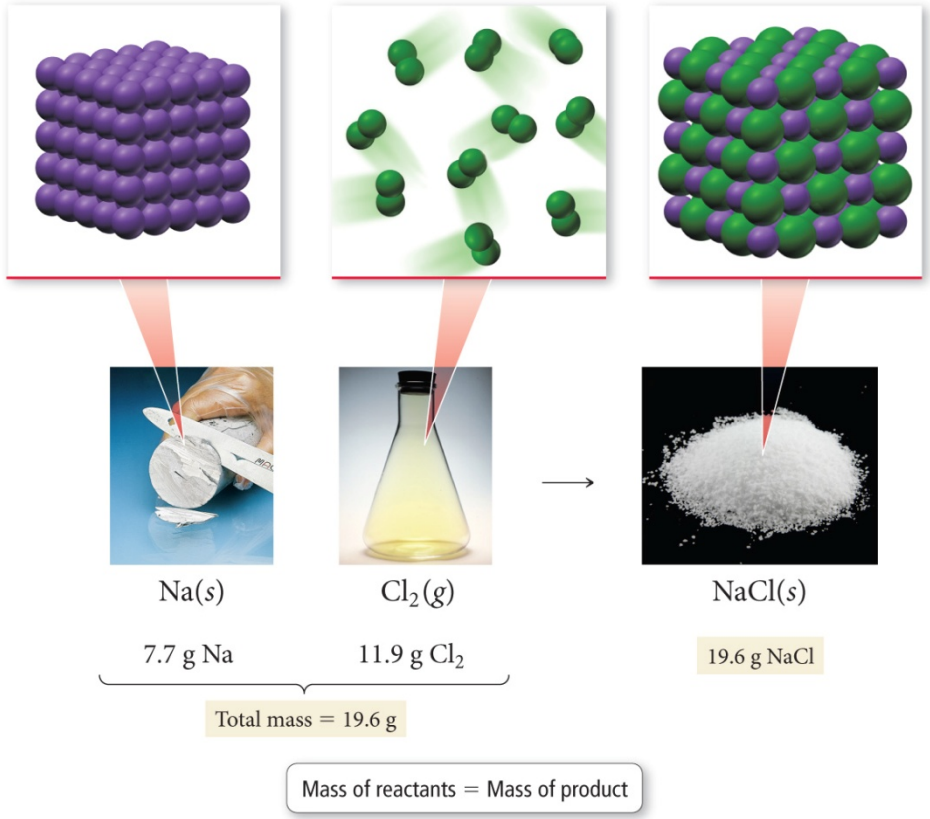
The Law of Conservation of Mass

- 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.

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The Law of Conservation of Mass



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

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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.***

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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:

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

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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.).

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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.

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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.

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

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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.

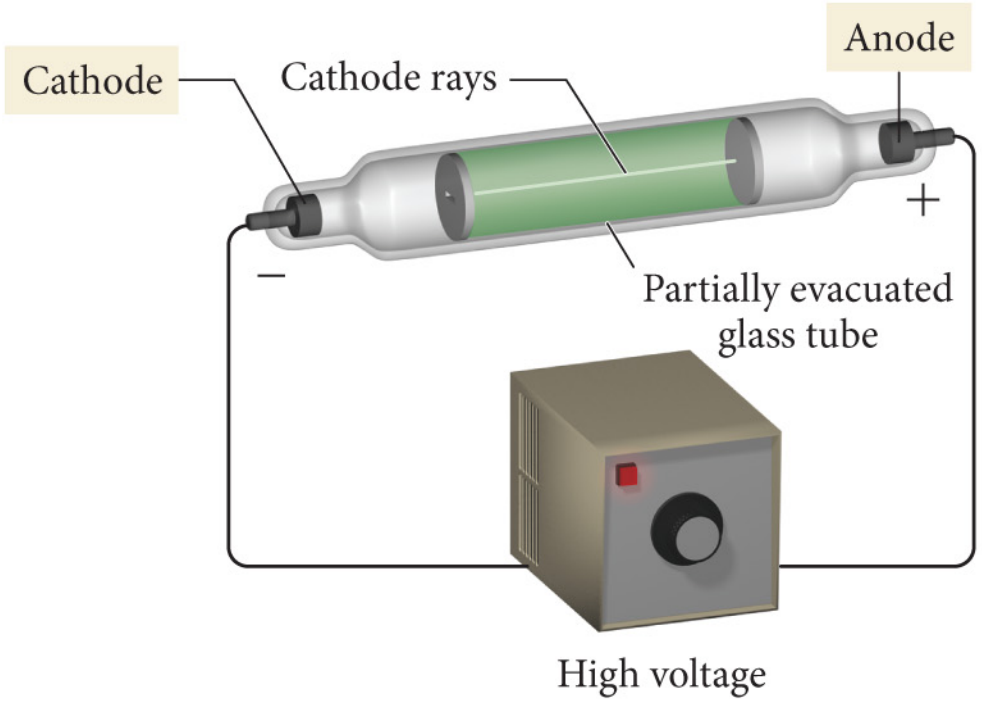
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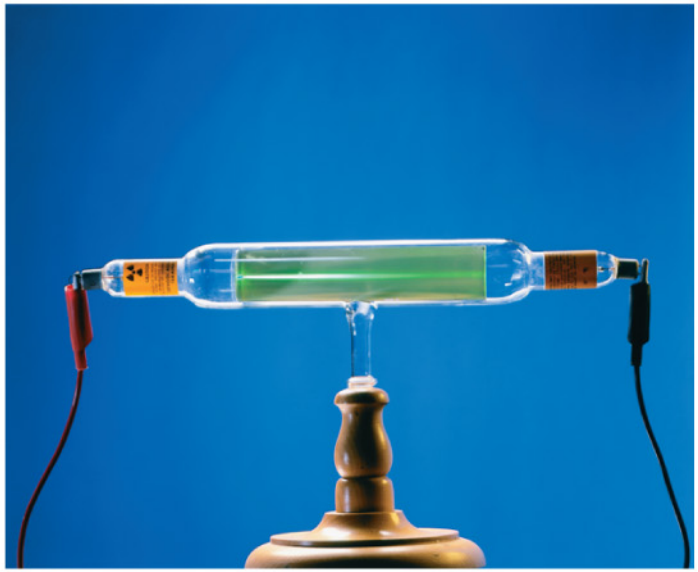
The Discovery of the Electron

- 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).

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The Discovery of the Electron





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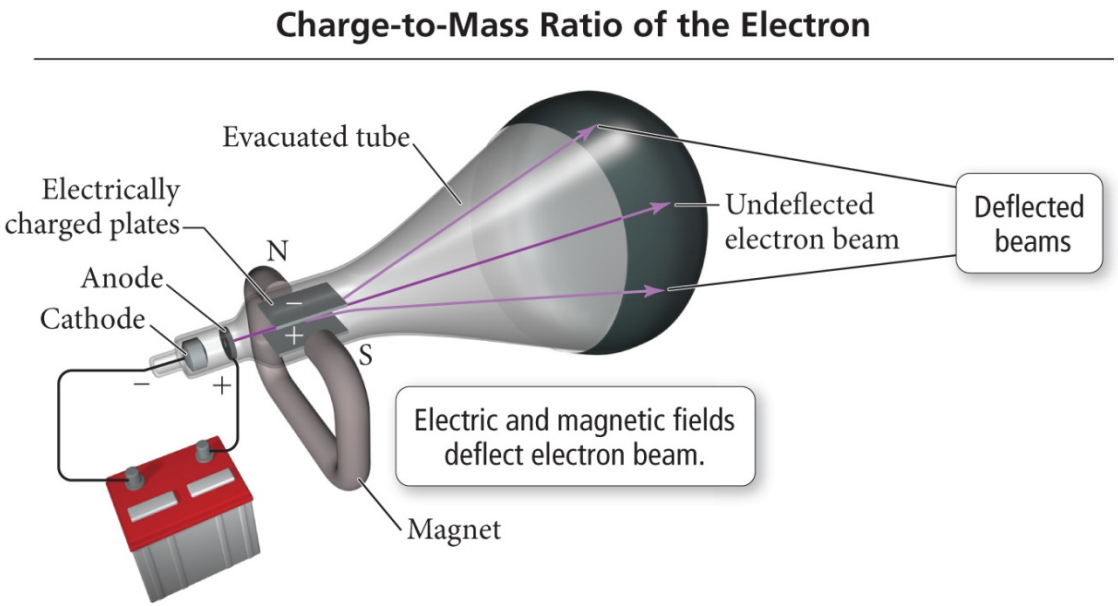
The Discovery of the Electron

- 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**.

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The Discovery of the Electron

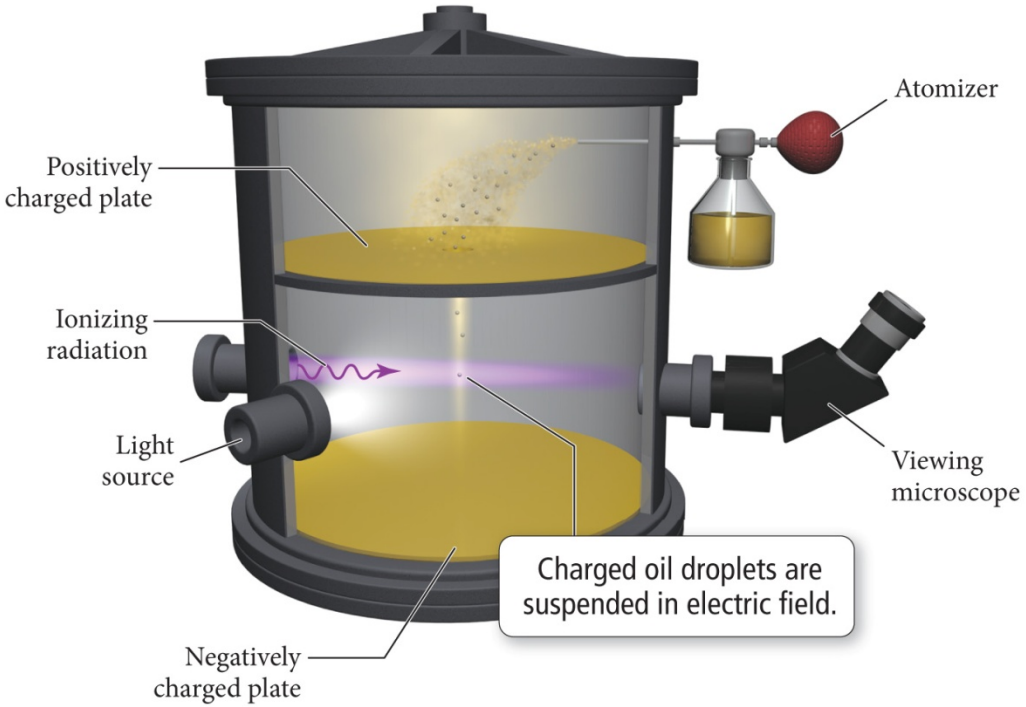
- J. J. Thomson measured the charge-to-mass ratio of the cathode ray particles by deflecting them using electric and magnetic fields, as shown in the figure.
- The value he measured was -1.76×10^3 coulombs (C) per gram.



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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.



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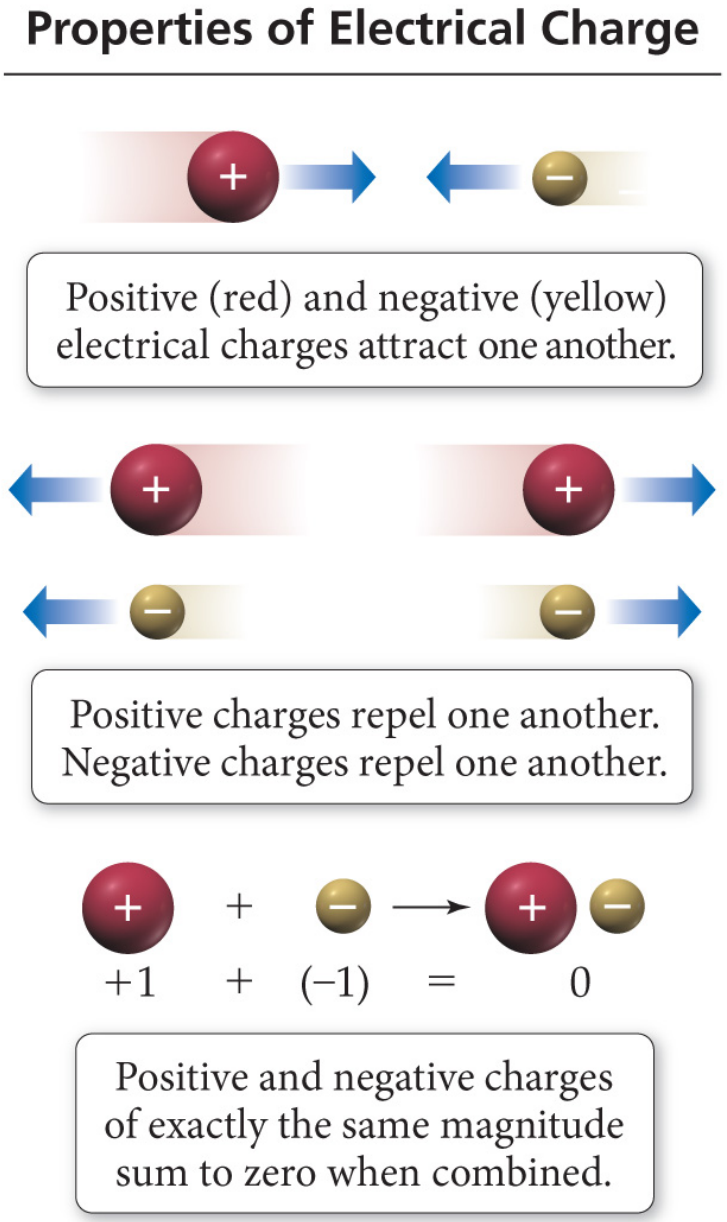
Millikan's Oil Drop Experiment

- 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.

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The Discovery of the Electron

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



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Millikan's Oil Drop Experiment

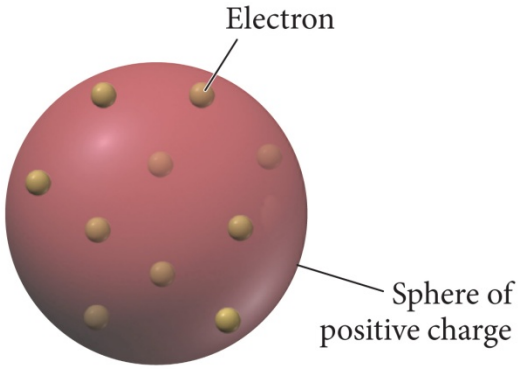
- With this number in hand, and knowing Thomson's mass-to-charge ratio for electrons, we can deduce the mass of an electron:

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

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The Structure of the Atom

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



Plum-pudding model

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

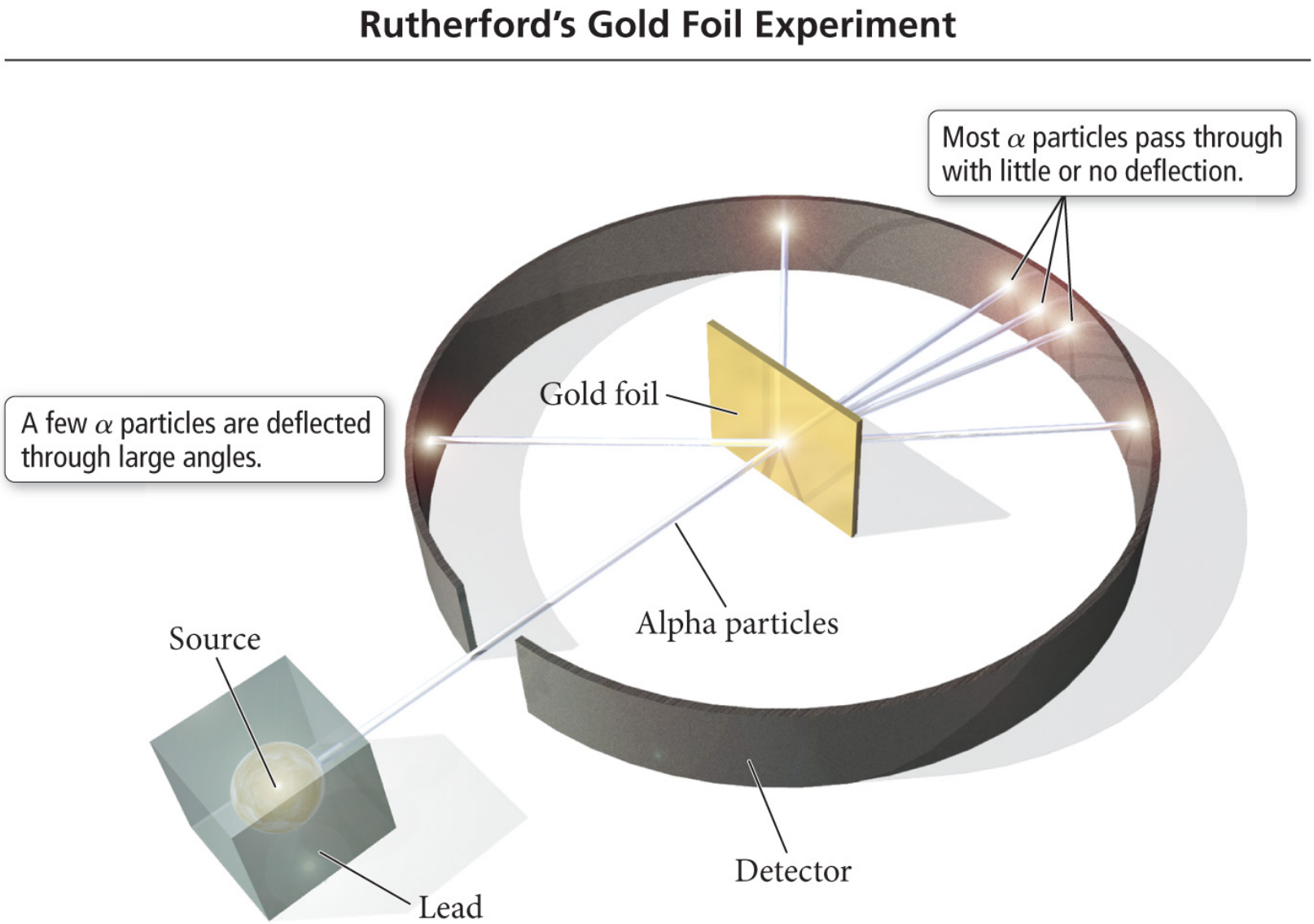
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Rutherford's Gold Foil Experiment

- 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.

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Rutherford's Gold Foil Experiment



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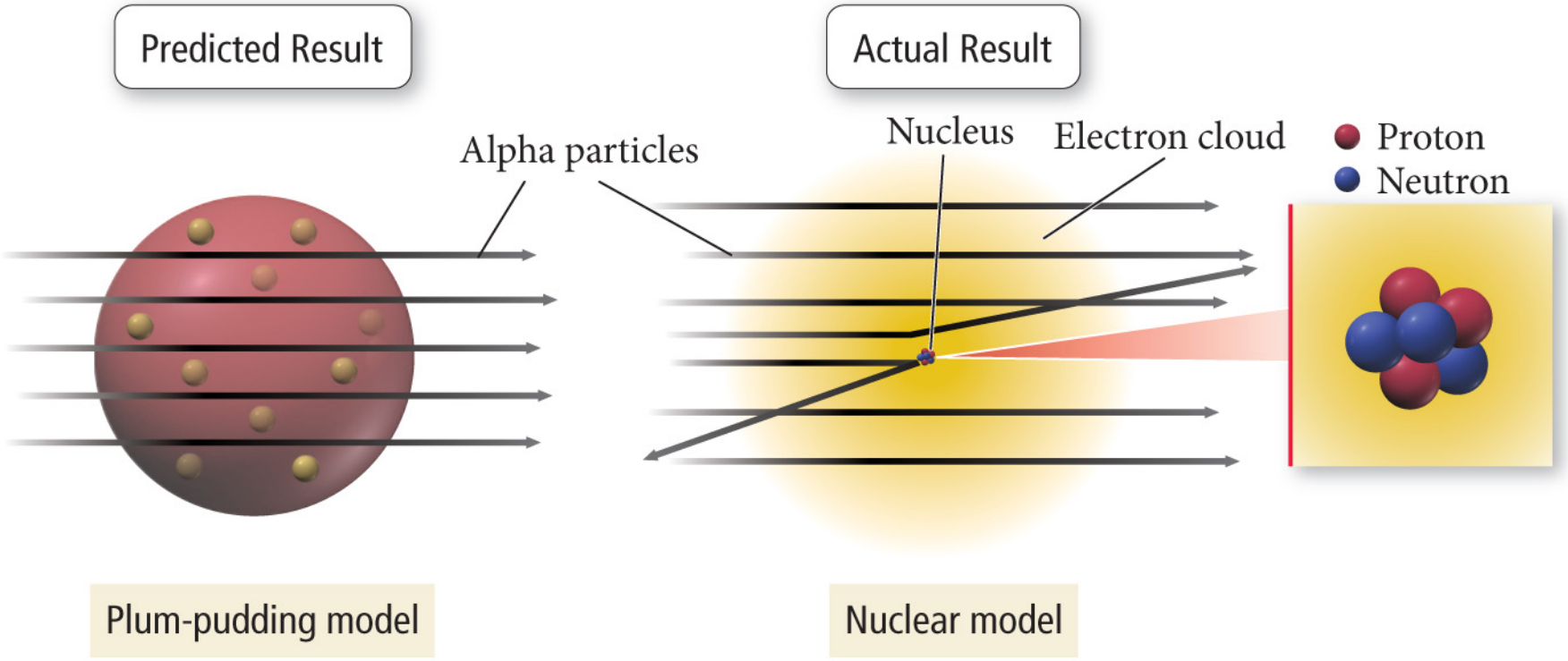
Rutherford's Gold Foil Experiment

- 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.

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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.

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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.

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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.

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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.

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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.

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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.

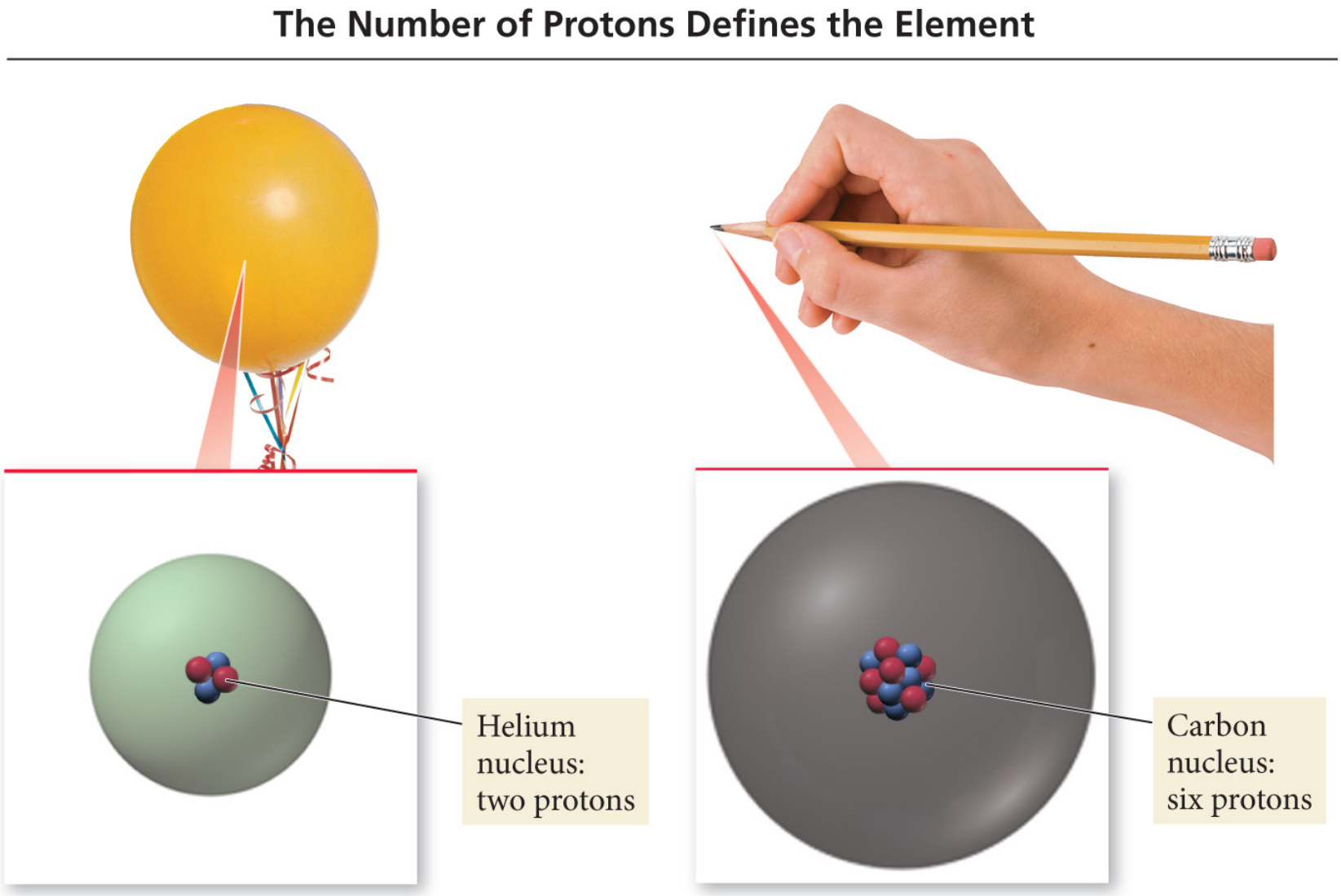
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Elements: Defined by Their Numbers of Protons

- 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***.

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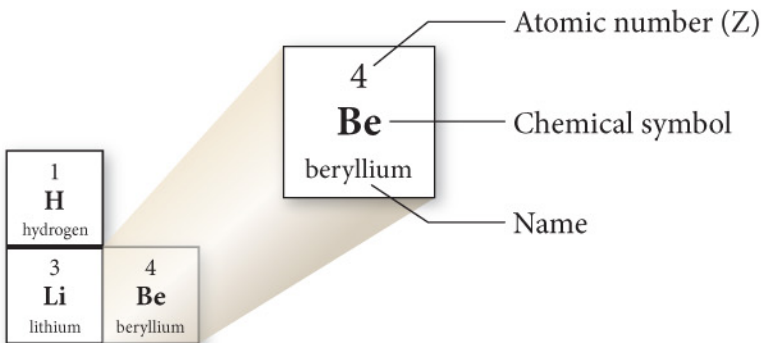
Elements: Defined by Their Numbers of Protons



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Periodic Table

The Periodic Table



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

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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.

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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**.

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Isotopes: Varied Number of Neutrons

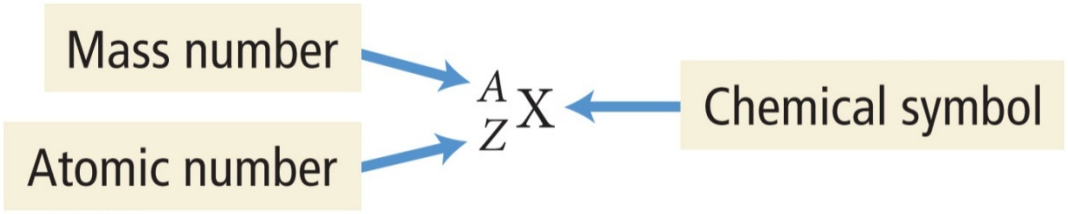
- The relative amount of each different isotope in a naturally occurring sample of a given element is roughly constant.
- These percentages are called the **natural abundance** of the isotopes.
  - Advances in mass spectrometry have allowed accurate measurements that reveal small but significant variations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

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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)



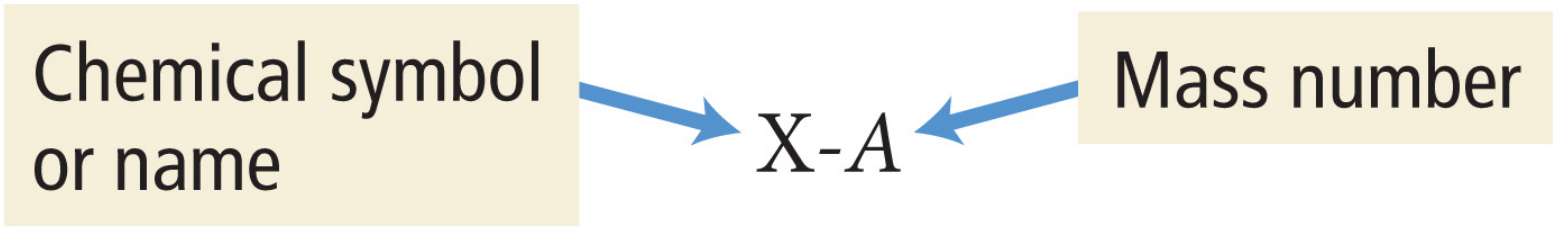
- where X is the chemical symbol, *A* is the mass number, and *Z* is the atomic number.

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

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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.



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

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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

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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**.

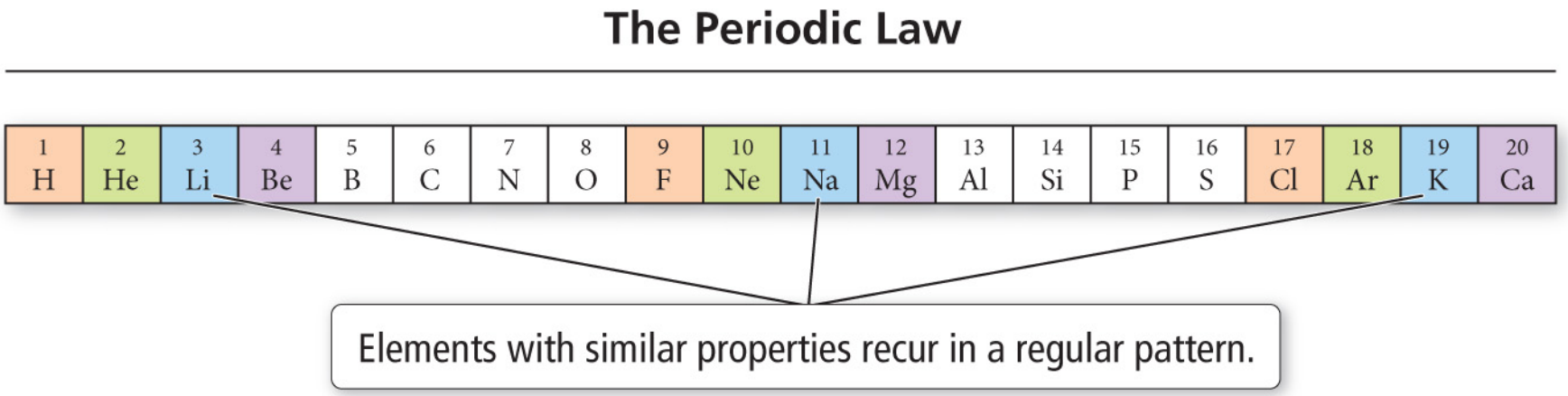
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Finding Patterns: The Periodic Law and the Periodic Table

- 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.

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The Periodic Law

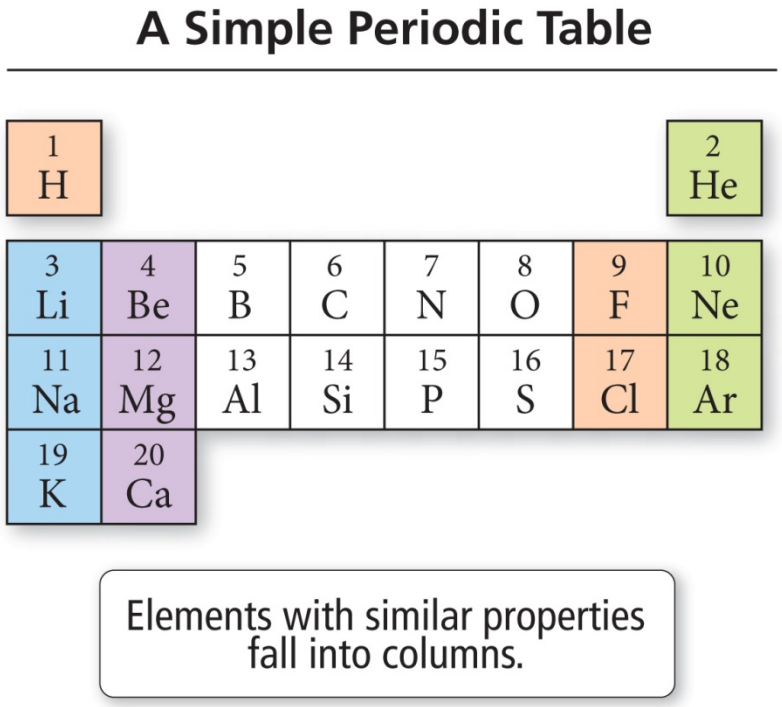


- Mendeleev summarized these observations in the **periodic law**:
  - **When the elements are arranged in order of increasing mass, certain sets of properties recur periodically.**

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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.



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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.

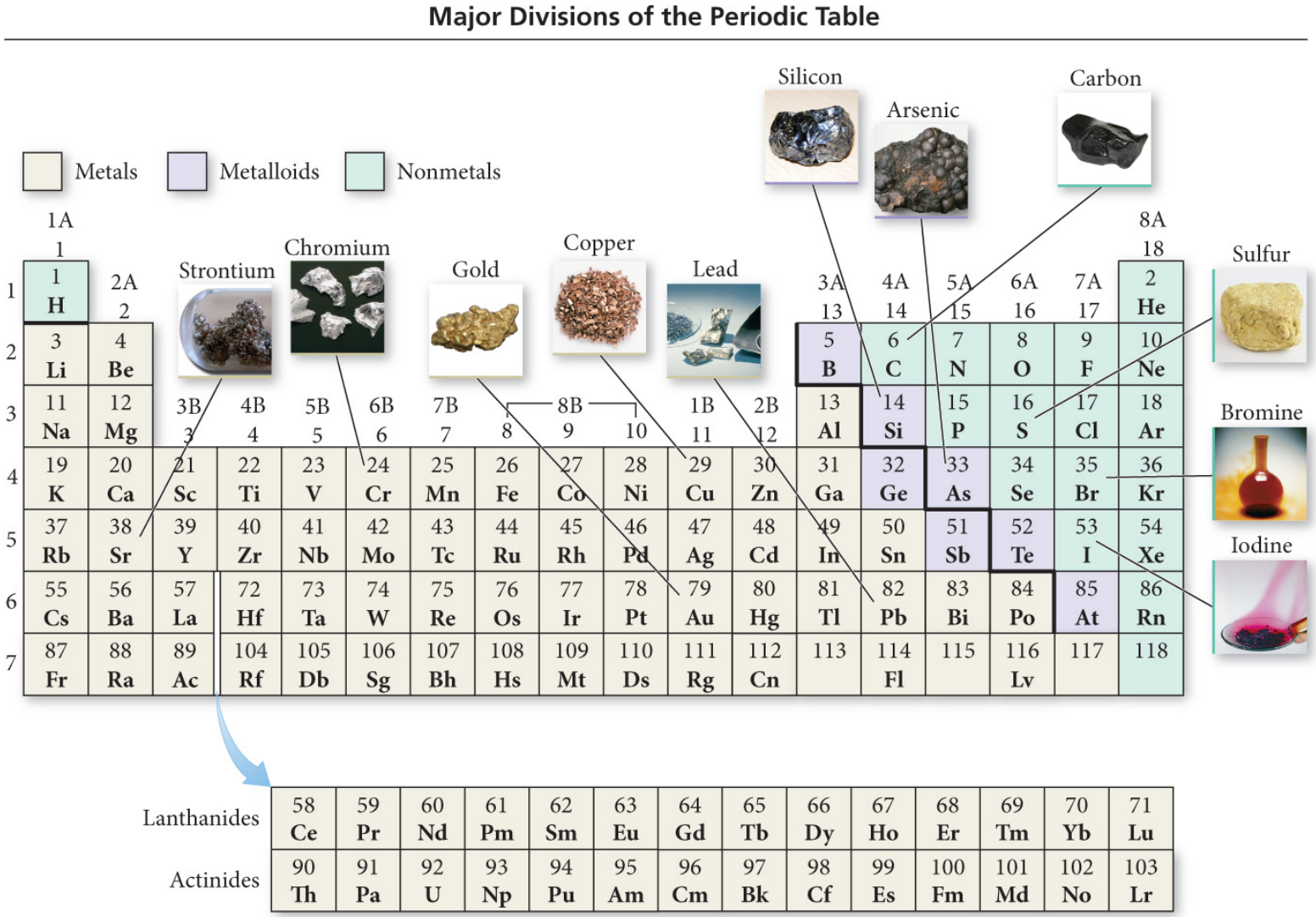
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Modern Periodic Table

- 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.

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Modern Periodic Table



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Classification of Elements

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

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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.

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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)

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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.

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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.

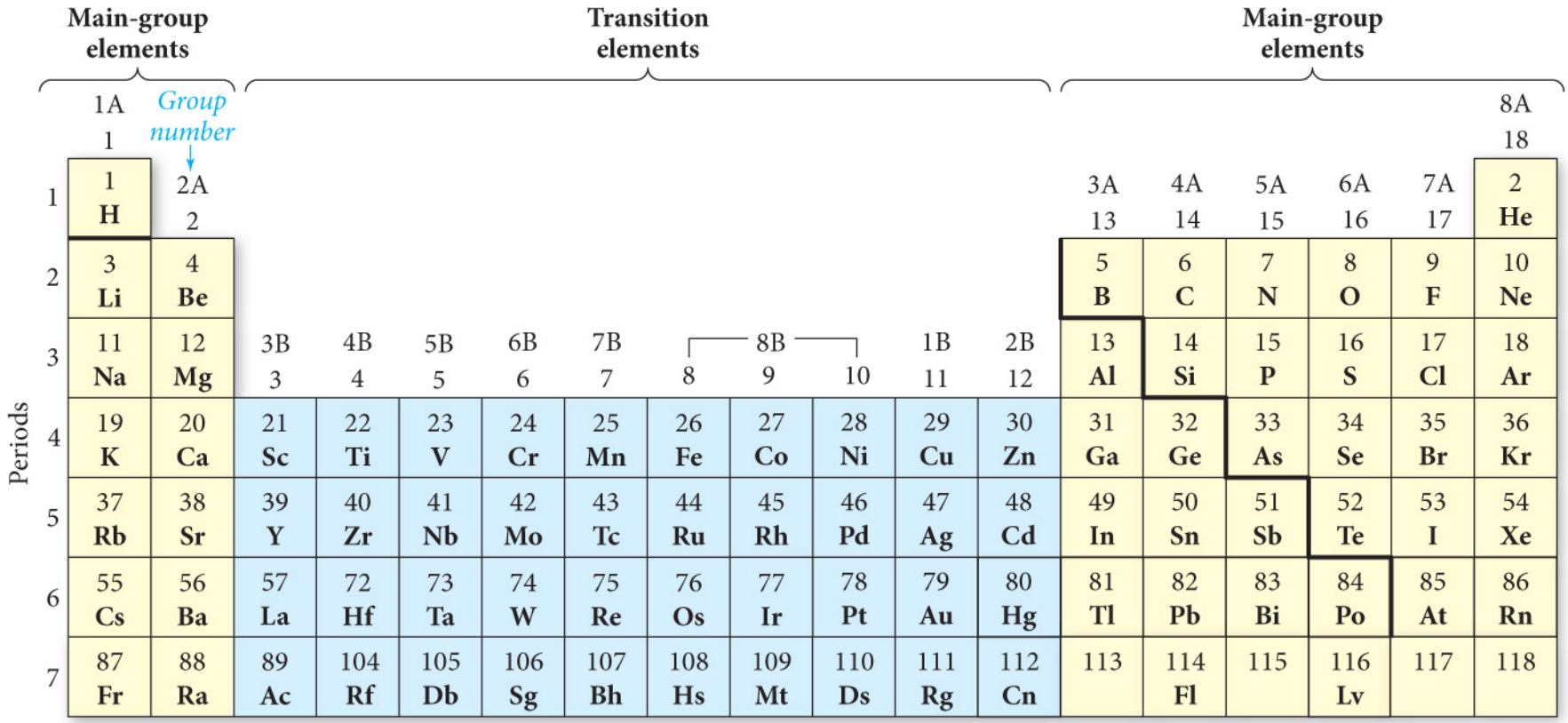
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Periodic Table

- 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.

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Periodic Table



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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).

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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).

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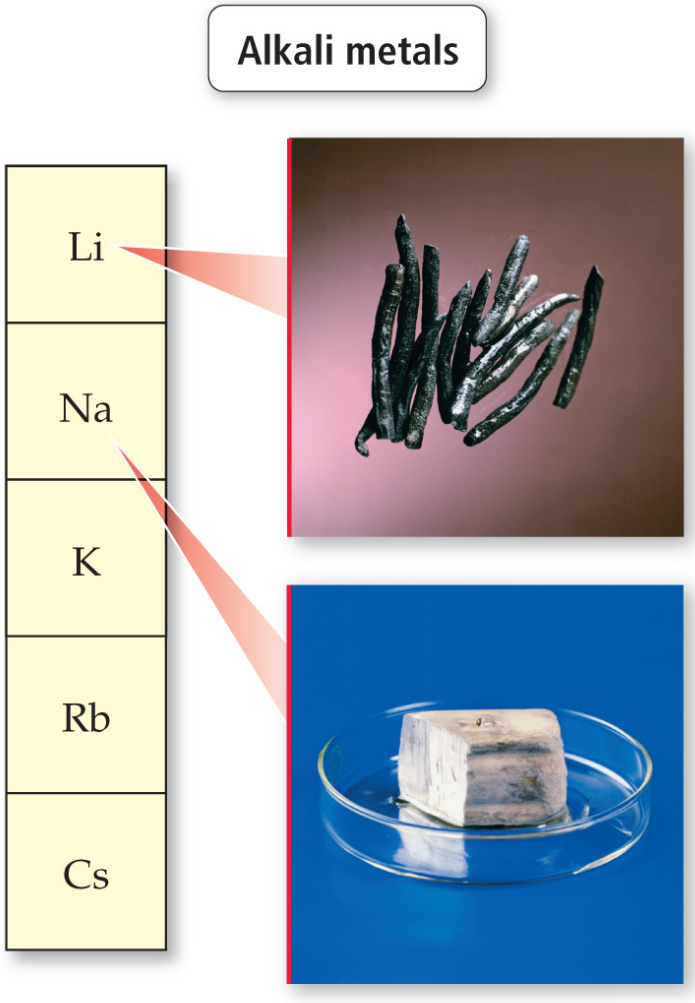
Noble Gas

- 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.

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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.



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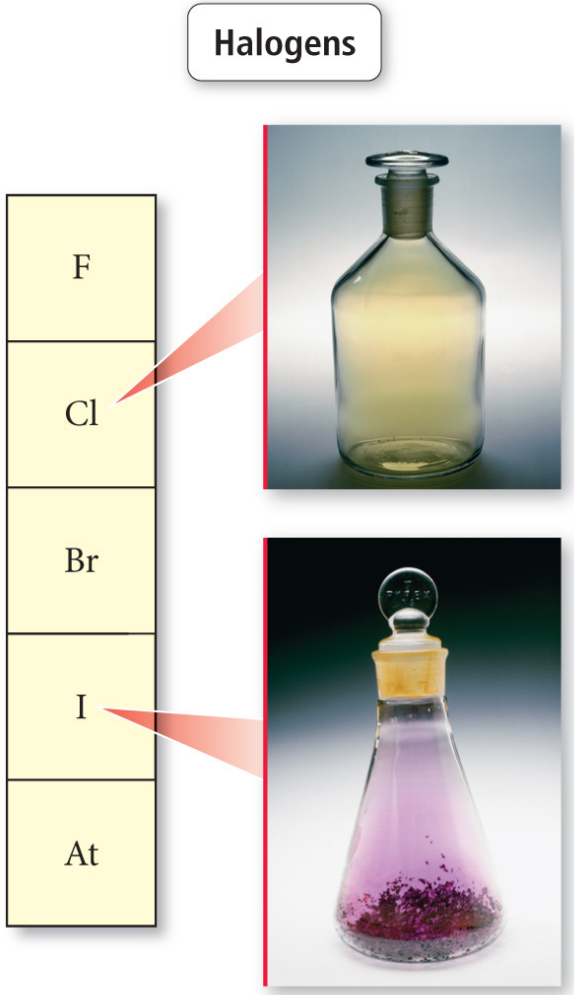
Alkaline Earth Metals

- 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.

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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



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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.**

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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.

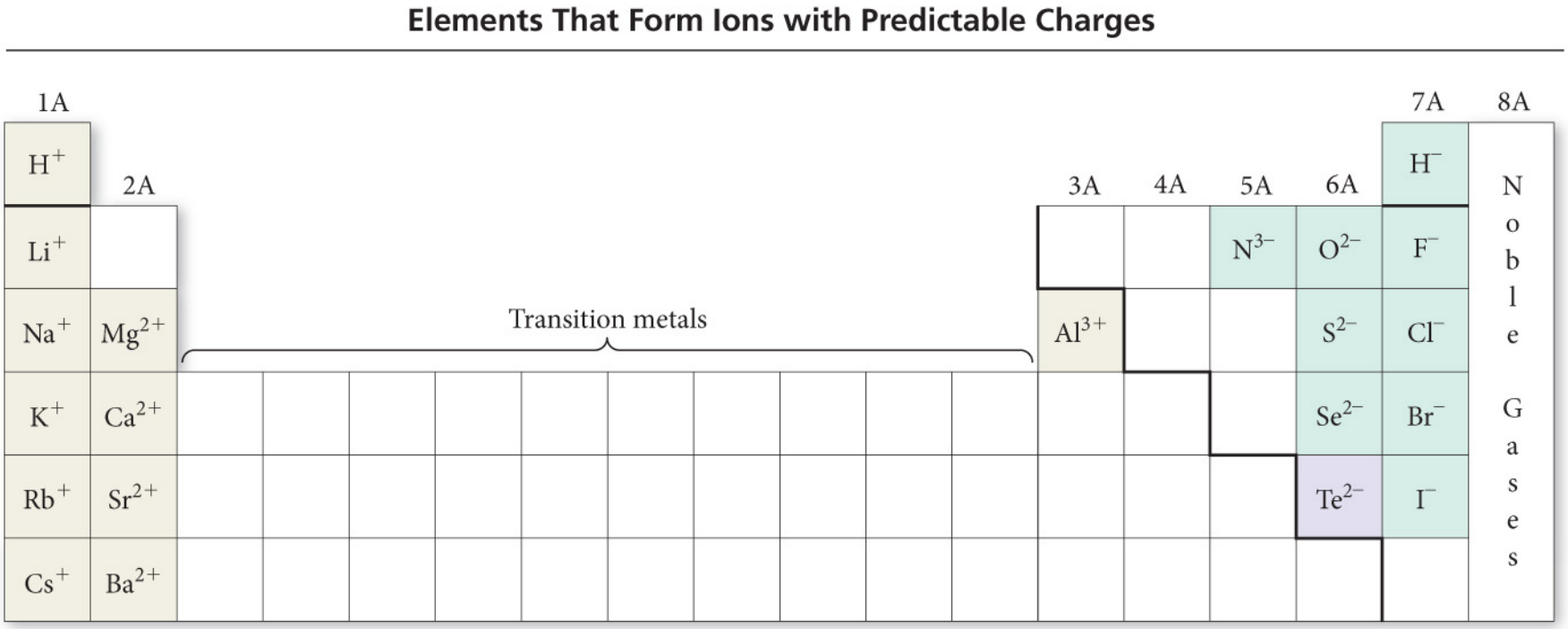
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Ions and the Periodic Table

- 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.

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Ions and the Periodic Table



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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*.

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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:

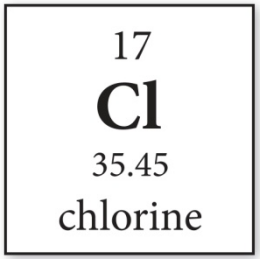
    Cl-37 = 0.2423(36.97 amu) = 8.9578 amu

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

    Atomic Mass Cl = 8.9578 + 26.4968 = 35.45 amu

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Atomic Mass



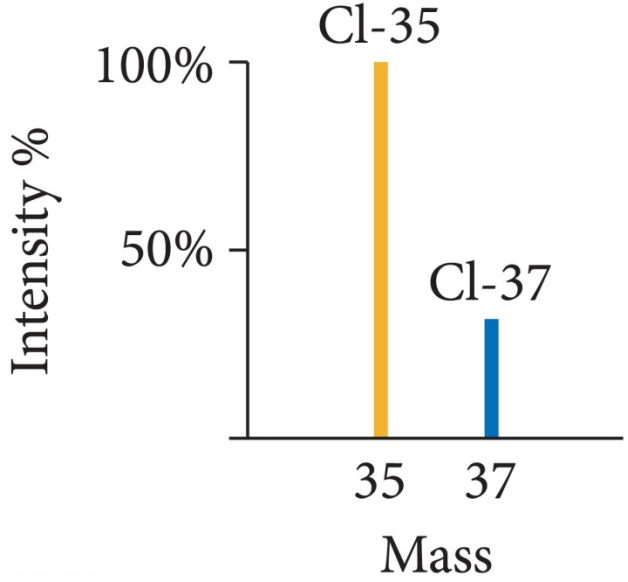
- In general, we calculate the atomic mass with the equation:

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

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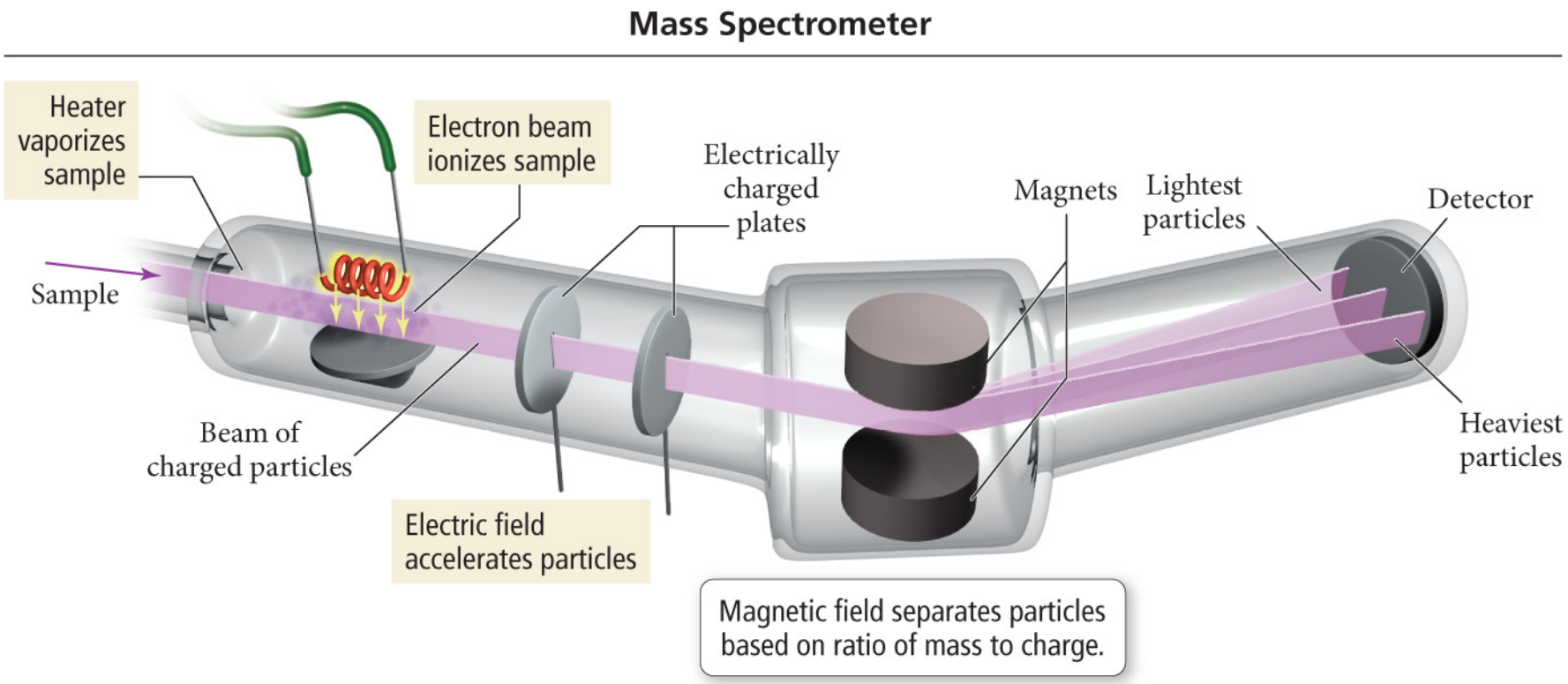
Mass Spectrometry: Measuring the Mass of Atoms and Molecules

- 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.



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Mass Spectrometry



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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.

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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 mole = 6.02214×10^{23} particles
- This number is **Avogadro's number**.

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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.*

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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**

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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:

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

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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).**

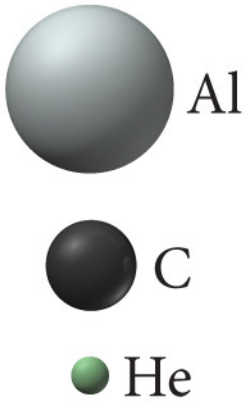
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Converting between Mass and Moles

$26.98\ \text{g aluminum} = 1\ \text{mol aluminum} = 6.022 \times 10^{23}\ \text{Al atoms}$

$12.01\ \text{g carbon} = 1\ \text{mol carbon} = 6.022 \times 10^{23}\ \text{C atoms}$

$4.003\ \text{g helium} = 1\ \text{mol helium} = 6.022 \times 10^{23}\ \text{He atoms}$



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

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Converting between Mass and Moles

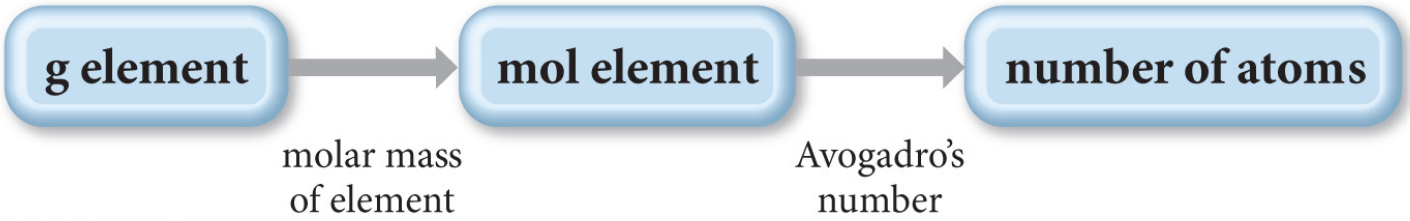
- The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element. For carbon,

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

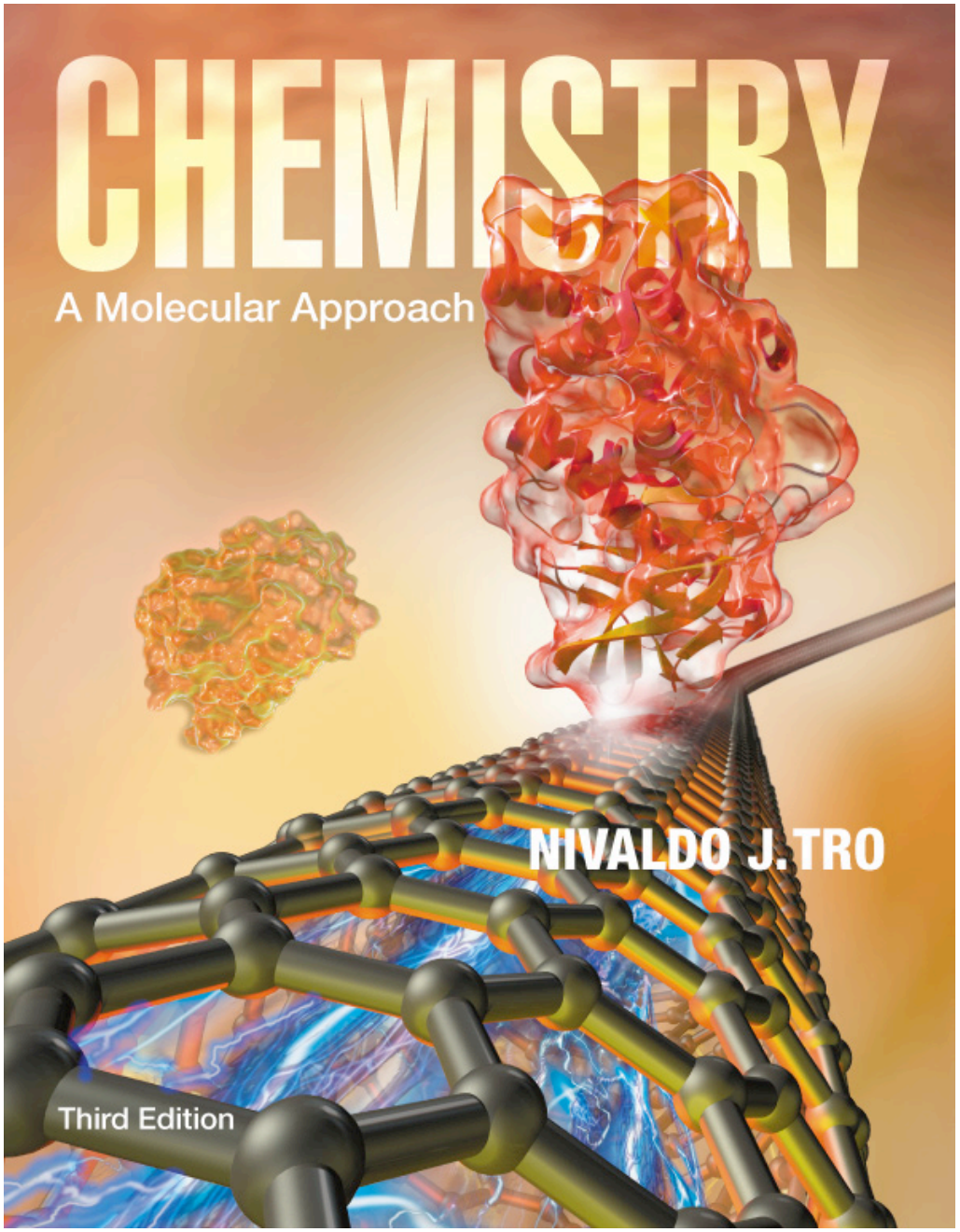
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Conceptual Plan

- 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:



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Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Christian Madu, Ph.D.
Collin College

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How Many Different Substances Exist?

- Elements combine with each other to form *compounds*.
- The great diversity of substances that we find in nature is a direct result of the ability of elements to form compounds.

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Hydrogen, Oxygen, and Water

- The dramatic difference between the elements hydrogen and oxygen and the compound water is typical of the differences between elements and the compounds that they form.

- *When two or more elements combine to form a compound, an entirely new substance results.*

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Hydrogen, Oxygen, and Water

Selected Properties	Hydrogen 	Oxygen 	Water 
Boiling Point	−253 °C	−183 °C	100 °C
State at Room Temperature	Gas	Gas	Liquid
Flammability	Explosive	Necessary for combustion	Used to extinguish flame

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Definite Proportion

- A hydrogen–oxygen mixture can have any proportions of hydrogen and oxygen gas.
- Water, by contrast, is composed of water molecules that always contain two hydrogen atoms to every one oxygen atom.
- Water has a definite proportion of hydrogen to oxygen.

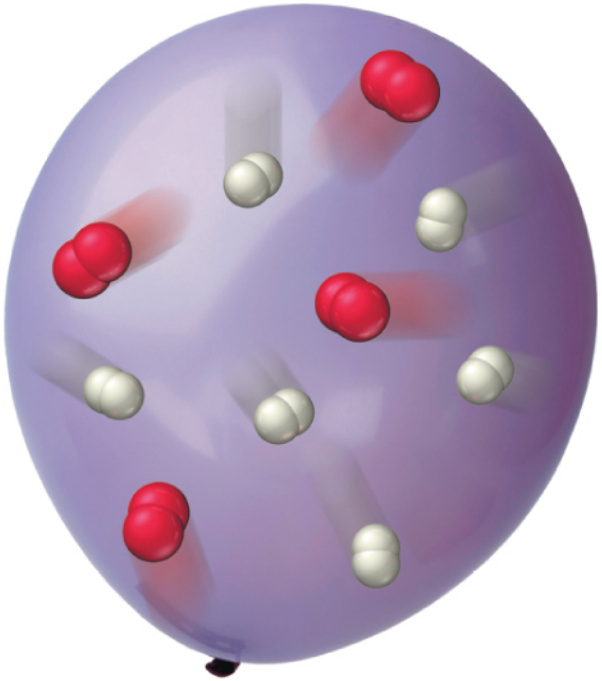
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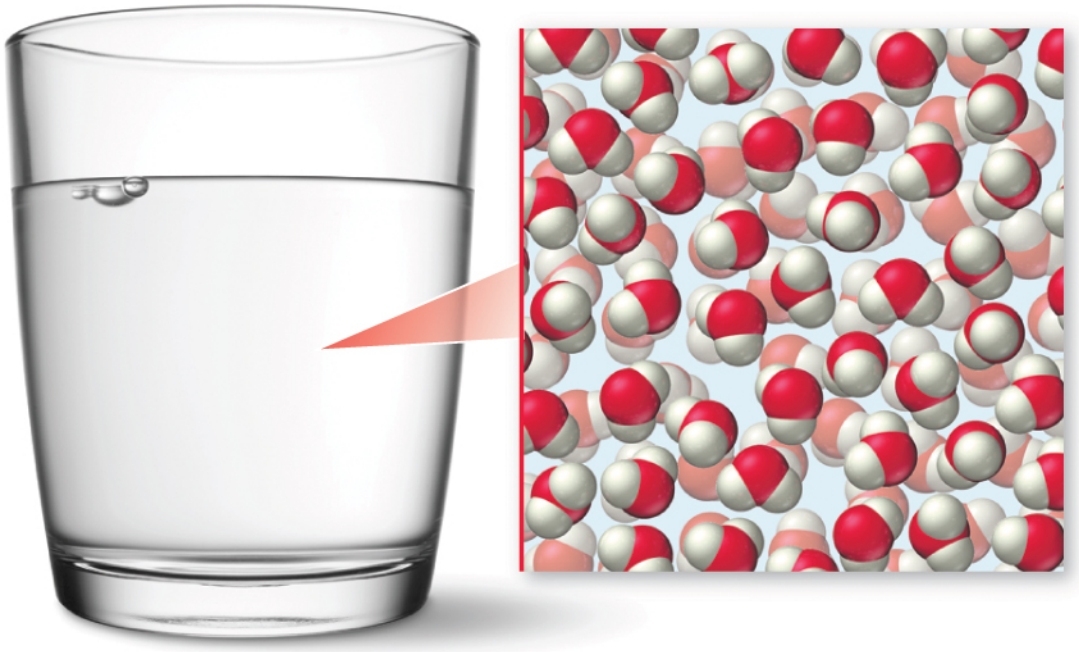
Definite Proportion

Mixtures and Compounds

Hydrogen and Oxygen Mixture
This can have any ratio of hydrogen to oxygen.

Water (A Compound)
Water molecules have a fixed ratio of hydrogen (two atoms) to oxygen (one atom).





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Chemical Bonds

- Compounds are composed of atoms held together by ***chemical bonds***.
- Chemical bonds result from the attractions between the charged particles (the electrons and protons) that compose atoms.
- Chemical bonds are classified into two types:
  - Ionic
  - Covalent

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Ionic Bonds

- *Ionic bonds*—which occur between metals and nonmetals—involve the *transfer* of electrons from one atom to another.
- When a metal interacts with a nonmetal, it can transfer one or more of its electrons to the nonmetal.
  - The metal atom then becomes a *cation*.
  - The nonmetal atom becomes an *anion*.

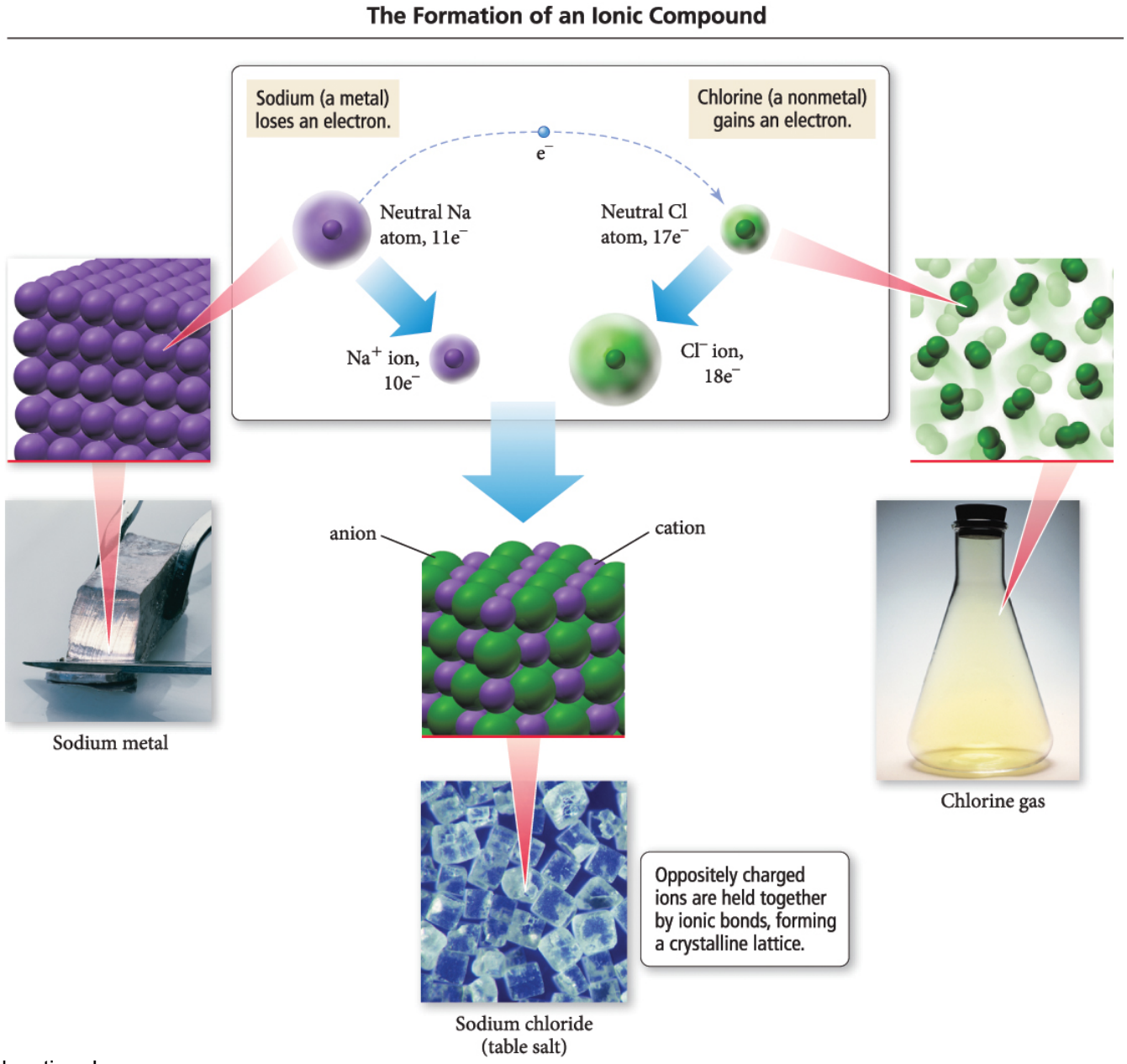
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Ionic Bonds

- These oppositely charged ions attract one another by electrostatic forces and form an **ionic bond**.
- The result is an ionic compound, which in the solid phase is composed of a lattice—a regular three-dimensional array—of alternating cations and anions.

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Ionic Bonds



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Covalent Bonds

- ***Covalent bonds***—which occur between two or more nonmetals—involve the *sharing* of electrons between two atoms.
    - When a nonmetal bonds with another nonmetal, neither atom transfers its electron to the other. Instead the bonding atoms *share* some of their electrons.
- The covalently bound atoms compose a *molecule*.
  – Hence, we call covalently bonded compounds **molecular compounds**.

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Representing Compounds: Chemical Formulas and Molecular Models

- A compound is represented with its **chemical formula.**
- Chemical formula indicates the elements present in the compound and the relative number of atoms or ions of each.
  - Water is represented as H_2O.
  - Carbon dioxide is represented as CO_2.
  - Sodium Chloride is represented as NaCl.
  - Carbon tetrachloride is represented as CCl_4.

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Types of Chemical Formulas

- Chemical formulas can generally be categorized into three different types:
    - Empirical formula
    - Molecular formula
    - Structural formula

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Types of Chemical Formulas

- An **empirical formula** gives the *relative* number of atoms of each element in a compound.
- **A molecular formula** gives the *actual* number of atoms of each element in a molecule of a compound.

(a) For C_4H_8, the greatest common factor is 4. The empirical formula is therefore CH_2.

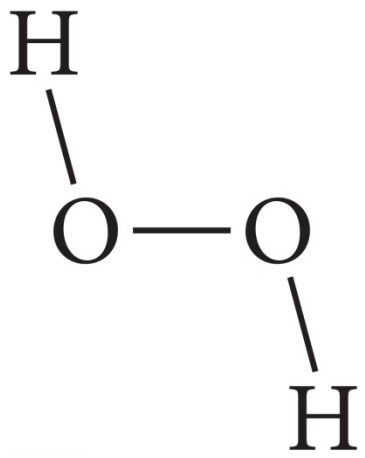
(b) For B_2H_6, the greatest common factor is 2. The empirical formula is therefore BH_3.

(c) For CCl_4, the only common factor is 1, so the empirical formula and the molecular formula are identical.

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Types of Chemical Formulas

- **A structural formula** uses lines to represent covalent bonds and shows how atoms in a molecule are connected or bonded to each other. The structural formula for H_2O_2 is shown below:



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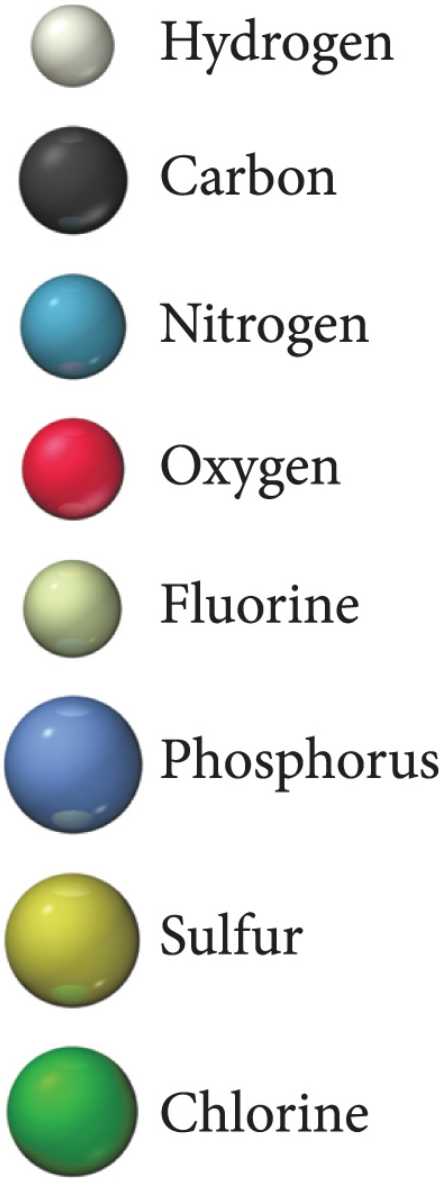
Types of Chemical Formulas

- The type of formula we use depends on how much we know about the compound and how much we want to communicate.
- A structural formula communicates the most information,
- while an empirical formula communicates the least.

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Molecular Models

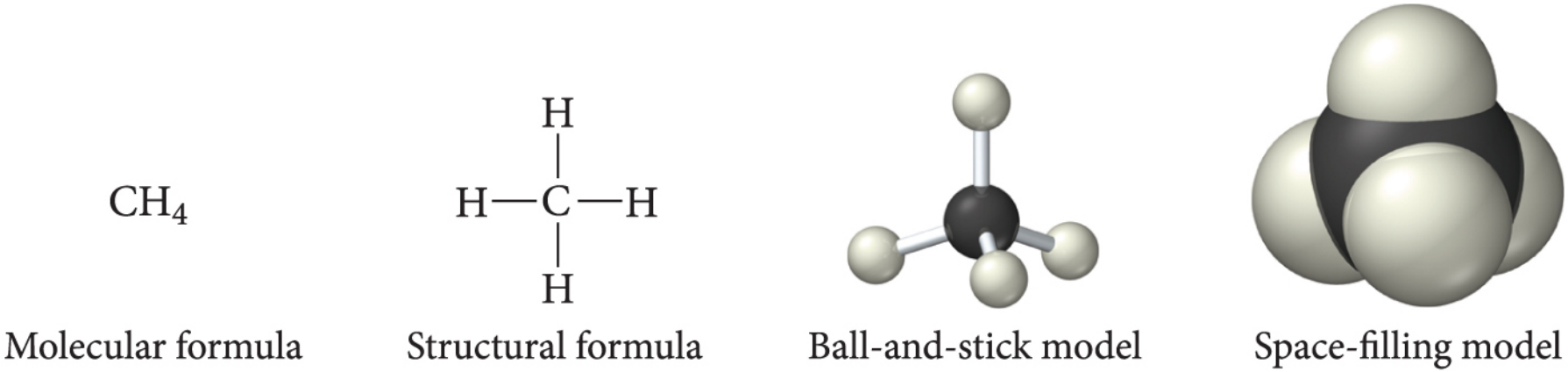
- A *molecular model* is a more accurate and complete way to specify a compound.
- A **ball-and-stick molecular model** represents atoms as balls and chemical bonds as sticks; how the two connect reflects a molecule's shape.
- The balls are typically color-coded to specific elements.



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Molecular Models

- In a **space-filling molecular model**, atoms fill the space between each other to more closely represent our best estimates for how a molecule might appear if scaled to visible size.

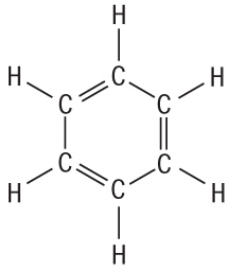
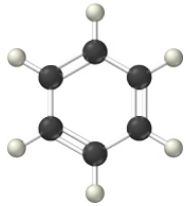
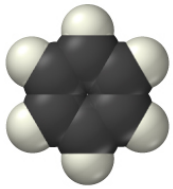
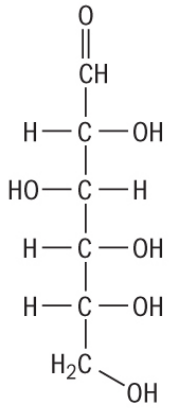
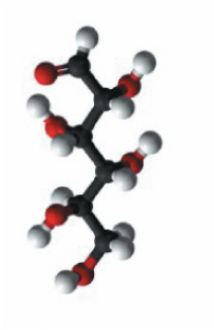
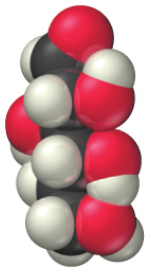
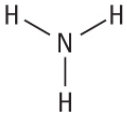
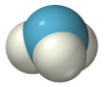


Molecular formula | Structural formula | Ball-and-stick model | Space-filling model

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Ways of Representing a Compound

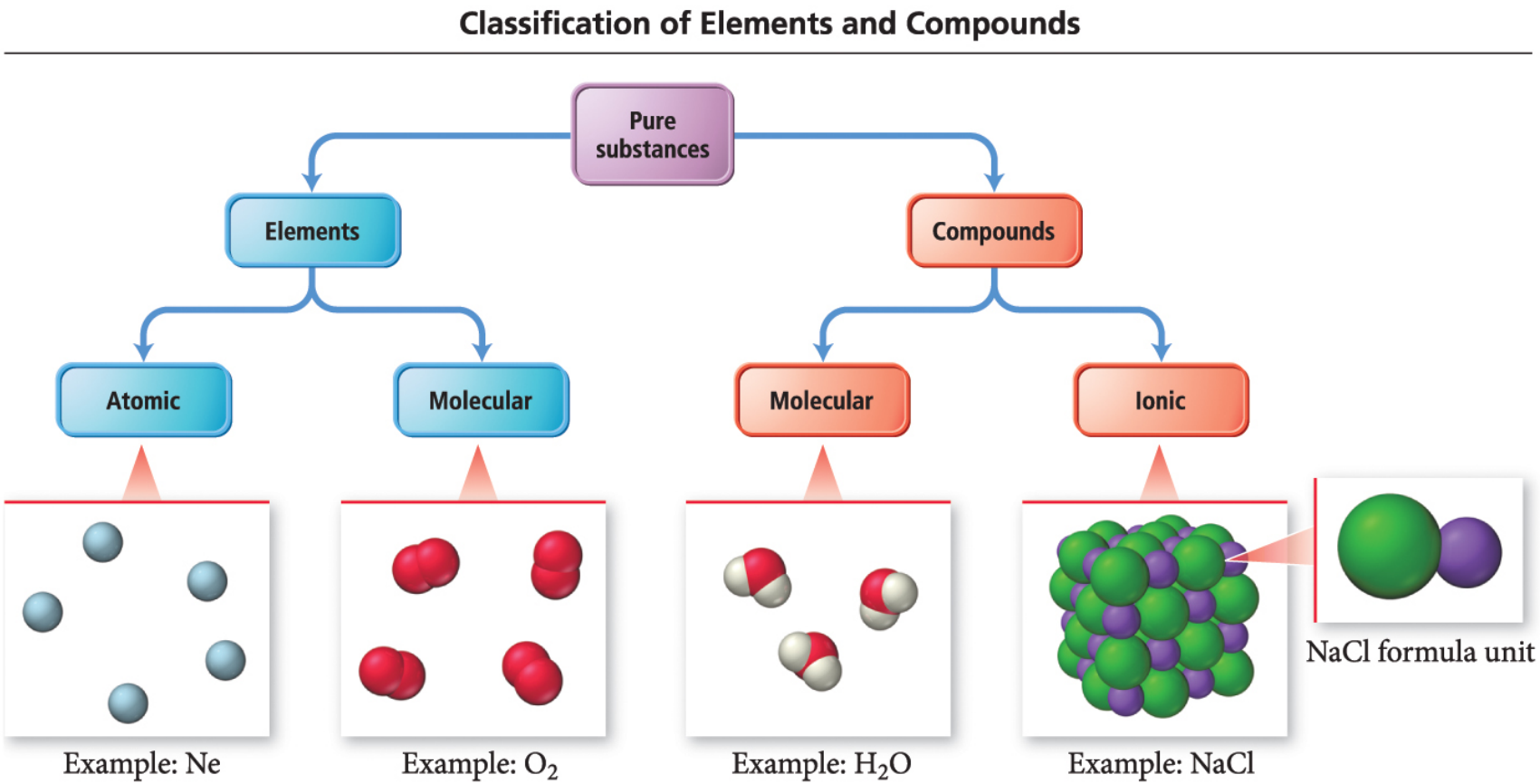
TABLE 3.1 Benzene, Acetylene, Glucose, and Ammonia

Name of Compound	Empirical Formula	Molecular Formula	Structural Formula	Ball-and-Stick Model	Space-Filling Model
Benzene	CH	C_6H_6			
Acetylene	CH	C_2H_2			
Glucose	CH_2O	$C_6H_{12}O_6$			
Ammonia	NH_3	NH_3			

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An Atomic-Level View of Elements and Compounds

- Elements may be either atomic or molecular. Compounds may be either molecular or ionic.



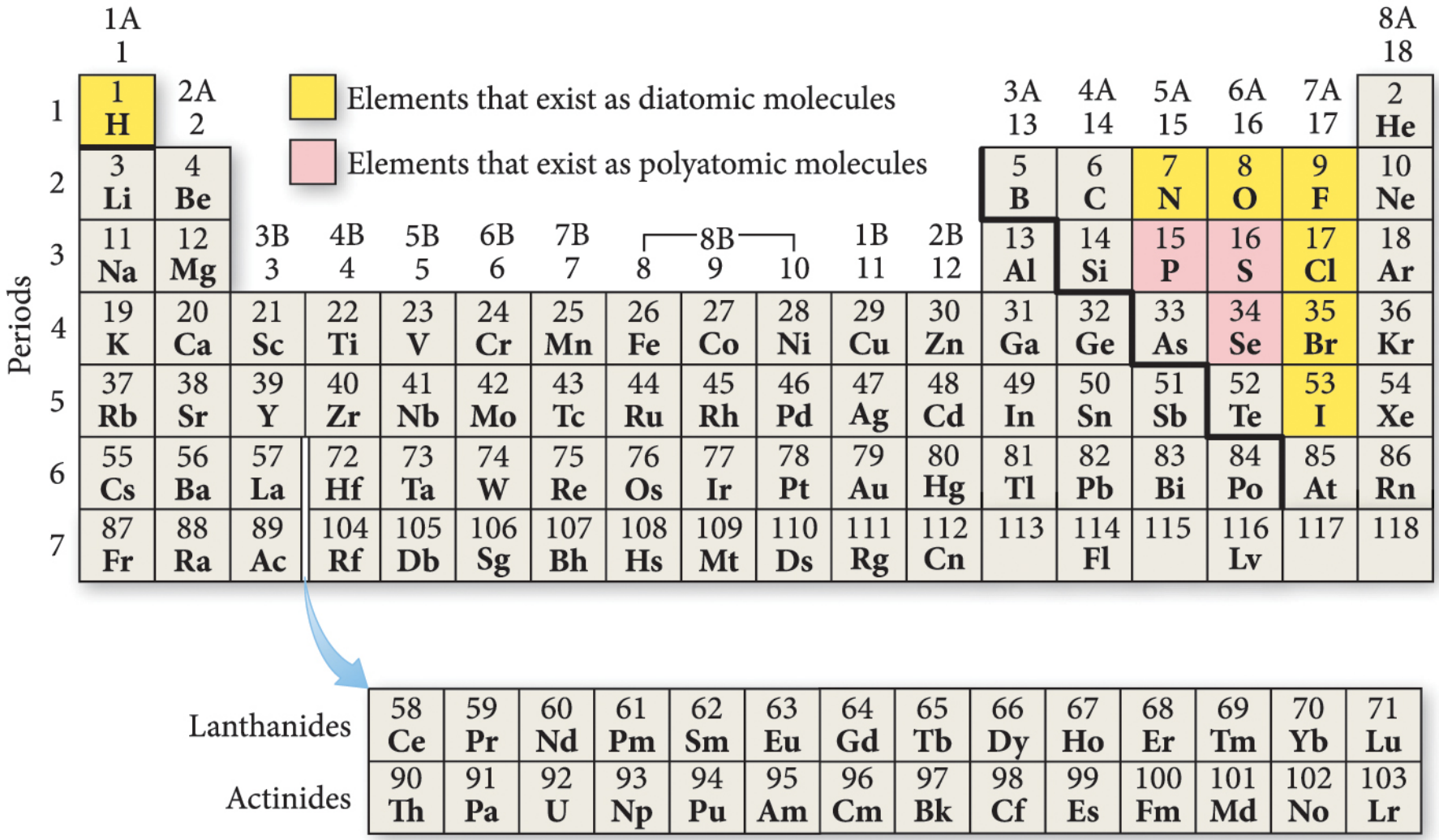
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View of Elements and Compounds

- **Atomic elements** exist in nature with single atoms as their basic units. Most elements fall into this category.
  - Examples are Na, Ne, C, K, Mg, etc.
- **Molecular elements** do not normally exist in nature with single atoms as their basic units; instead, they exist as molecules—two or more atoms of the element bonded together.
  - There only seven diatomic elements and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.
  - Also, P_4 and S_8 are polyatomic elements.

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Molecular Elements



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Molecular Compounds

- **Molecular compounds** are usually composed of two or more covalently bonded nonmetals.
- The basic units of molecular compounds are molecules composed of the constituent atoms.
  - Water is composed of H_2O molecules.
  - Dry ice is composed of CO_2 molecules.
  - Propane (often used as a fuel for grills) is composed of C_3H_8 molecules.

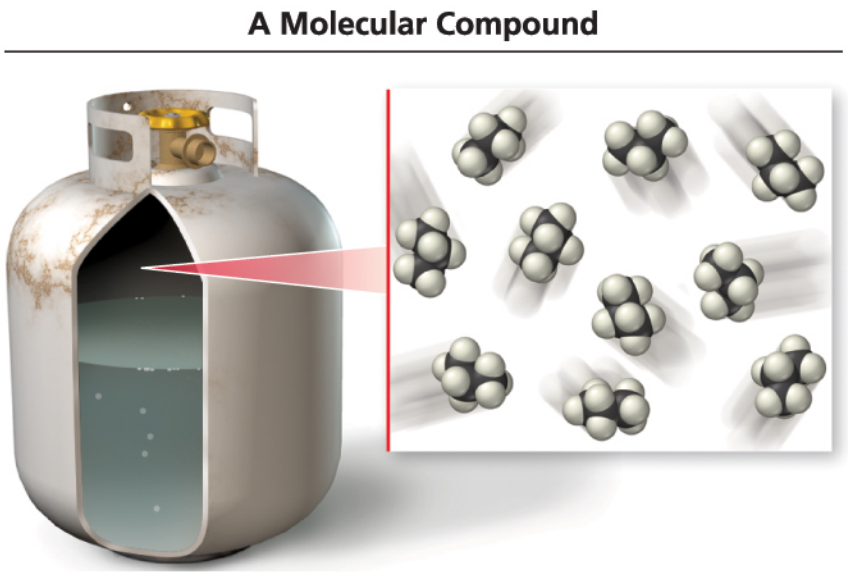
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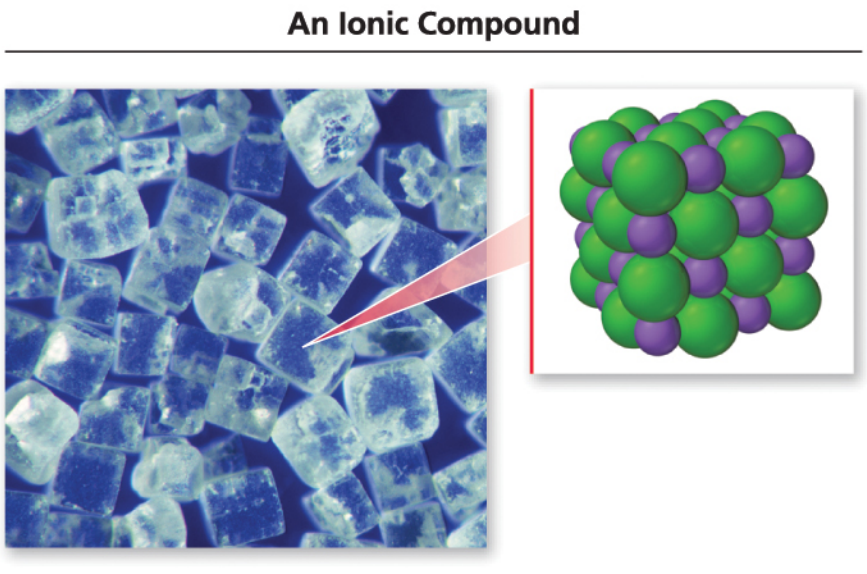
Ionic Compounds

- **Ionic compounds** are composed of cations (usually a metal) and anions (usually one or more nonmetals) bound together by ionic bonds.
- The basic unit of an ionic compound is the **formula unit**, the smallest, electrically neutral collection of ions.
- The ionic compound table salt, with the formula unit NaCl, is composed of Na^+ and Cl^- ions in a one-to-one ratio.

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Molecular and Ionic Compounds





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Polyatomic Ion

- Many common ionic compounds contain ions that are themselves composed of a group of covalently bonded atoms with an overall charge.
- This group of charged species is called polyatomic ions.
  - $NaNO_3$ contains Na^+ and NO_3^-.
  - $CaCO_3$ contains Ca^{2+} and CO_3^{2-}.
  - KClO Contains K^+ and ClO^-.

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Ionic Compounds: Formulas and Names

- ***Summarizing Ionic Compound Formulas:***
  - Ionic compounds always contain positive and negative ions.
  - In a chemical formula, the sum of the charges of the positive ions (cations) must equal the sum of the charges of the negative ions (anions).
  - The formula of an ionic compound reflects the smallest whole-number ratio of ions.

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Ionic Compounds: Formulas and Names

- The charges of the representative elements can be predicted from their group numbers.
- The representative elements forms only one type of charge.
- Transition metals tend to form multiple types of charges.
- Hence, their charge cannot be predicted as in the case of most representative elements.

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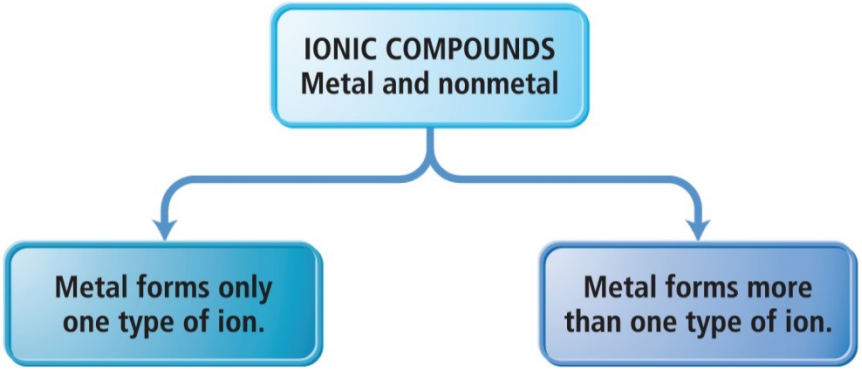
Naming Ionic Compounds

- *Ionic compounds are usually composed of metals and nonmetals*.
- Anytime you see a metal and one or more nonmetals together in a chemical formula, assume that you have an ionic compound.
- NaBr, $Al_2(CO_3)_3$, $CaHPO_4$, and $MgSO_4$ are some examples of ionic compounds.

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Naming Ionic Compounds

- Ionic compounds can be categorized into two types, depending on the metal in the compound.



- The first type contains a metal whose charge is invariant from one compound to another.
- Whenever the metal in this first type of compound forms an ion, the ion always has the same charge.

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Naming Type I Ionic Compounds

TABLE 3.2 Metals Whose Charge Is Invariant from One Compound to Another

Metal	Ion	Name	Group Number
Li	Li^+	Lithium	1A
Na	Na^+	Sodium	1A
K	K^+	Potassium	1A
Rb	Rb^+	Rubidium	1A
Cs	Cs^+	Cesium	1A
Be	Be^{2+}	Beryllium	2A
Mg	Mg^{2+}	Magnesium	2A
Ca	Ca^{2+}	Calcium	2A
Sr	Sr^{2+}	Strontium	2A
Ba	Ba^{2+}	Barium	2A
Al	Al^{3+}	Aluminum	3A
Zn	Zn^{2+}	Zinc	*
Sc	Sc^{3+}	Scandium	*
Ag**	Ag^+	Silver	*

*The charge of these metals cannot be inferred from their group number.

**Silver sometimes forms compounds with other charges, but these are rare.

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Naming Type II Ionic Compounds

- The second type of ionic compound contains a metal with a charge that can differ in different compounds.
- The metals in this second type of ionic compound can form more than one kind of cation (depending on the compound).
- Its charge must therefore be specified for a given compound.

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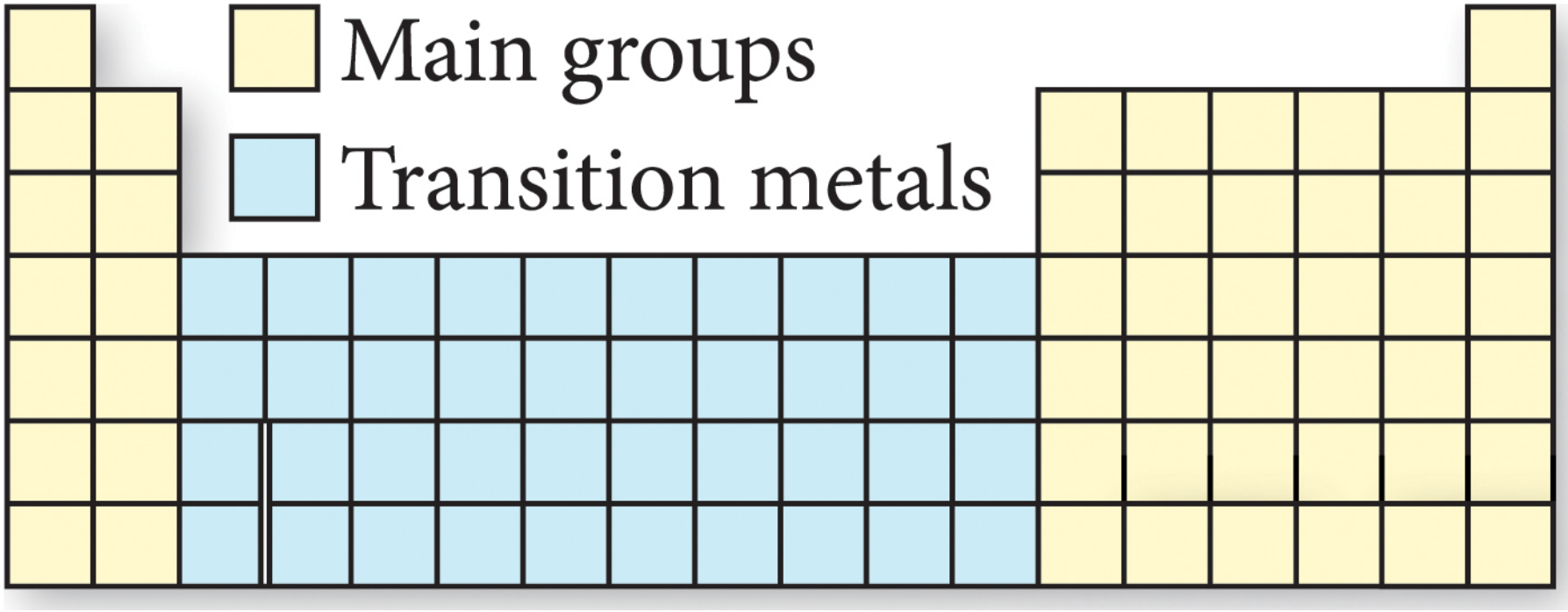
Type II Ionic Compounds

- Iron, for instance, forms a 2+ cation in some of its compounds and a 3+ cation in others.
- Metals of this type are often *transition metals*.
  - $FeSO_4$ Here iron is +2 cation (Fe^{2+}).
  - $Fe_2(SO_4)_3$ Here iron is +3 cation (Fe^{3+}).
  - Cu_2O Here copper is +1 cation (Cu^{+}).
  - CuO Here copper is +2 cation (Cu^{2+}).

Some main group metals, such as Pb, Tl, and Sn, form more than one type of cation.

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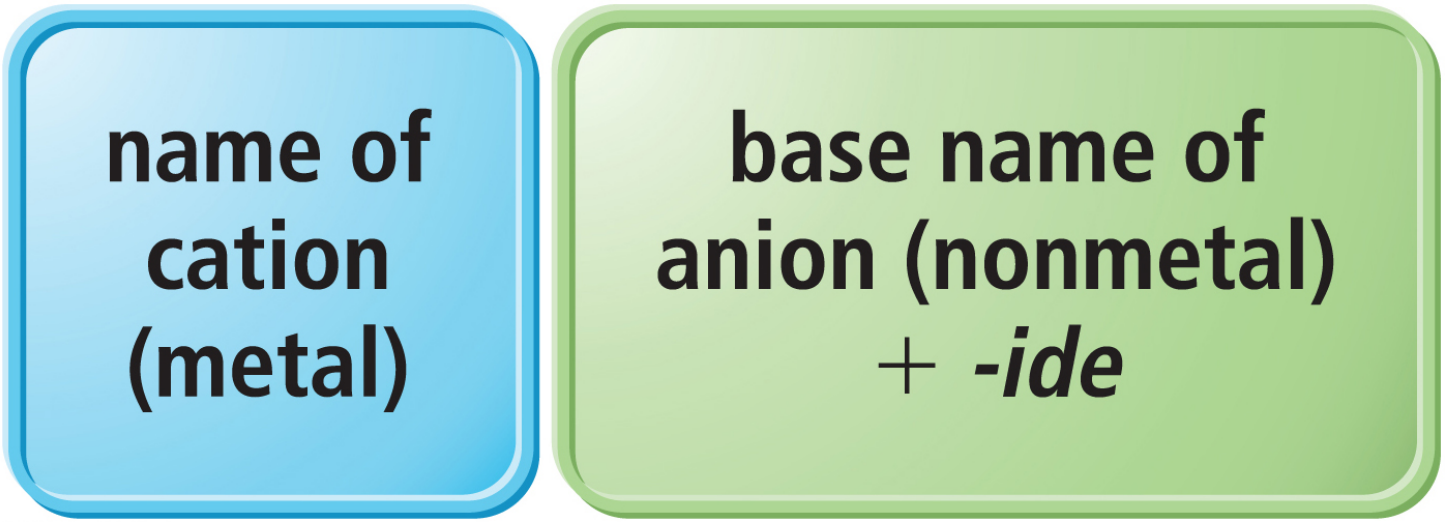
Type II Ionic Compounds



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Naming Binary Ionic Compounds of Type I Cations

- **Binary compounds** contain only two different elements. The names of binary ionic compounds take the following form:



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Naming Type I Binary Ionic Compounds

- For example, the name for KCl consists of the name of the cation, *potassium*, followed by the base name of the anion, *chlor*, with the ending *-ide*.
  - KCl is *potassium chloride*.
- The name for CaO consists of the name of the cation, *calcium*, followed by the base name of the anion, *ox*, with the ending *-ide.*
  - CaO is *calcium oxide*.

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Base Names of Monoatomic Anions

- The base names of some nonmetals, and their most common charges in ionic compounds, are shown in Table 3.3.

TABLE 3.3 Some Common Monoatomic Anions

Nonmetal	Symbol for Ion	Base Name	Anion Name
Fluorine	F^-	fluor	Fluoride
Chlorine	Cl^-	chlor	Chloride
Bromine	Br^-	brom	Bromide
Iodine	I^-	iod	Iodide
Oxygen	O^{2-}	ox	Oxide
Sulfur	S^{2-}	sulf	Sulfide
Nitrogen	N^{3-}	nitr	Nitride
Phosphorus	P^{3-}	phosph	Phosphide

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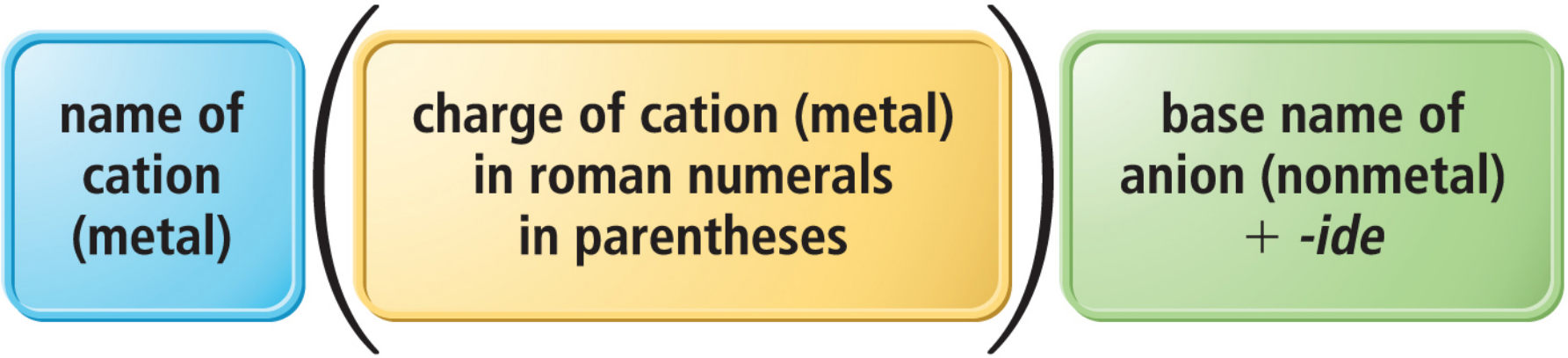
Naming Type II Binary Ionic Compounds

- For these types of metals, the name of the cation is followed by a roman numeral (in parentheses) that indicates the charge of the metal in that particular compound.
  - For example, we distinguish between Fe^{2+} and Fe^{3+} as follows:
    - Fe^{2+} Iron(II)
    - Fe^{3+} Iron(III)

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Naming Type II Binary Ionic Compounds

- The full names for compounds containing metals that form more than one kind of cation have the following form:



The charge of the metal cation can be determined by inference from the sum of the charges of the nonmetal.

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Naming Type II Binary Ionic Compounds

- For example, to name $CrBr_3$ determine the charge on the chromium.
  - Total charge on cation + total anion charge = 0.
  - Cr charge + 3(Br^- charge) = 0.
  - Since each Br has a –1 charge, then
    - Cr charge + 3(–1) = 0
    - Cr charge –3 = 0
    - Cr = +3
  - Hence, the cation Cr^{3+} is called chromium(III), while Br^- is called bromide.

Therefore, $CrBr_3$ is **chromium(III) bromide.**

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Type II Cation

TABLE 3.4 Some Metals That Form Cations with Different Charges

Metal	Ion	Name	Older Name*
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	Chromium(III)	Chromic
Iron	Fe^{2+}	Iron(II)	Ferrous
	Fe^{3+}	Iron(III)	Ferric
Cobalt	Co^{2+}	Cobalt(II)	Cobaltous
	Co^{3+}	Cobalt(III)	Cobaltic
Copper	Cu^{+}	Copper(I)	Cuprous
	Cu^{2+}	Copper(II)	Cupric
Tin	Sn^{2+}	Tin(II)	Stannous
	Sn^{4+}	Tin(IV)	Stannic
Mercury	Hg_2^{2+}	Mercury(I)	Mercurous
	Hg^{2+}	Mercury(II)	Mercuric
Lead	Pb^{2+}	Lead(II)	Plumbous
	Pb^{4+}	Lead(IV)	Plumbic

*An older naming system substitutes the names found in this column for the name of the metal and its charge. Under this system, chromium(II) oxide is named chromous oxide. Additionally, the suffix *-ous* indicates the ion with the lesser charge and *-ic* indicates the ion with the greater charge. We will *not* use the older system in this text.

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Naming Ionic Compounds Containing Polyatomic Ions

- We name ionic compounds that contain a polyatomic ion in the same way as other ionic compounds, except that we use the name of the polyatomic ion whenever it occurs.
- For example, $NaNO_2$ is named according to
  - its cation, Na^+, *sodium*, and
  - its polyatomic anion, NO_2^-, *nitrite*.
- Hence, $NaNO_2$ is *sodium nitrite*.

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Common Polyatomic Ions

TABLE 3.5 Some Common Polyatomic Ions

Name	Formula	Name	Formula
Acetate	$C_2H_3O_2^-$	Hypochlorite	ClO^-
Carbonate	CO_3^{2-}	Chlorite	ClO_2^-
Hydrogen carbonate (or bicarbonate)	HCO_3^-	Chlorate	ClO_3^-
Hydroxide	OH^-	Perchlorate	ClO_4^-
Nitrite	NO_2^-	Permanganate	MnO_4^-
Nitrate	NO_3^-	Sulfite	SO_3^{2-}
Chromate	CrO_4^{2-}	Hydrogen sulfite (or bisulfite)	HSO_3^-
Dichromate	$Cr_2O_7^{2-}$	Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}	Hydrogen sulfate (or bisulfate)	HSO_4^-
Hydrogen phosphate	HPO_4^{2-}	Cyanide	CN^-
Dihydrogen phosphate	$H_2PO_4^-$	Peroxide	O_2^{2-}
Ammonium	NH_4^+		

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Oxyanions

- Most polyatomic ions are **oxyanions**, anions containing oxygen and another element.
- Notice that when a series of oxyanions contains different numbers of oxygen atoms, they are named according to the number of oxygen atoms in the ion.
- If there are two ions in the series,
  - the one with more oxygen atoms has the ending *-ate*, and
  - the one with fewer has the ending *-ite*.
- For example,
  - NO_3^- is ***nitrate***
  - NO_2^- is ***nitrite***

  SO_4^{2-} is ***sulfate***

  SO_3^{2-} is ***sulfite***

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Oxyanions

- If there are more than two ions in the series then the prefixes *hypo-*, meaning *less than*, and *per-*, meaning *more than*, are used.

ClO^-	*hypo*chlor*ite*	BrO^-	*hypo*brom*ite*
ClO_2^-	chlor*ite*	BrO_2^-	brom*ite*
ClO_3^-	chlor*ate*	BrO_3^-	brom*ate*
ClO_4^-	*per*chlor*ate*	BrO_4^-	*per*brom*ate*

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Hydrated Ionic Compounds

- **Hydrates** are ionic compounds containing a specific number of water molecules associated with each formula unit.
  - For example, the formula for epsom salts is $MgSO_4 \cdot 7H_2O$.
  - Its systematic name is magnesium sulfate heptahydrate.
  - $CoCl_2 \cdot 6H_2O$ is cobalt(II)chloride hexahydrate.

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Hydrates

Common hydrate prefixes

- hemi = ½
- mono = 1
- di = 2
- tri = 3
- tetra = 4
- penta = 5
- hexa = 6
- hepta = 7
- octa = 8

Other common hydrated ionic compounds and their names are as follows:

- $CaSO_4 \cdot 1/2H_2O$ is called calcium sulfate hemihydrate.
- $BaCl_2 \cdot 6H_2O$ is called barium chloride hexahydrate.
- $CuSO_4 \cdot 6H_2O$ is called copper sulfate hexahydrate.

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Molecular Compounds: Formulas and Names

- The formula for a molecular compound *cannot* readily be determined from its constituent elements because the same combination of elements may form many different molecular compounds, each with a different formula.
  - Nitrogen and oxygen form all of the following unique molecular compounds: NO, NO_2, N_2O, N_2O_3, N_2O_4, and N_2O_5.

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Molecular Compounds

- *Molecular compounds are composed of two or more nonmetals*.
- Generally, write the name of the element with the smallest group number first.
- If the two elements lie in the same group, then write the element with the greatest row number first.
  - The prefixes given to each element indicate the number of atoms present.

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Binary Molecular Compounds



- These prefixes are the same as those used in naming hydrates:

mono = 1	hexa = 6
di = 2	hepta = 7
tri = 3	octa = 8
tetra = 4	nona = 9
penta = 5	deca = 10

If there is only one atom of the *first element* in the formula, the prefix *mono*- is normally omitted.

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Acids

- Acids are molecular compounds that release hydrogen ions (H^+) when dissolved in water.
- Acids are composed of hydrogen, usually written first in their formula, and one or more nonmetals, written second.
  - HCl is a molecular compound that, when dissolved in water, forms $H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions, where *aqueous* (*aq*) means *dissolved in water*.

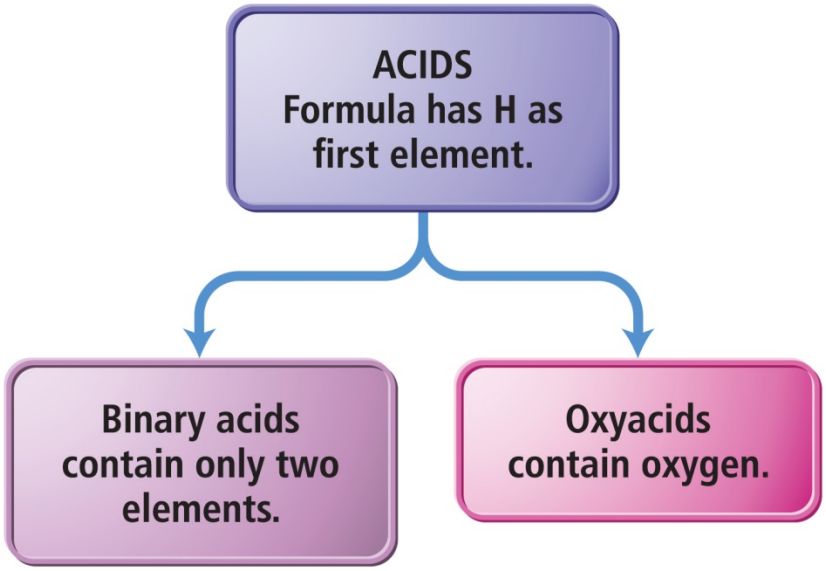
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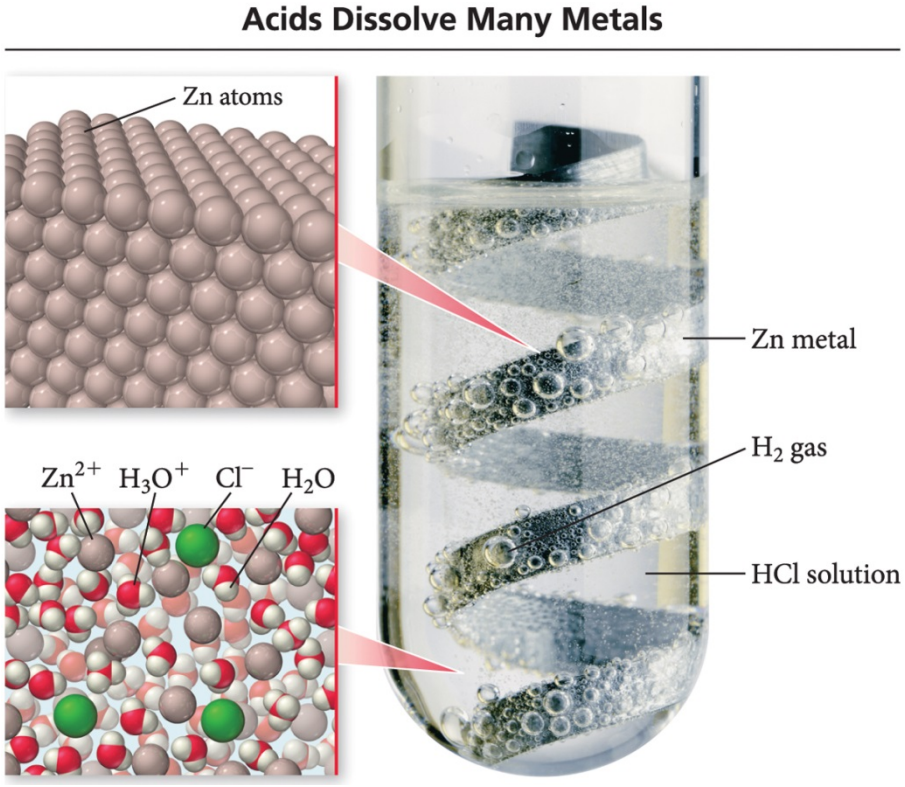
Acids

- Acids are molecular compounds that form H^+ when dissolved in water.
  - To indicate the compound is dissolved in water (*aq*) is written after the formula.
    - » A compound is not considered an acid if it does not dissolve in water.
- Sour taste
- Dissolve many metals
  - such as Zn, Fe, Mg; but not Au, Ag, Pt
- Formula generally starts with H
  - e.g., HCl, H_2SO_4

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Acids



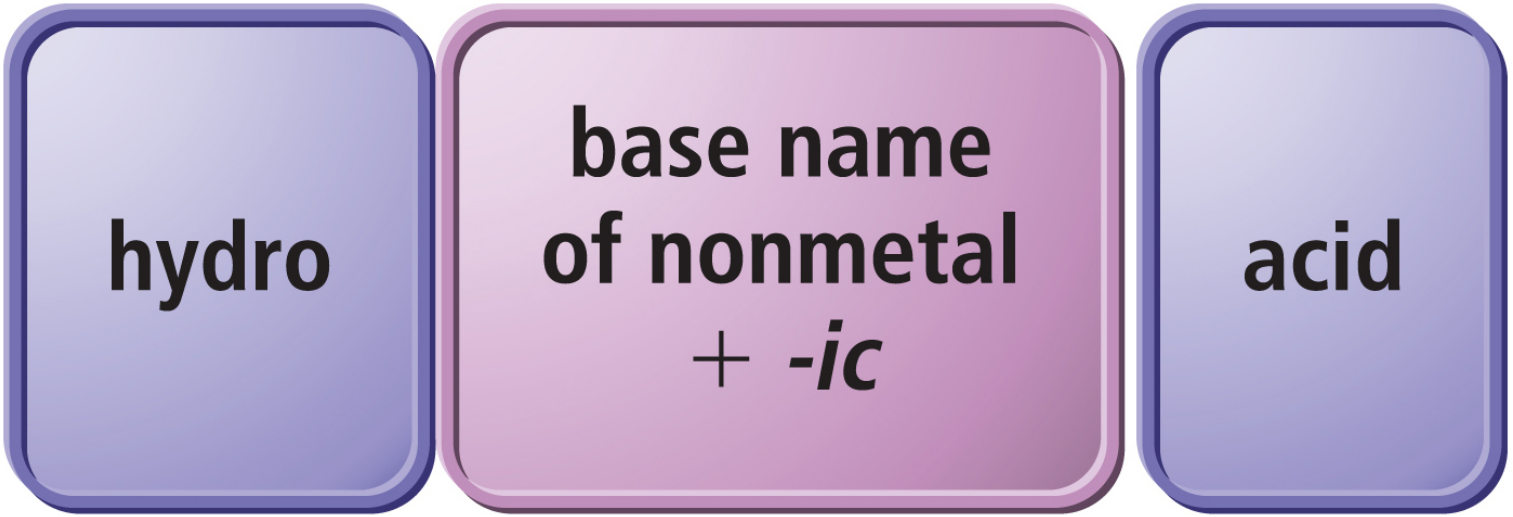


- Binary acids have H^{+1} cation and nonmetal anion.
- Oxyacids have H^+ cation and polyatomic anion.

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Naming Binary Acids

- Write a ***hydro-*** prefix.
- Follow with the nonmetal name.
- Change ending on nonmetal name to *–ic*.
- Write the word **acid** at the end of the name.



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Naming Oxyacids

- If polyatomic ion name ends in *–ate*, then change ending to *–ic* suffix.
- If polyatomic ion name ends in *–ite*, then change ending to *–ous* suffix.
- Write word **acid** at the end of all names.

oxyanions ending with *-ate*	base name of oxyanion + *-ic* acid
oxyanions ending with *-ite*	base name of oxyanion + *-ous* acid

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Name the Following

1. H_2S

2. $HClO_3$

3. $HC_2H_3O_2$

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Name the Following

1. H_2S — hydrosulfuric acid

2. $HClO_3$ — chloric acid

3. $HC_2H_3O_2$ — acetic acid

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Writing Formulas for Acids

- When name ends in **acid**, formulas starts with **H**.
- Write formulas as if ionic, even though it is molecular.
- *Hydro-* prefix means it is binary acid; no prefix means it is an oxyacid.
- For oxyacid
  - if ending is ***–ic***, polyatomic ion ends in *–ate*.
  - if ending is ***–ous***, polyatomic ion ends in *–ous*.

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Acid Rain

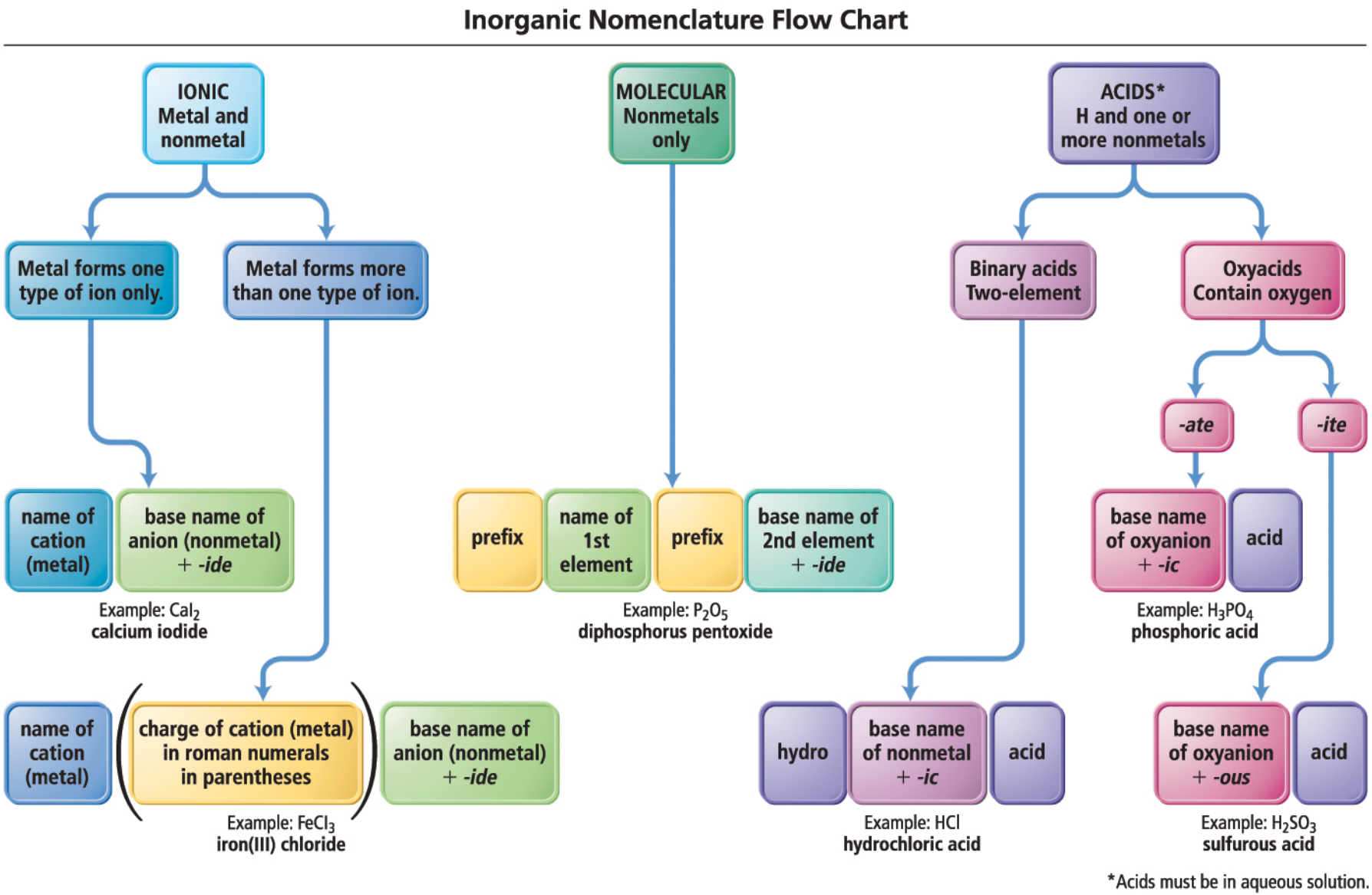
- Certain pollutants—such as NO, NO_2, SO_2, SO_3—form acids when mixed with water, resulting in acidic rainwater.
- Acid rain can fall or flow into lakes and streams, making these bodies of water more acidic.





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Inorganic Nomenclature Flow Chart

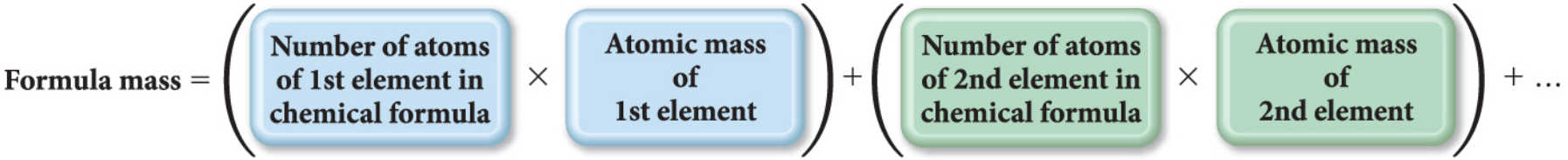


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Formula Mass

- The mass of an individual molecule or formula unit
  - ✓ also known as molecular mass or molecular weight
- Sum of the masses of the atoms in a single molecule or formula unit
  - ✓ whole = sum of the parts!
    - Mass of 1 molecule of H_2O

= 2(1.01 amu H) + 16.00 amu O = 18.02 amu



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Molar Mass of Compounds

The *molar mass of a compound*—the mass in grams of 1 mol of its molecules or formula units—is numerically equivalent to its formula mass.

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Molar Mass of Compounds

- The relative masses of molecules can be calculated from atomic masses:

  formula mass = 1 molecule of H_2O = 2(1.01 amu H) + 16.00 amu O = 18.02 amu

- 1 mole of H_2O contains 2 moles of H and 1 mole of O:

  molar mass = 1 mole H_2O

  = 2(1.01 g H) + 16.00 g O = 18.02 g

  so the molar mass of H_2O is 18.02 g/mole

- Molar mass = formula mass (in g/mole)

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Using Molar Mass to Count Molecules by Weighing

- Molar mass in combination with Avogadro's number can be used to determine the number of atoms in a given mass of the element.
  - Use molar mass to convert to the amount in moles. Then use Avogadro's number to convert to number of molecules.

$$\#\,molecule\,A = \frac{gram\,A}{1} * \frac{1\,mole\,A}{molar\,mass\,A} * \frac{6.02\,x\,10^{23}\,molecules\,A}{1\,mole\,A}$$

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Composition of Compounds

A chemical formula, in combination with the molar masses of its constituent elements, indicates the relative quantities of each element in a compound, which is extremely useful information.

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Composition of Compounds

- Percentage of each element in a compound
  - ✓ by mass
- Can be determined from
  1. the formula of the compound and
  2. the experimental mass analysis of the compound.
- The percentages may not always total to 100% due to rounding.

$$\text{mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100$$

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Conversion Factors from Chemical Formula

- Chemical formulas contain within them inherent relationships between numbers of atoms and molecules.
  - Or moles of atoms and molecules

$$1\ \text{mol}\ CCl_2F_2 : 2\ \text{mol}\ Cl$$

- These relationships can be used to determine the amounts of constituent elements and molecules.
  - Like percent composition

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Determining a Chemical Formula from Experimental Data

Empirical Formula

- Simplest, whole-number ratio of the atoms of elements in a compound
- Can be determined from elemental analysis
  - Masses of elements formed when a compound is decomposed, or that react together to form a compound
    - Combustion analysis
  - Percent composition

Note: An empirical formula represents a ratio of atoms or a ratio of moles of atoms, *not a ratio of masses*.

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Finding an Empirical Formula

1. Convert the percentages to grams.
    a) Assume you start with 100 g of the compound.
    b) Skip if already grams.
2. Convert grams to moles.
    a) Use molar mass of each element.
3. Write a pseudoformula using moles as subscripts.

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Finding an Empirical Formula

4. Divide all by smallest number of moles.
   a) If the result is within 0.1 of a whole number, round to the whole number.
5. Multiply all mole ratios by a number to make all whole numbers.
   a) If ratio .5, multiply all by 2.
   b) if ratio .33 or .67, multiply all by 3.
   c) If ratio 0.25 or 0.75, multiply all by 4, etc.
   d) Skip if already whole numbers.

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Molecular Formulas for Compounds

- The molecular formula is a multiple of the empirical formula.
- To determine the molecular formula you need to know the empirical formula and the molar mass of the compound.

Molecular formula = (empirical formula)***n***,

*where **n** is a positive integer.*

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Molecular Formulas for Compounds

- The molar mass is a whole-number multiple of the **empirical formula molar mass**, the sum of the masses of all the atoms in the empirical formula:

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

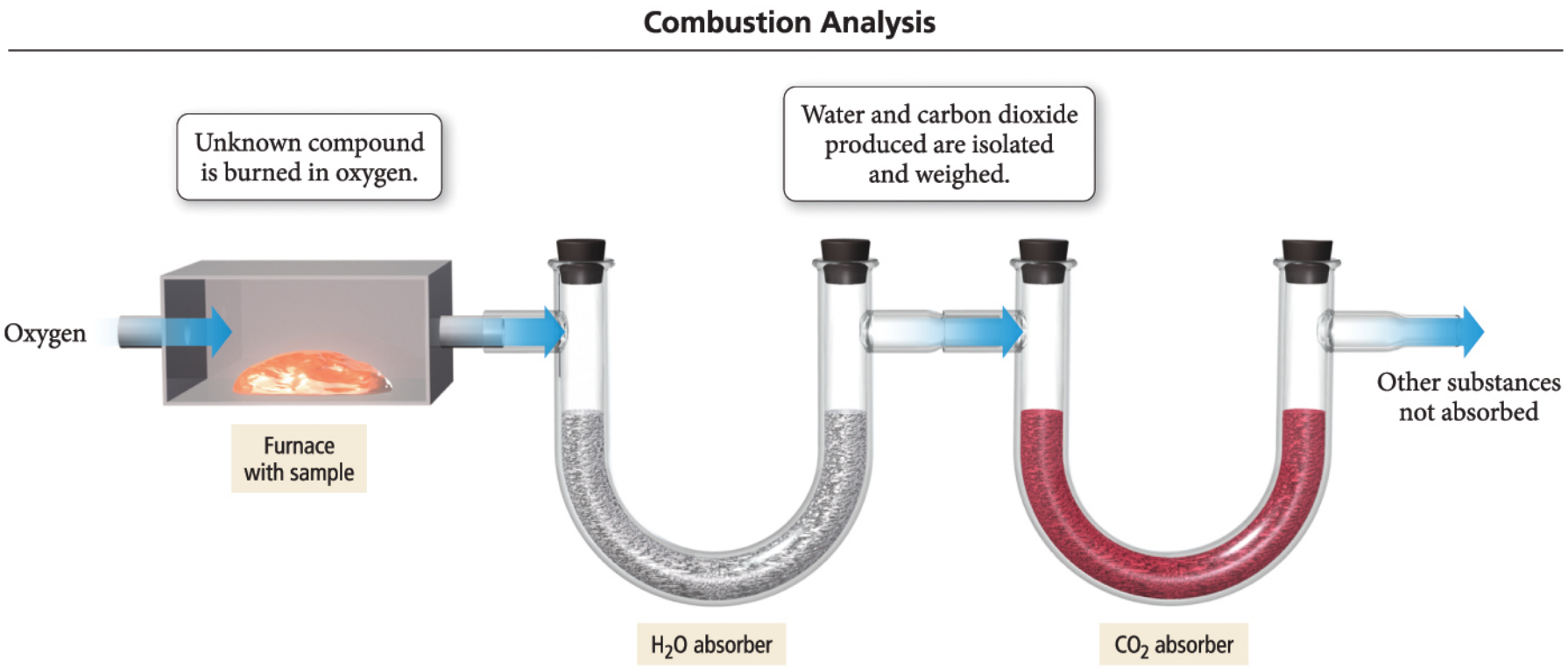
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Combustion Analysis

- A common technique for analyzing compounds is to burn a known mass of compound and weigh the amounts of product made.
  - This is generally used for organic compounds containing C, H, O.
- By knowing the mass of the product and composition of constituent element in the product, the original amount of constituent element can be determined.
  - All the original C forms CO_2, the original H forms H_2O, and the original mass of O is found by subtraction.
- Once the masses of all the constituent elements in the original compound have been determined, the empirical formula can be found.

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Combustion Analysis



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Chemical Reactions

- Reactions involve chemical changes in matter resulting in new substances.
- Reactions involve rearrangement and exchange of atoms to produce new molecules.
  - Elements are not transmuted during a reaction.

Reactants → Products

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Chemical Equations

- Shorthand way of describing a reaction
- Provide information about the reaction
  - Formulas of reactants and products
  - States of reactants and products
  - Relative numbers of reactant and product molecules that are required
  - Can be used to determine weights of reactants used and products that can be made

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TABLE 3.6 States of Reactants and Products in Chemical Equations

Abbreviation	State
(*g*)	Gas
(*l*)	Liquid
(*s*)	Solid
(*aq*)	Aqueous (water solution)

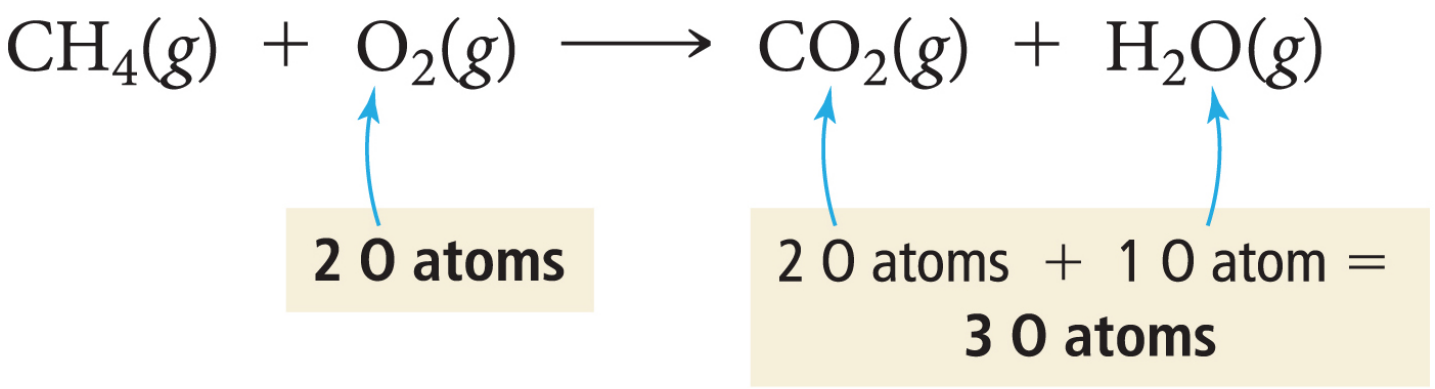
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Combustion of Methane

- Methane gas burns to produce carbon dioxide gas and gaseous water.
  - Whenever something burns it combines with $O_2(g)$.

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

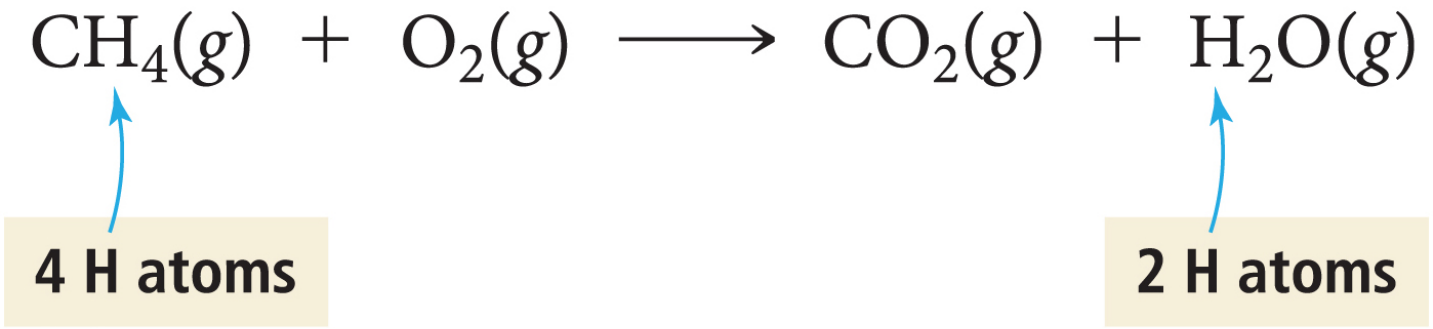
- If you look closely, you should immediately spot a problem.



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Combustion of Methane

- Notice also that the left side has four hydrogen atoms while the right side has only two.



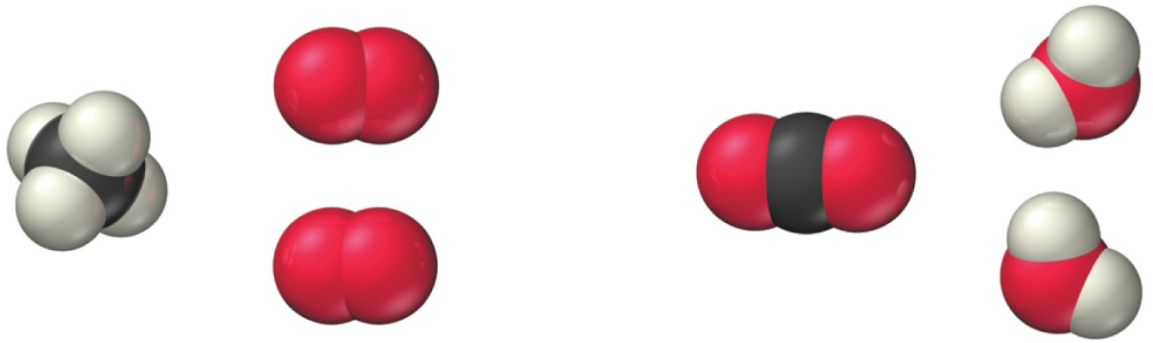
- To correct these problems, we must **balance** the equation by changing the coefficients, not the subscripts.

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Combustion of Methane, Balanced

- To show the reaction obeys the Law of Conservation of Mass the equation must be **balanced**.
  - We adjust the numbers of molecules so there are equal numbers of atoms of each element on both sides of the arrow.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$



1 C + 4 H + 4 O ⟶ 1 C + 4 H + 4 O

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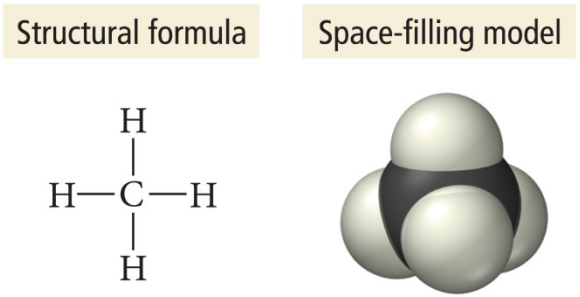
Organic Compounds

- Early chemists divided compounds into two types: organic and inorganic.
- Compounds from living things were called organic; compounds from the nonliving environment were called inorganic.
- Organic compounds are easily decomposed and could not be made in the lab.
- Inorganic compounds are very difficult to decompose, but are able to be synthesized.

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Modern Organic Compounds

- Today organic compounds are commonly made in the lab and we find them all around us.
- Organic compounds are mainly made of C and H, sometimes with O, N, P, S, and trace amounts of other elements
- The main element that is the focus of organic chemistry is **carbon**.



Methane, CH_4

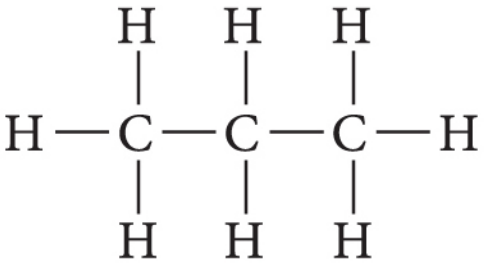
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Carbon Bonding

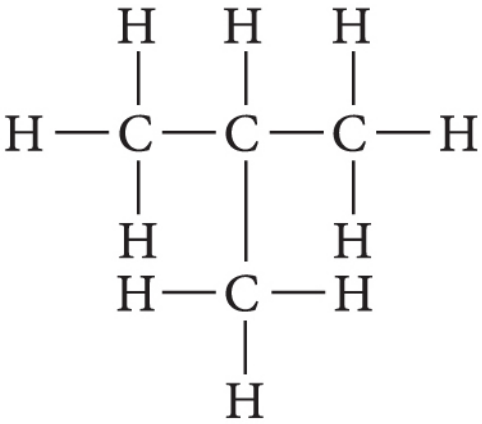
- Carbon atoms bond almost exclusively covalently.
  - Compounds with ionic bonding C are generally inorganic.
- When C bonds, it forms four covalent bonds:
  - 4 single bonds, 2 double bonds, 1 triple + 1 single, etc.
- Carbon is unique in that it can form limitless chains of C atoms, both straight and branched, and rings of C atoms.

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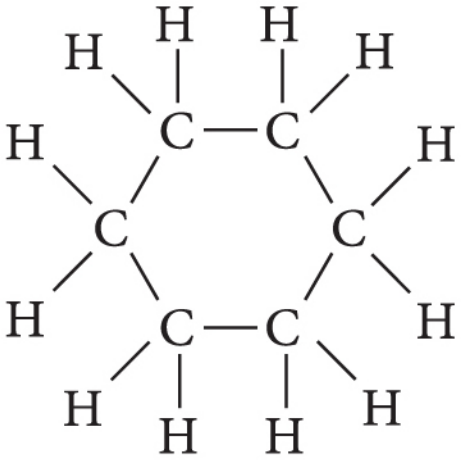
Carbon Bonding



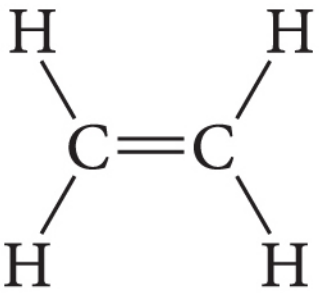
Propane (C_3H_8)



Isobutane (C_4H_{10})



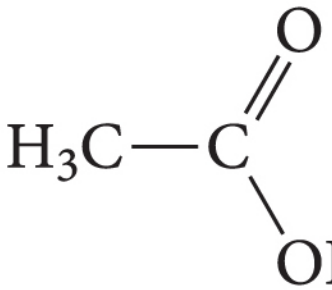
Cyclohexane (C_6H_{12})



Ethene (C_2H_4)



Ethyne (C_2H_2)

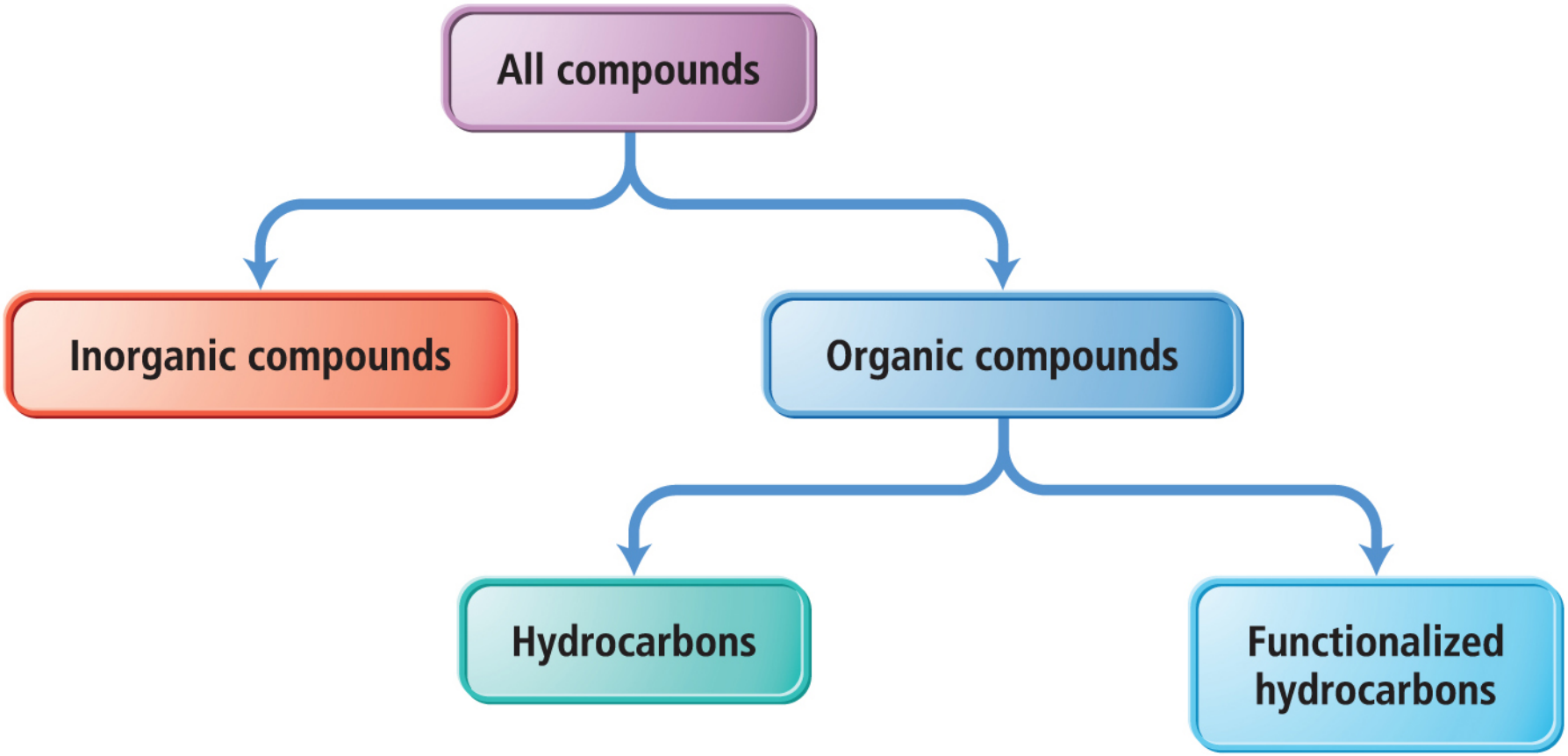


Acetic acid (CH_3COOH)

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Hydrocarbons

- Organic compounds can be categorizing into types: hydrocarbons and functionalized hydrocarbons.



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Hydrocarbons

- **Hydrocarbons** are organic compounds that contain only carbon and hydrogen.
- Hydrocarbons compose common fuels such as
  - oil,
  - gasoline,
  - liquid propane gas,
  - and natural gas.



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Naming of Hydrocarbons

- Hydrocarbons containing only single bonds are called **alkanes**,
- while those containing double or triple bonds are **alkenes** and **alkynes**, respectively.
- Hydrocarbons consist of a base name and a suffix.
  - alkane (-*ane*)
  - alkene (-*ene*)
  - alkyne (-*yne*)

- The base names for a number of hydrocarbons are listed here:
  - 1 meth 2 eth
  - 3 prop 4 but
  - 5 pent 6 hex
  - 7 hept 8 oct
  - 9 non 10 dec

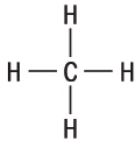
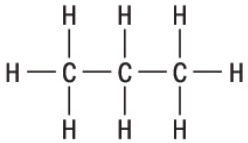
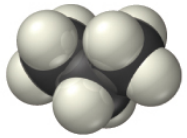
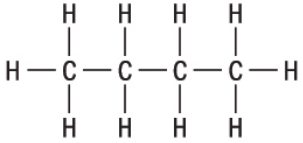
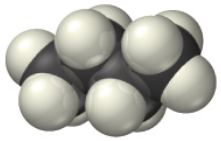
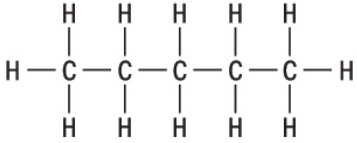
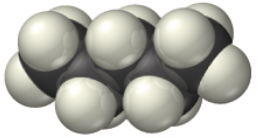
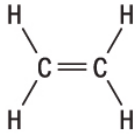
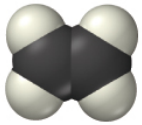
Base name determined by number of C atoms

Suffix determined by presence of multiple bonds

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Common Hydrocarbons

TABLE 3.7 Common Hydrocarbons

Name	Molecular Formula	Structural Formula	Space-filling Model	Common Uses
Methane	CH_4			Primary component of natural gas
Propane	C_3H_8			LP gas for grills and outdoor stoves
n-Butane*	C_4H_{10}			Common fuel for lighters
n-Pentane*	C_5H_{12}			Component of gasoline
Ethene	C_2H_4			Ripening agent in fruit
Ethyne	C_2H_2			Fuel for welding torches

*The "*n*" in the names of these hydrocarbons stands for "normal", which means straight chain.

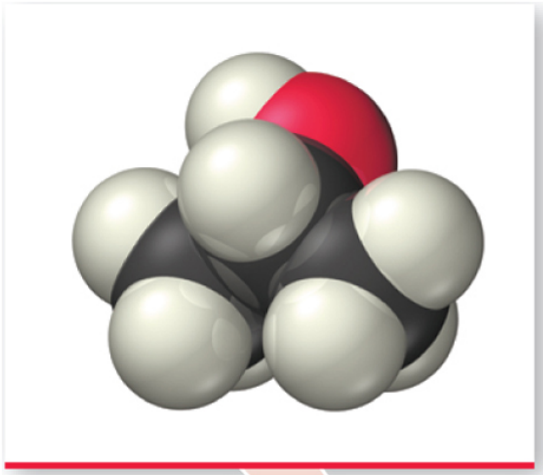
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Functionalized Hydrocarbons

- The term ***functional group*** derives from the functionality or chemical character that a specific atom or group of atoms imparts to an organic compound.
  - Even a carbon–carbon double or triple bond can justifiably be called a "functional group."
- A group of organic compounds with the same functional group forms a **family**.

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Functionalized Hydrocarbons



Hydrocarbon (R) group → CH_3OH ← OH functional group

Methanol

Hydrocarbon (R) group → CH_3CHOH ← OH functional group

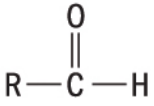
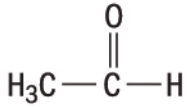
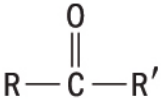
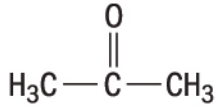
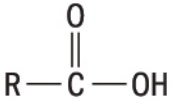
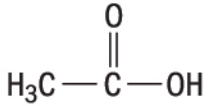
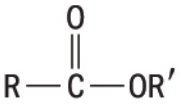
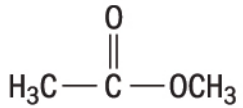
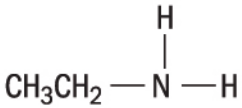
Isopropanol
(2-propanol)



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Families in Organic Compounds

TABLE 3.8 Families of Organic Compounds

Family	Name Ending	General Formula	Example	Name	Occurrence/Use
Alcohols	-ol	R—OH	CH_3CH_2—OH	Ethanol (ethyl alcohol)	Alcohol in fermented beverages
Ethers	ether	R—O—R′	CH_3CH_2—O—CH_2CH_3	Diethyl ether	Anesthetic; laboratory solvent
Aldehydes	-al			Ethanal (acetaldehyde)	Perfumes; flavors
Ketones	-one			Propanone (acetone)	Fingernail polish remover
Carboxylic acids	acid			Acetic acid	Vinegar
Esters	-ate			Methyl acetate	Laboratory solvent
Amines	amine	RNH_2		Ethyl amine	Smell of rotten fish

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