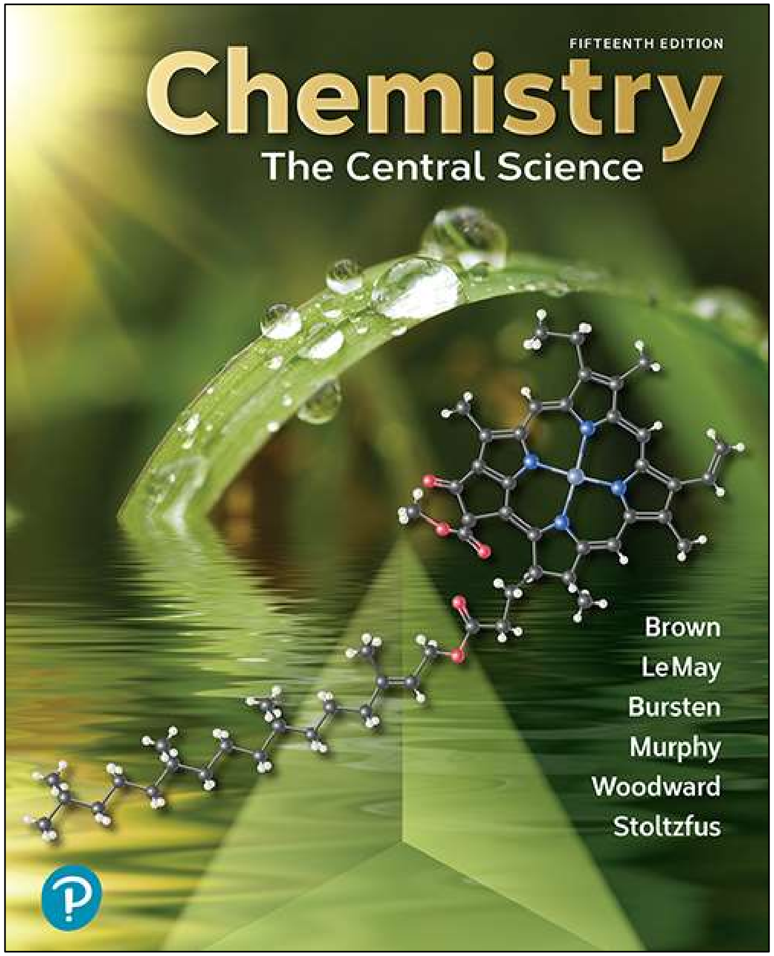
Chemistry: The Central Science



Chapter 1

Introduction:
Matter, Energy,
and Measurement

Chemistry the “central” science.

- Understanding what substances are made of and how they behave is the essence of the field of chemistry.
- Regardless of your science interest, chemistry is essential.
  - Examples
    - Health Care
    - Physics, Construction, Engineering
    - Geology, Agriculture
    - Environmental (CSN)
- Chemistry is sometimes referred to as “the central science,” because it interfaces with every other field in the sciences. Even if your passion lies in art, music, business, or sports, you can find ways that chemistry overlaps with these disciplines.

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So how is chemistry defined?

1.1 Chemistry

- Chemistry is the study of matter and its changes.
- Matter is defined as anything that has mass and takes up volume.
  - For example, in this image, we see clouds, air, water, and a distant sailboat. All of these are composed of matter. As we explore chemistry, we'll explore what these different forms of matter are made of, and how these unique substances behave.

Composition and Structure

- When describing matter, we often refer to something’s composition or to its structure.
  a. Composition refers to what something is made of.
  b. Structure is a slightly broader term – it refers to what something is made of, but it also refers to how the components are arranged.
- Let's look as some examples

Composition

What something is made of?

All made of wood.

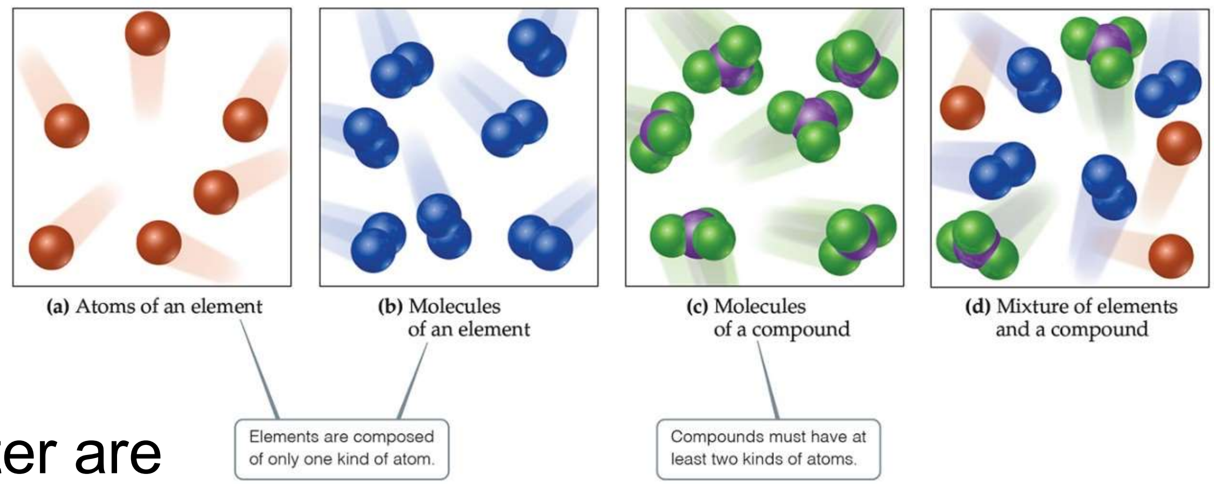
Structure

How the components are arranged?

Arrangement of the wood components allows each one to fulfill a different function.

1.2 Classifications of Matter

Matter is anything that has mass and takes up space.

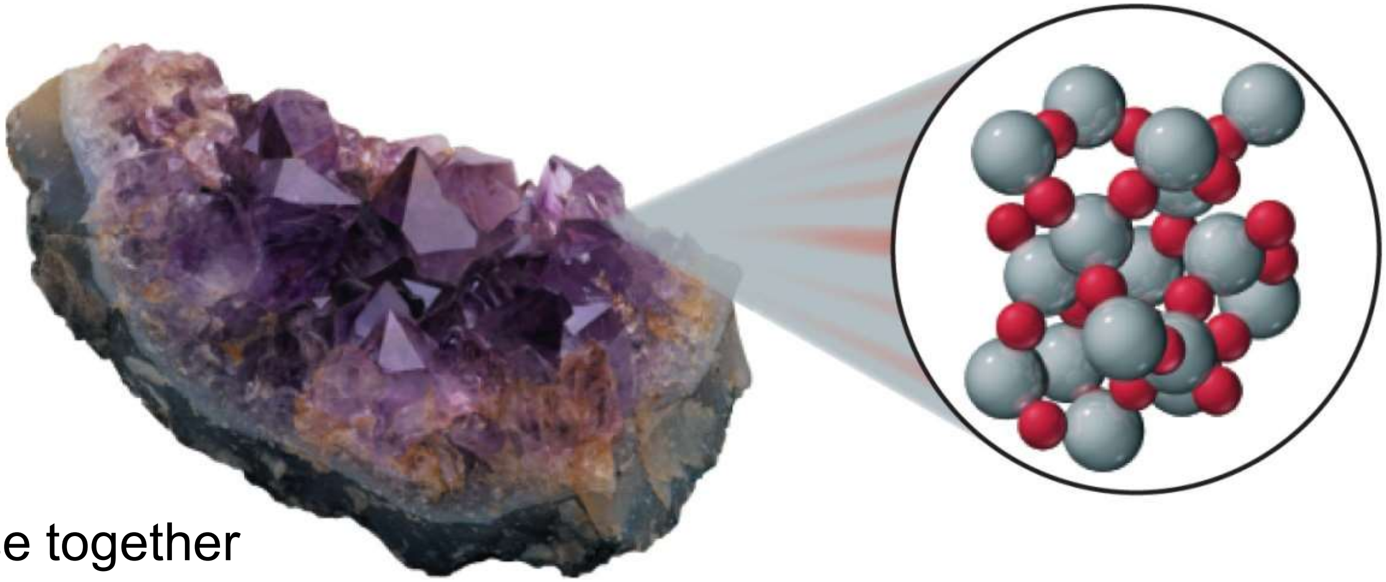


- States of matter
- Composition of matter
- Examples
- The three states of matter are
  1) **solid.**
  2) **liquid.**
  3) **gas.**

Solids

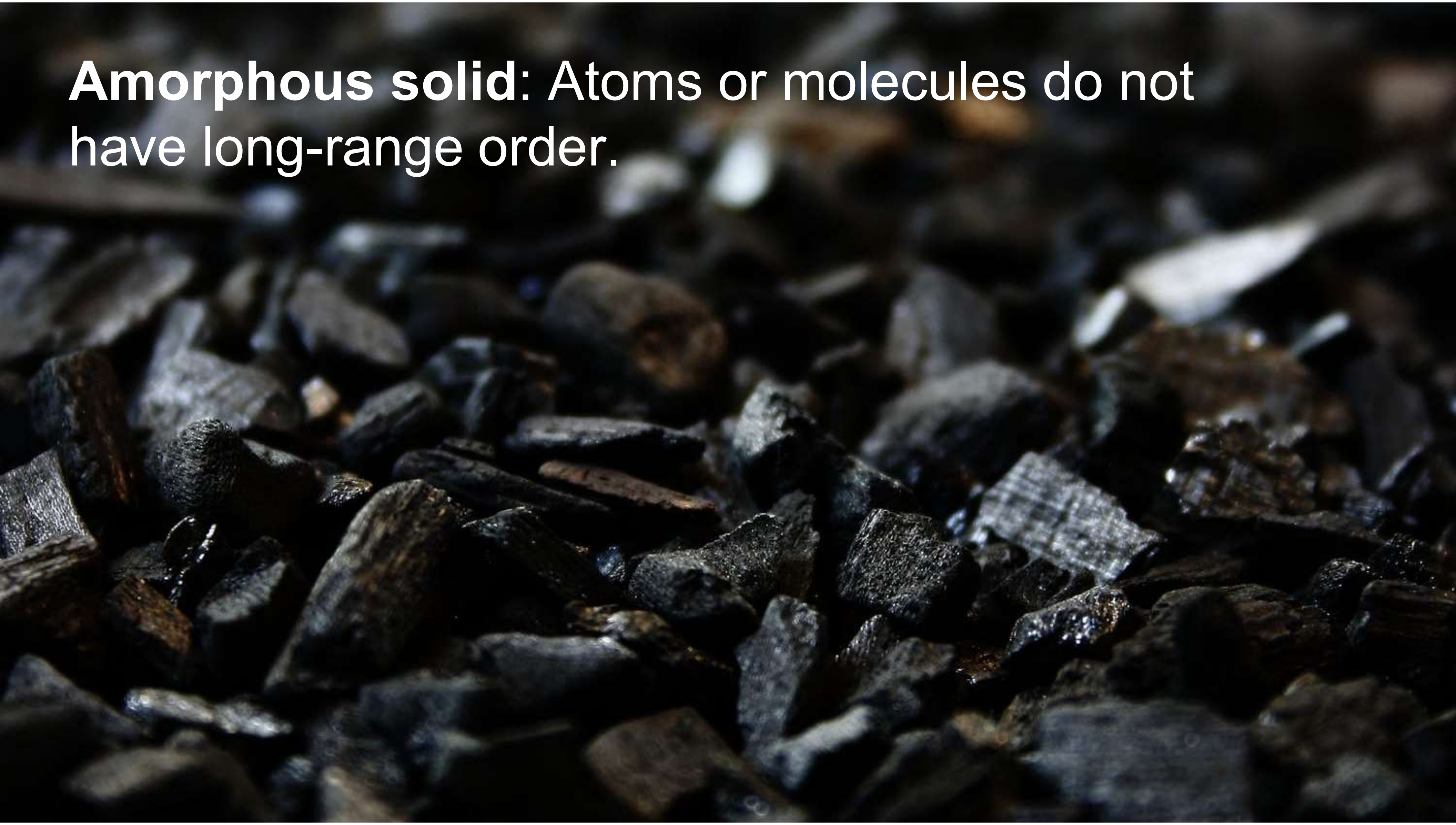
Solids have

- a definite shape
- a definite volume
- particles that are close together in a fixed arrangement
- particles that move very slowly
- Two types of solids (next slides)



Amethyst, a solid, is a purple form of quartz (SiO_2).

Crystalline solid: Atoms or molecules are arranged in geometric patterns with long-range, repeating order.

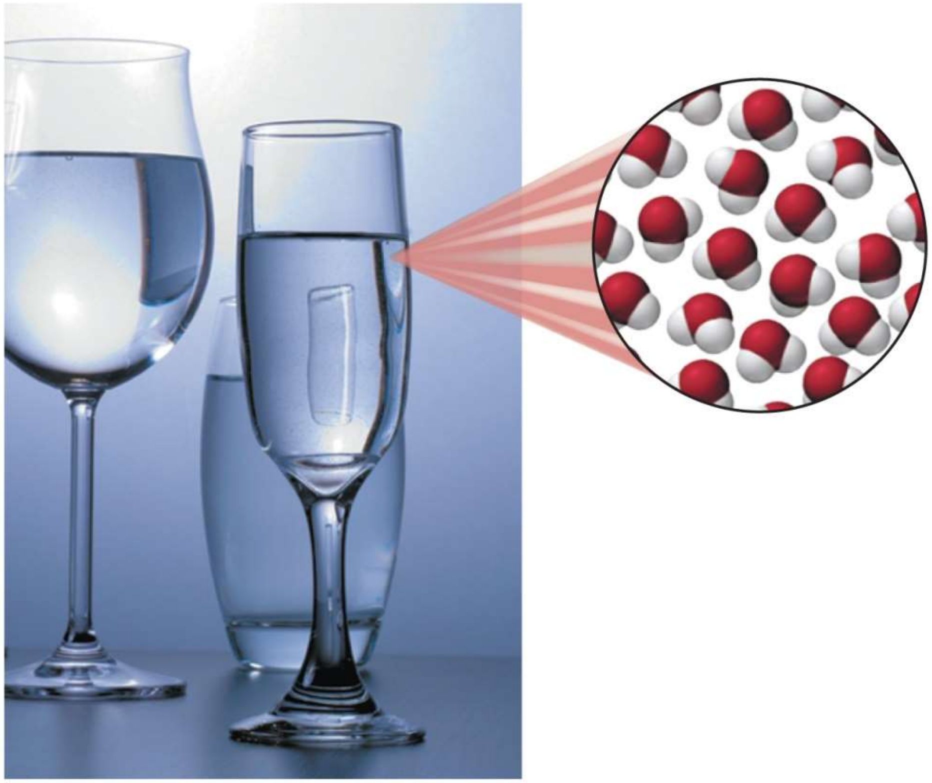

Amorphous solid: Atoms or molecules do not have long-range order.

Liquids

Liquids have

- an indefinite shape but a definite volume
- the same shape as their container
- particles that are close together but mobile
- particles that move slowly

A liquid has a definite volume but takes the shape of its container.



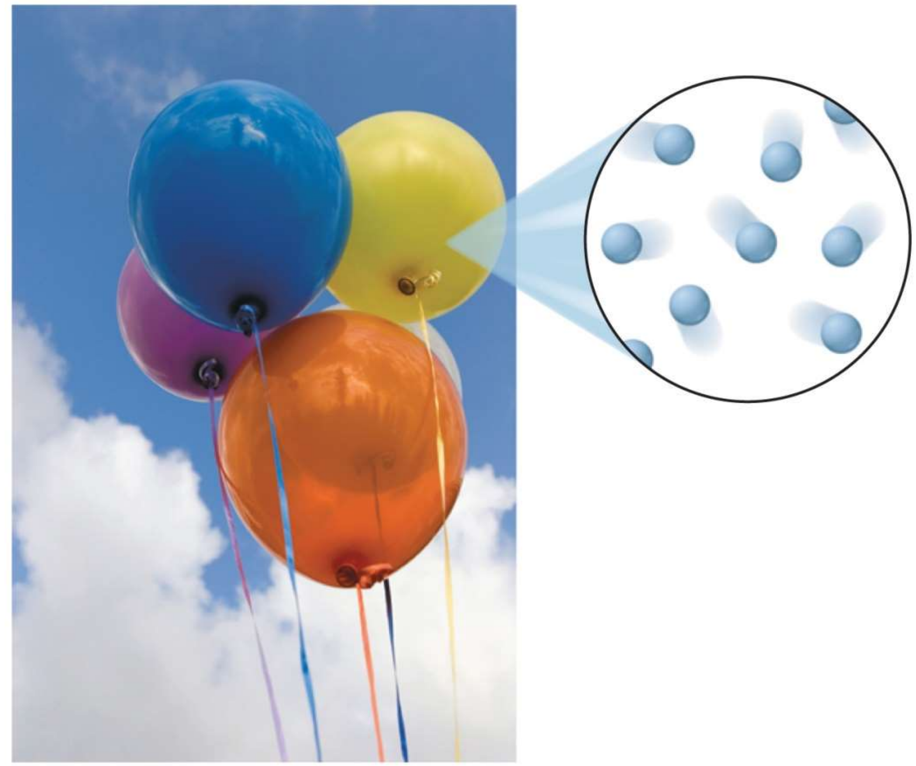
IMPORTANT - liquid takes the shape of its container.

Gases

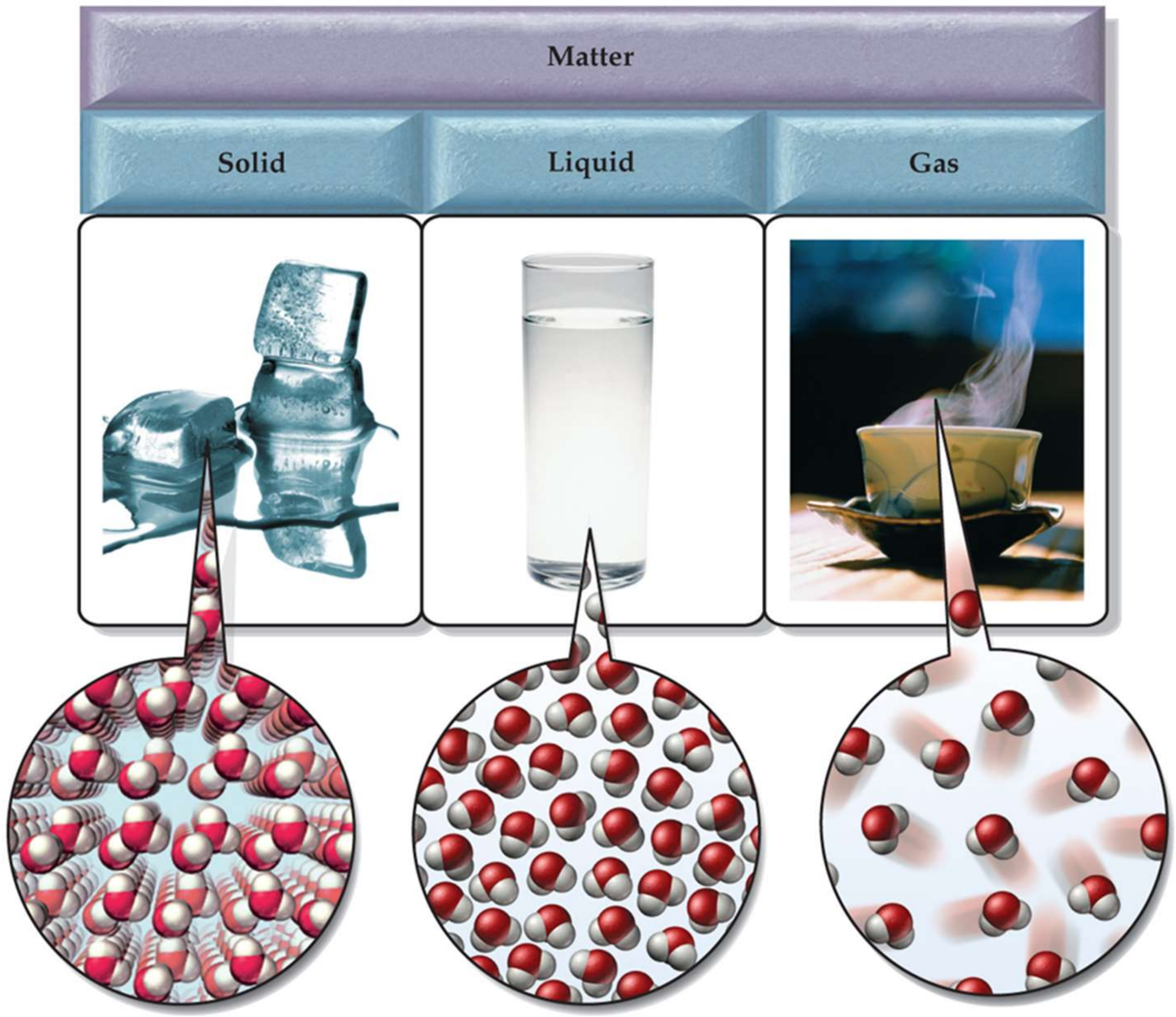
Gases have

- an indefinite shape
- an indefinite volume
- the same shape and volume as their container
- particles that are far apart
- particles that move very fast

A gas takes the shape and volume of its container.



IMPORTANT *- A gas takes the shape and volume of its*



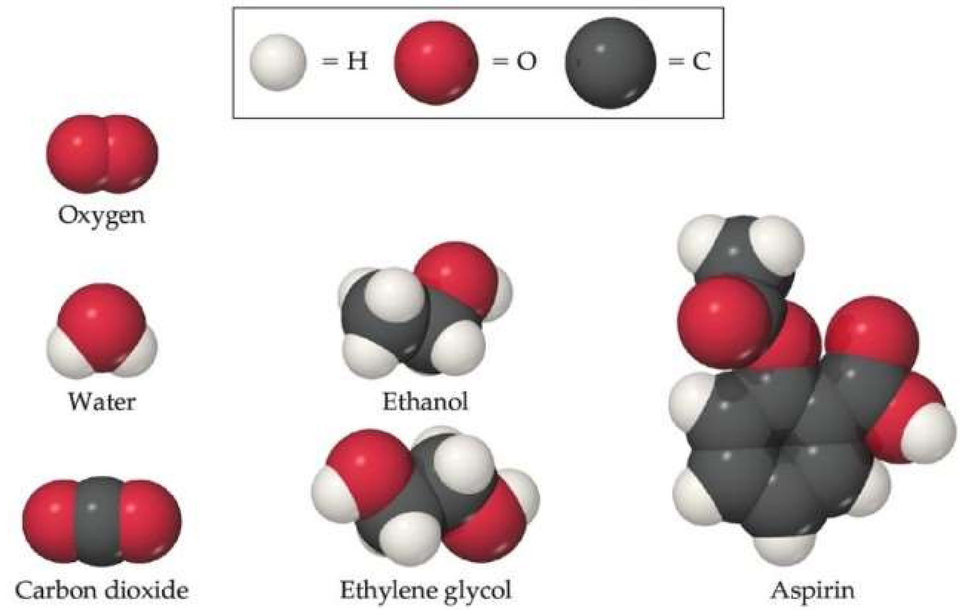
Matter
Solid
Liquid
Gas

Classification of Matter as Substances

- A **substance** has distinct properties and a composition that does not vary from sample to sample.
- The two types of substances are **elements** and **compounds**.
  - An **element** is a substance which can **not** be decomposed to simpler substances.
  - A **compound** is a substance which **can** be decomposed to simpler substances because it is made up of more than one element.

Classification of Matter Based on Composition

- **Atoms** are the building blocks of matter.
- Each **element** is made of a unique kind of atom, but can be made of more than one atom of that kind.
- A **compound** is made of atoms from two or more different elements.



Note: Balls of different colors are used to represent atoms of different elements. Attached balls represent connections between atoms that are seen in nature. These groups of atoms are called **molecules**.

Representing Elements

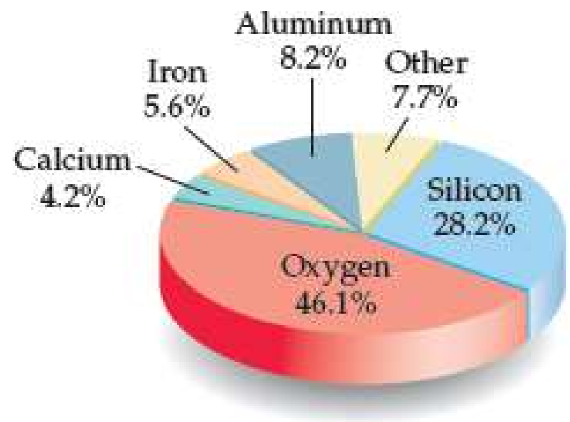
Table 1.1 Some Common Elements and Their Symbols

Carbon	C	Aluminum	Al	Copper	Cu (from **cuprum**)
Fluorine	F	Bromine	Br	Iron	Fe (from **ferrum**)
Hydrogen	H	Calcium	Ca	Lead	Pb (from **plumbum**)
Iodine	I	Chlorine	Cl	Mercury	Hg (from **hydrargyrum**)
Nitrogen	N	Helium	He	Potassium	K (from **kalium**)
Oxygen	O	Lithium	Li	Silver	Ag (from **argentum**)
Phosphorus	P	Magnesium	Mg	Sodium	Na (from **natrium**)
Sulfur	S	Silicon	Si	Tin	Sn (from **stannum**)

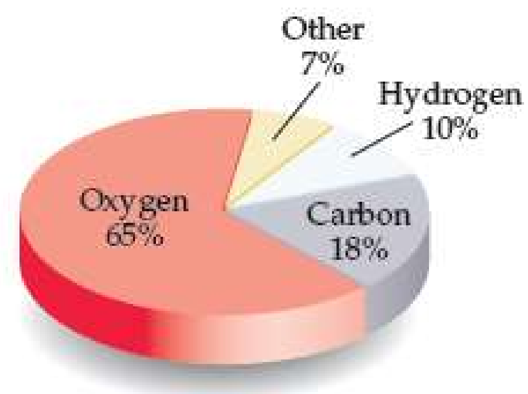
- Chemists usually represent elements as **symbols**.
- **Symbols** are one or two letters; the first is always capitalized.
- Some elements are based on Latin, Greek, or other foreign language names.

Elements and Composition

- There are currently 118 named elements.
- Only five elements make up 90% of the Earth's crust by mass.
- Only three elements make up 90% of the human body by mass.
- Note the importance of oxygen.

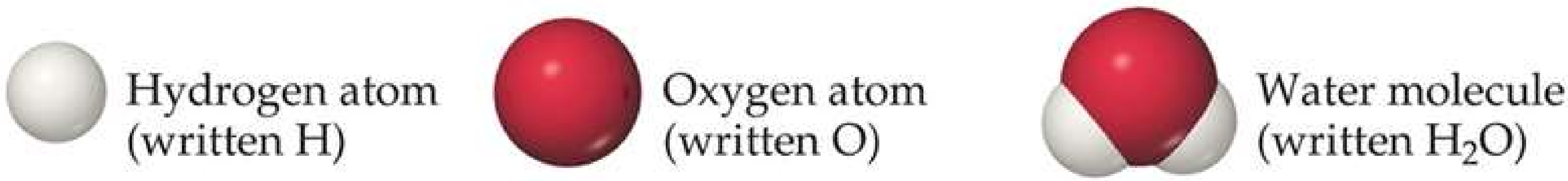


Earth's crust

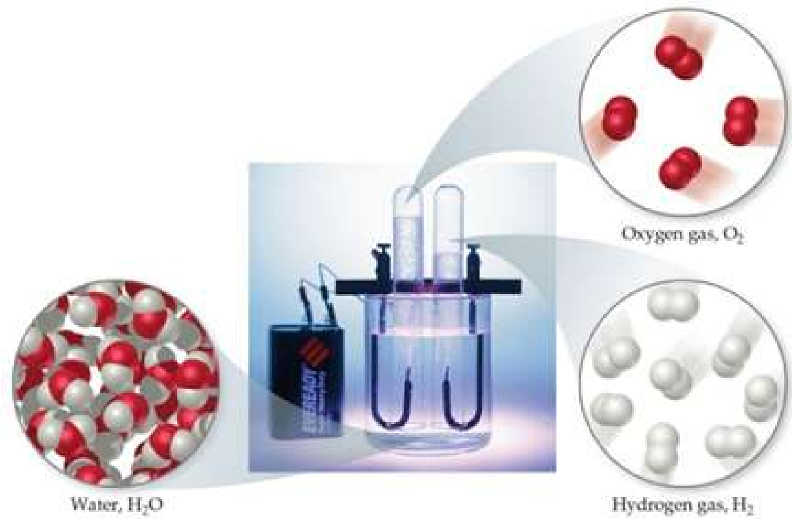


Human body

Compounds and Composition



- Compounds have a definite composition. That means that the relative number of atoms of each element in the compound is the same in any sample.
- This is the **Law of Constant Composition** (or the **Law of Definite Proportions**).



Mixtures

- **Mixtures** exhibit the properties of the substances that make them.
- Mixtures can vary in composition throughout a sample (**heterogeneous**) or can have the same composition throughout the sample (**homogeneous**).
- A homogeneous mixture is also called a **solution**.

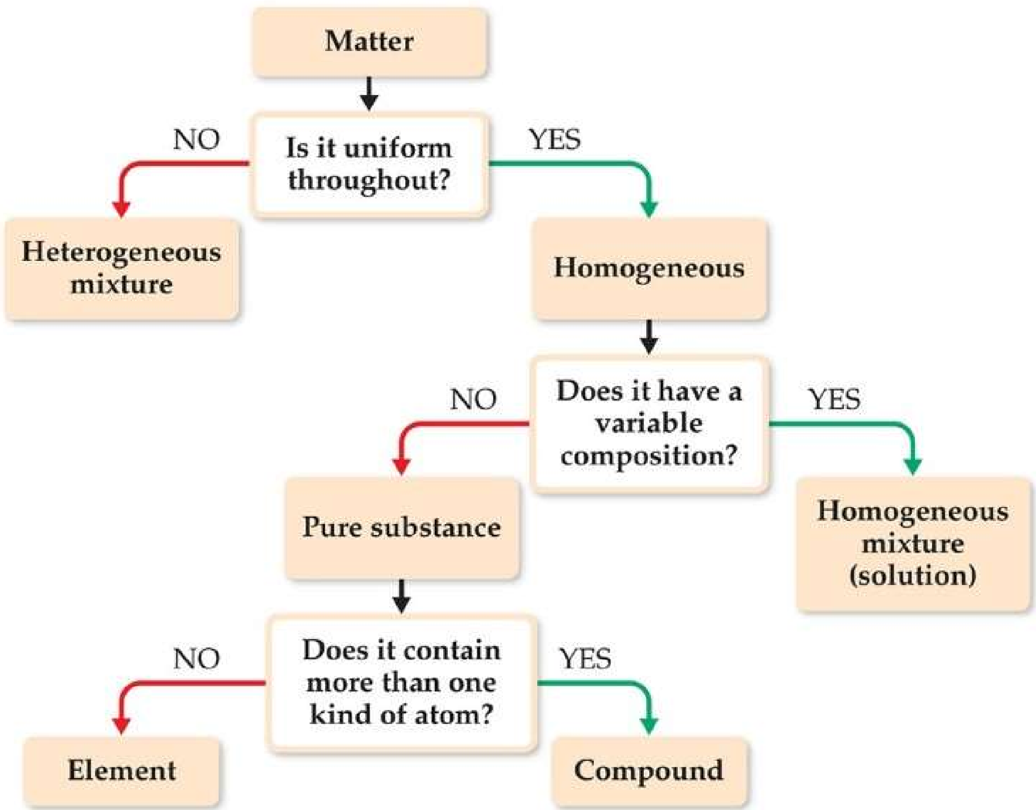
(a) (b)

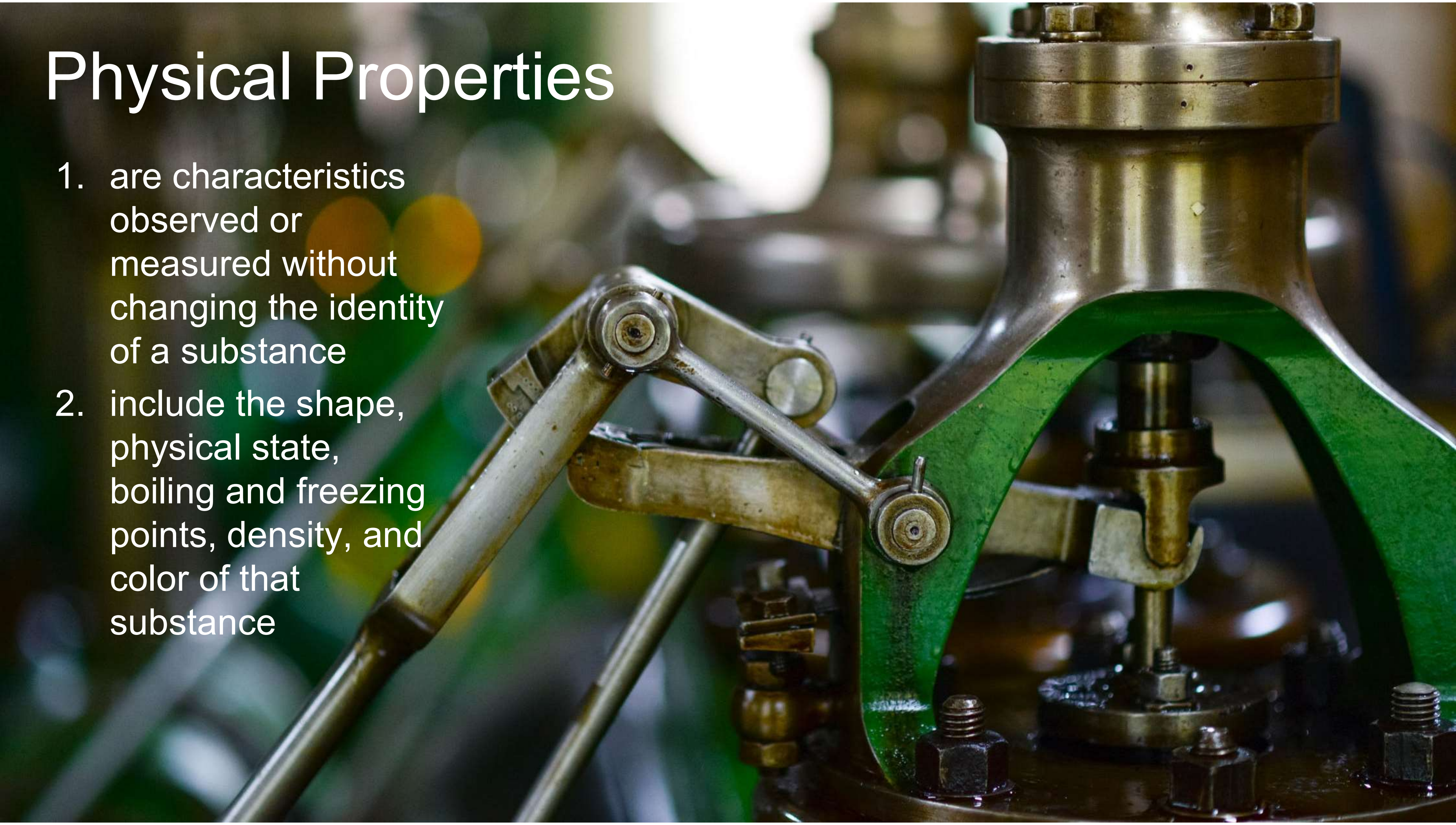
Separating Mixtures

- Mixtures can be separated based on physical properties of the components of the mixture. Some methods used are:
  - filtration
  - distillation
  - chromatography

Making a Decision

- If you follow this scheme, you can determine how to classify any type of matter:
  - Homogeneous mixture
  - Heterogeneous mixture
  - Element
  - Compound

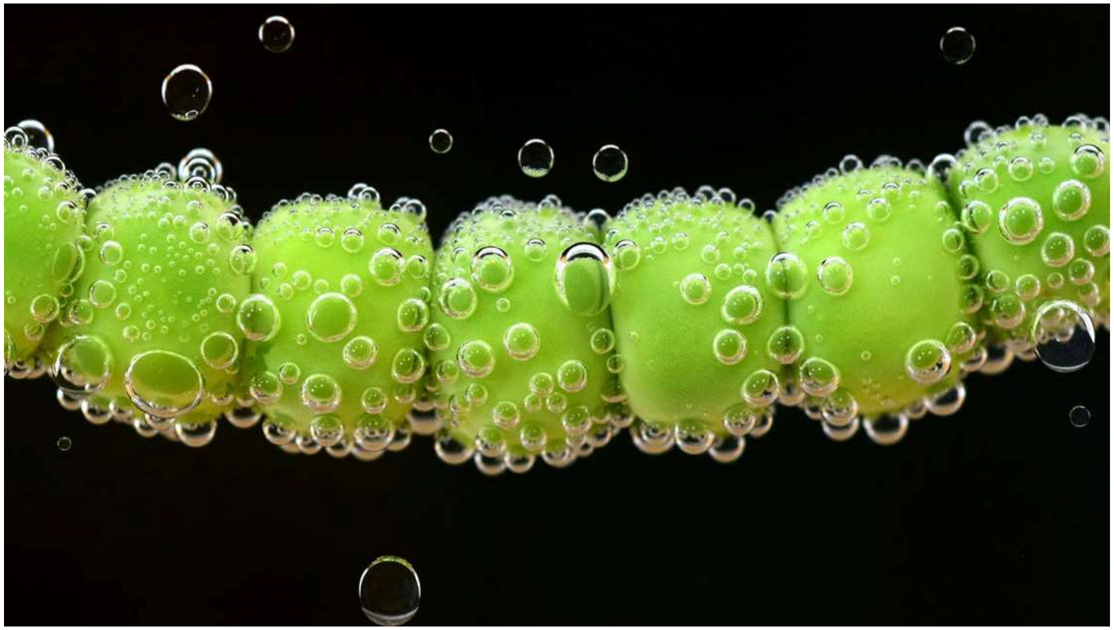




Physical Properties

1. are characteristics observed or measured without changing the identity of a substance
2. include the shape, physical state, boiling and freezing points, density, and color of that substance

Physical Change



A **physical change** occurs in a substance if there is

- a change in the state
- a change in the physical shape
- no change in the identity and composition of the substance

Examples of Physical Changes

Water boils to form water vapor.
Sugar dissolves in water to form a solution.
Copper is drawn into thin copper wires.
Paper is cut into tiny pieces of confetti.
Pepper is ground into flakes.

Chemical Properties and Changes

Chemical properties *describe the ability of a substance to interact with other substances to change into a new substance.*

When a ***chemical change*** *takes place, the original substance is turned into one or more new substances with new chemical and physical properties.*

$$\text{Reactants} \xrightarrow[\text{Chemical Change}]{} \text{Products}$$

Chemical change

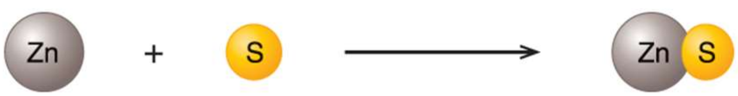

Chemical change

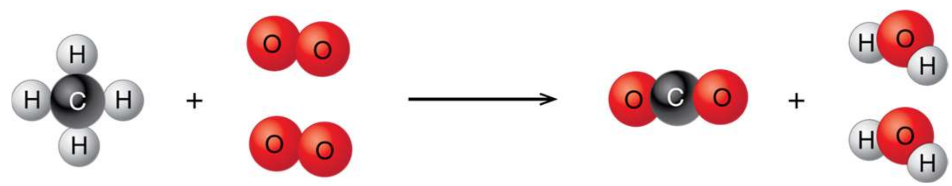
Examples of Chemical Changes

Shiny, silver metal reacts in air to give a black, grainy coating.
A piece of wood burns with a bright flame, and produces heat, ashes, carbon dioxide, and water vapor.
Heating white, granular sugar forms a smooth, caramel-colored substance.
Iron, which is gray and shiny, combines with oxygen to form orange-red rust.

Elements combine to form compounds: a chemical change.

Macmillan Learning

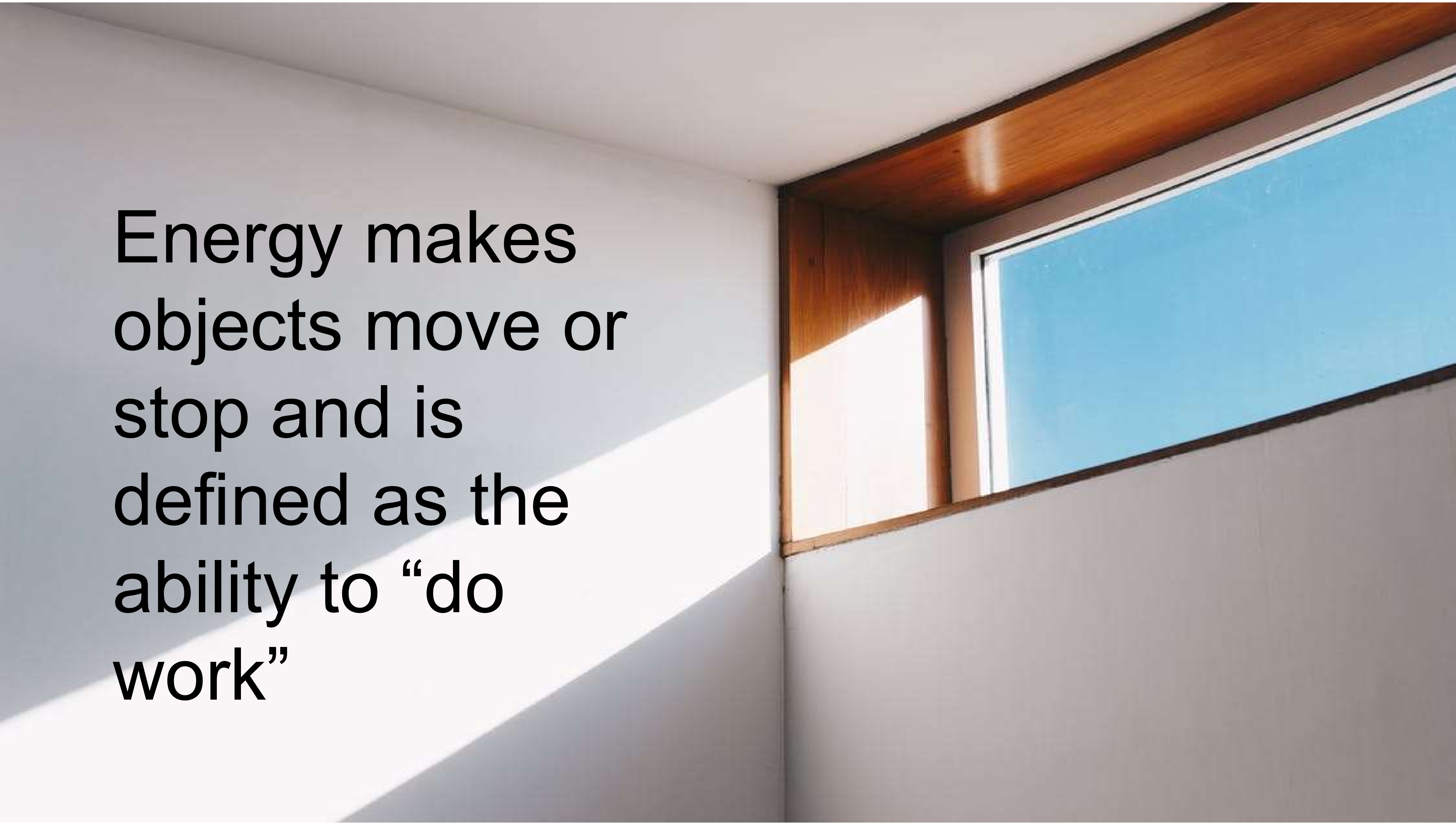




Property Type—Further Distinction

- **Intensive properties** are independent of the amount of the substance that is present.
  - Examples include density, boiling point, or color.
  - These are important for **identifying** a substance.
- **Extensive properties** depend upon the amount of the substance present.
  - Examples include mass, volume, or energy.

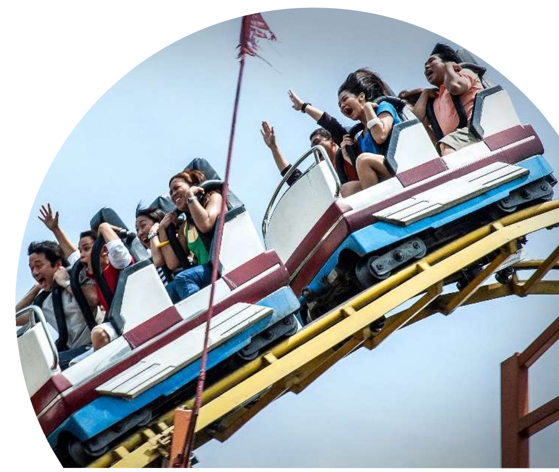
Energy



Energy makes objects move or stop and is defined as the ability to "do work"


Like mass, energy obeys
the law of conservation

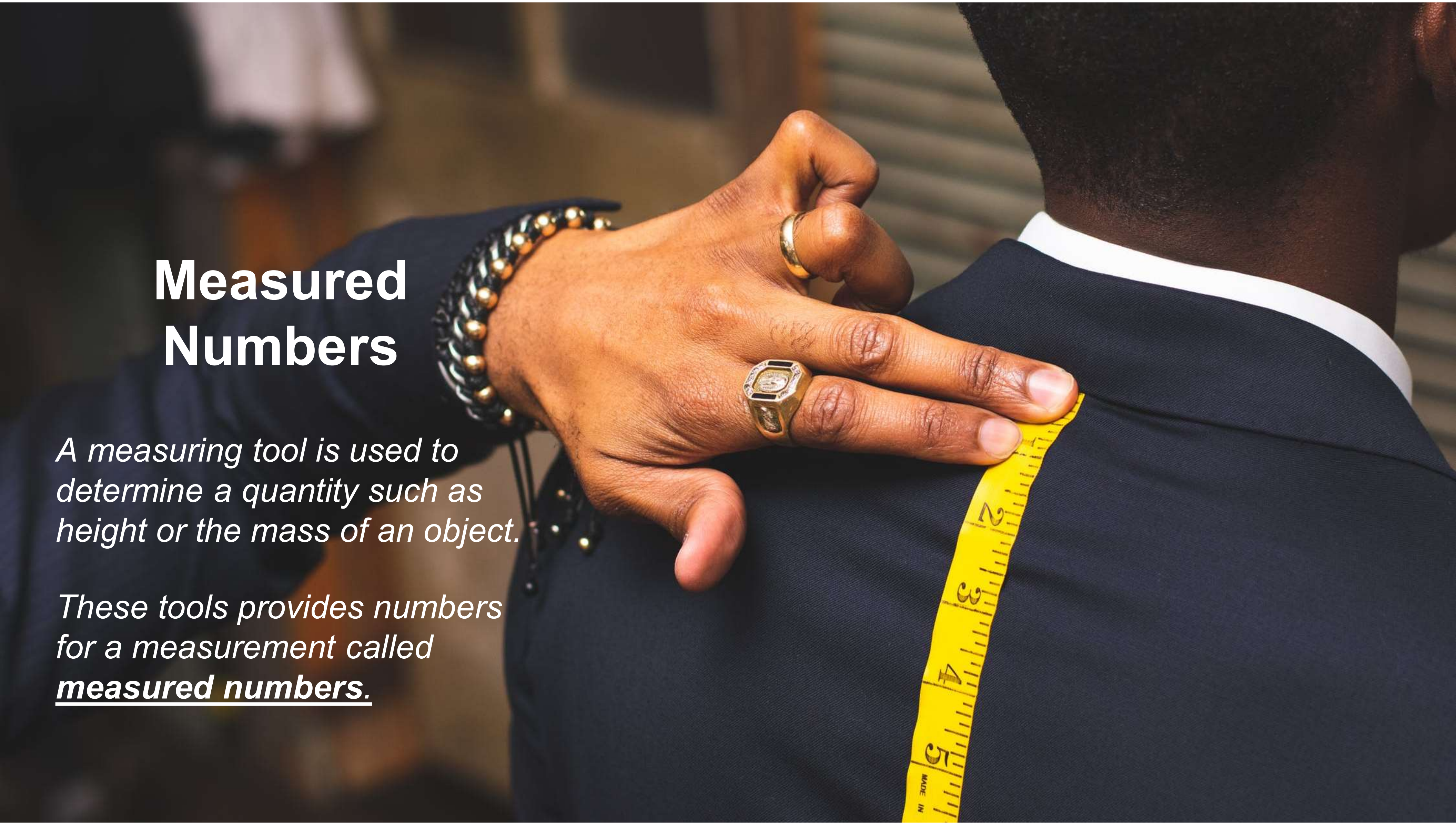
<u>Kinetic energy</u> is the energy of motion.



<u>Potential energy</u> is energy stored for use later.

Energy Units

- The unit of energy: Joule (J). It is a derived unit:
  - $KE = \frac{1}{2} m v^2$
  - If the object is 2 kg, and it moves at 1 m/s, it will posses 1 J of kinetic energy:
  - $1 J = \frac{1}{2}(2\ kg)(1\ m/s)^2 \quad OR: 1\ J \equiv 1\ kg \cdot m^2/s^2$
- The kJ is commonly used for chemical change.
- Historically, the calorie was used: 1 cal = 4.184 J.
- This calorie is not the nutritional calorie. That one is a kcal.
- 1 nutritional calorie = 1 cal = 1000 cal.



Measured Numbers

A measuring tool is used to determine a quantity such as height or the mass of an object.

These tools provides numbers for a measurement called ***measured numbers.***

Units of Measurement

Units, Continued

Fundamental Units

Measurement	Unit
Mass	kilogram (kg)
Length	meter (m)
Time	second (s)
Temperature	kelvin (K)
Light Intensity	candela (cd)
Electric current	ampere (A)
Amount	mole (mol)

Derived Units

Measurement	Units
Volume	m^3
Velocity	m/s
Density	kg/m^3

In order to facilitate communication, the international scientific community has developed an accepted set of 7 fundamental units of measurement.

From these, a host of ***derived units*** can be produced.

Units

Measurement	Metric Unit	English Unit	Relationship
Length	meter (m)	foot (ft) mile (mi)	1 m = 3.280 ft 1 km = 0.621 mi
Mass or Weight	kilogram (kg)	pound (lb)	1 kg = 2.204 lb
Volume	liter (L)	gallon (gal)	1 liter = 0.264 gal

In the United States, people commonly use English units, such as feet and miles for length, pounds for weight, and gallons for volume. Most of the world uses a different system, called the metric system.

Prefixes

Prefix	Symbol	Numerical Value	Scientific Notation	Equality
Prefixes That Increase the Size of the Unit				
tera	T	1 000 000 000 000	10^{12}	$1\ \text{Tg} = 1 \times 10^{12}\ \text{g}$
giga	G	1 000 000 000	10^{9}	$1\ \text{Gm} = 1 \times 10^{9}\ \text{m}$
mega	M	1 000 000	10^{6}	$1\ \text{Mg} = 1 \times 10^{6}\ \text{g}$
kilo	k	1 000	10^{3}	$1\ \text{km} = 1 \times 10^{3}\ \text{m}$
Prefixes That Decrease the Size of the Unit				
deci	d	0.1	10^{-1}	$1\ \text{dL} = 1 \times 10^{-1}\ \text{L}$ $1\ \text{L} = 10\ \text{dL}$
centi	c	0.01	10^{-2}	$1\ \text{cm} = 1 \times 10^{-2}\ \text{m}$ $1\ \text{m} = 100\ \text{cm}$
milli	m	0.001	10^{-3}	$1\ \text{ms} = 1 \times 10^{-3}\ \text{s}$ $1\ \text{s} = 1 \times 10^{3}\ \text{ms}$
micro	μ	0.000 001	10^{-6}	$1\ \mu\text{g} = 1 \times 10^{-6}\ \text{g}$ $1\ \text{g} = 1 \times 10^{6}\ \mu\text{g}$
nano	n	0.000 000 001	10^{-9}	$1\ \text{nm} = 1 \times 10^{-9}\ \text{m}$ $1\ \text{m} = 1 \times 10^{9}\ \text{nm}$
pico	p	0.000 000 000 001	10^{-12}	$1\ \text{ps} = 1 \times 10^{-12}\ \text{s}$ $1\ \text{s} = 1 \times 10^{12}\ \text{ps}$

Using Common Metric Prefixes

1. *How many meters are in a kilometer?*

$$1 \text{ km} = 1{,}000 \text{ m}$$

2. *How many A are in a MA?*

$$1 \text{ MA} = 1{,}000{,}000 \text{ A}$$

3. *How many mg are in a g?*

$$1 \text{ mg} = \frac{1}{1{,}000} \text{ g}$$

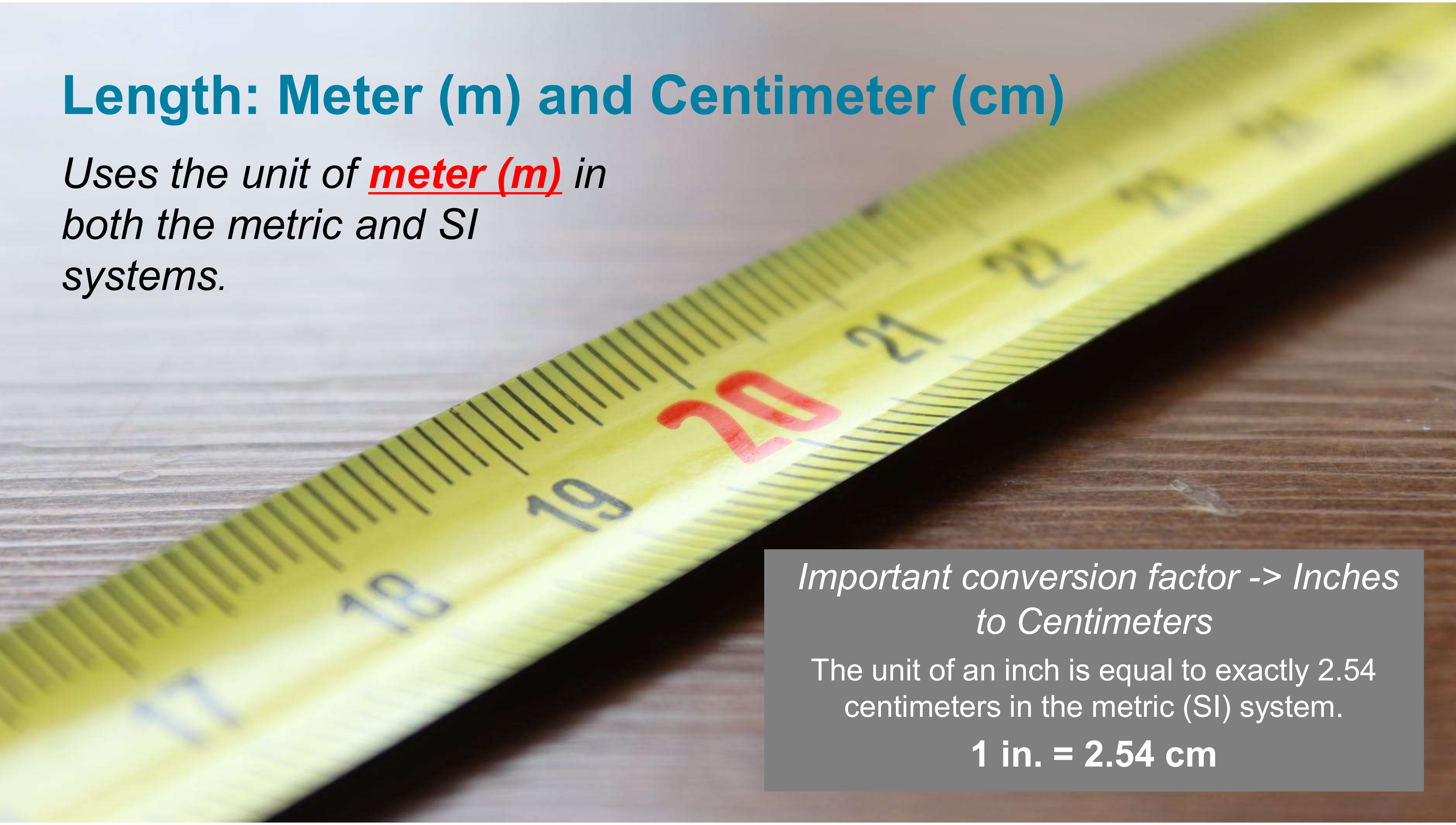
$$1{,}000 \text{ mg} = 1 \text{ g}$$

Table 2.5 Common Metric Prefixes

Prefix	Symbol	Meaning	
Mega-	M	10^6	1,000,000
Kilo-	k	10^3	1,000
Milli-	m	10^{-3}	$\frac{1}{1{,}000}$

Length: Meter (m) and Centimeter (cm)

Uses the unit of ***meter (m)*** *in both the metric and SI systems.*



Important conversion factor -> Inches to Centimeters

The unit of an inch is equal to exactly 2.54 centimeters in the metric (SI) system.

1 in. = 2.54 cm

Volume: Liter (L) and Milliliter (mL)

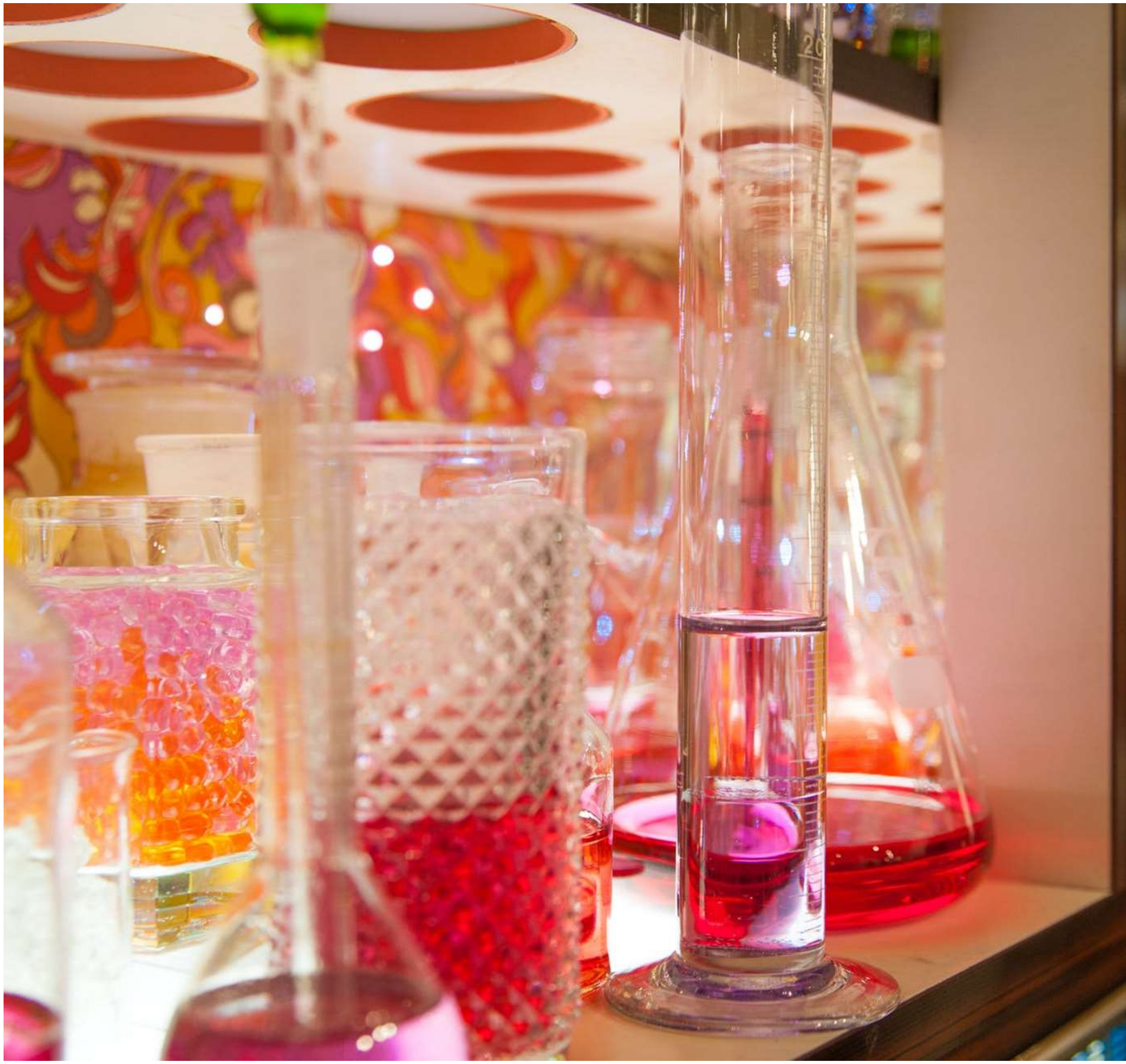
Is defined as the amount of space occupied by a substance.

*Si system is the unit **m³ (cubic meter).***

Metric** system is unit **liter (L).

Gases and liquids are measured by volume.

Important conversion factor
-> quart to liters
1 qt = 946 mL or 0.946 L



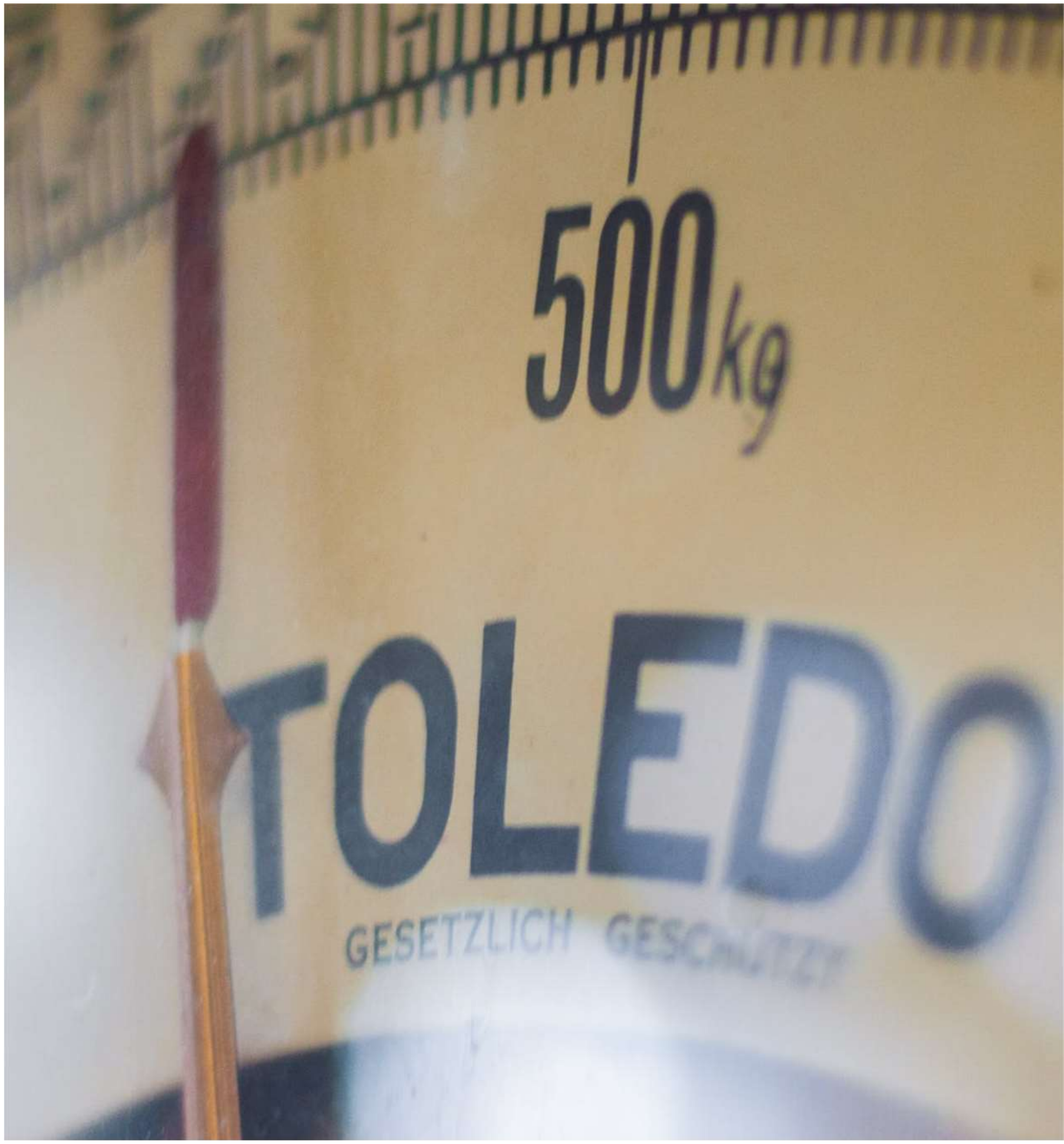
Mass: Gram (g) and Kilogram (kg)

*The **mass** of an object (different from weight)*

*Metric systems uses the unit **gram (g)***

*The SI system uses the unit **kilogram (kg)***

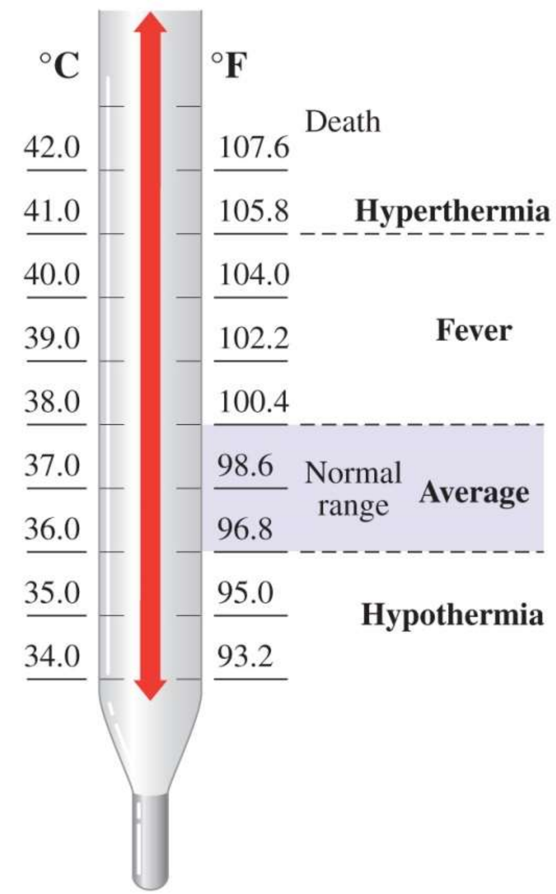
Important conversion factor
-> kilogram/gram to pound(s)
1 kg = 2.20 lb and 454 g = 1 lb

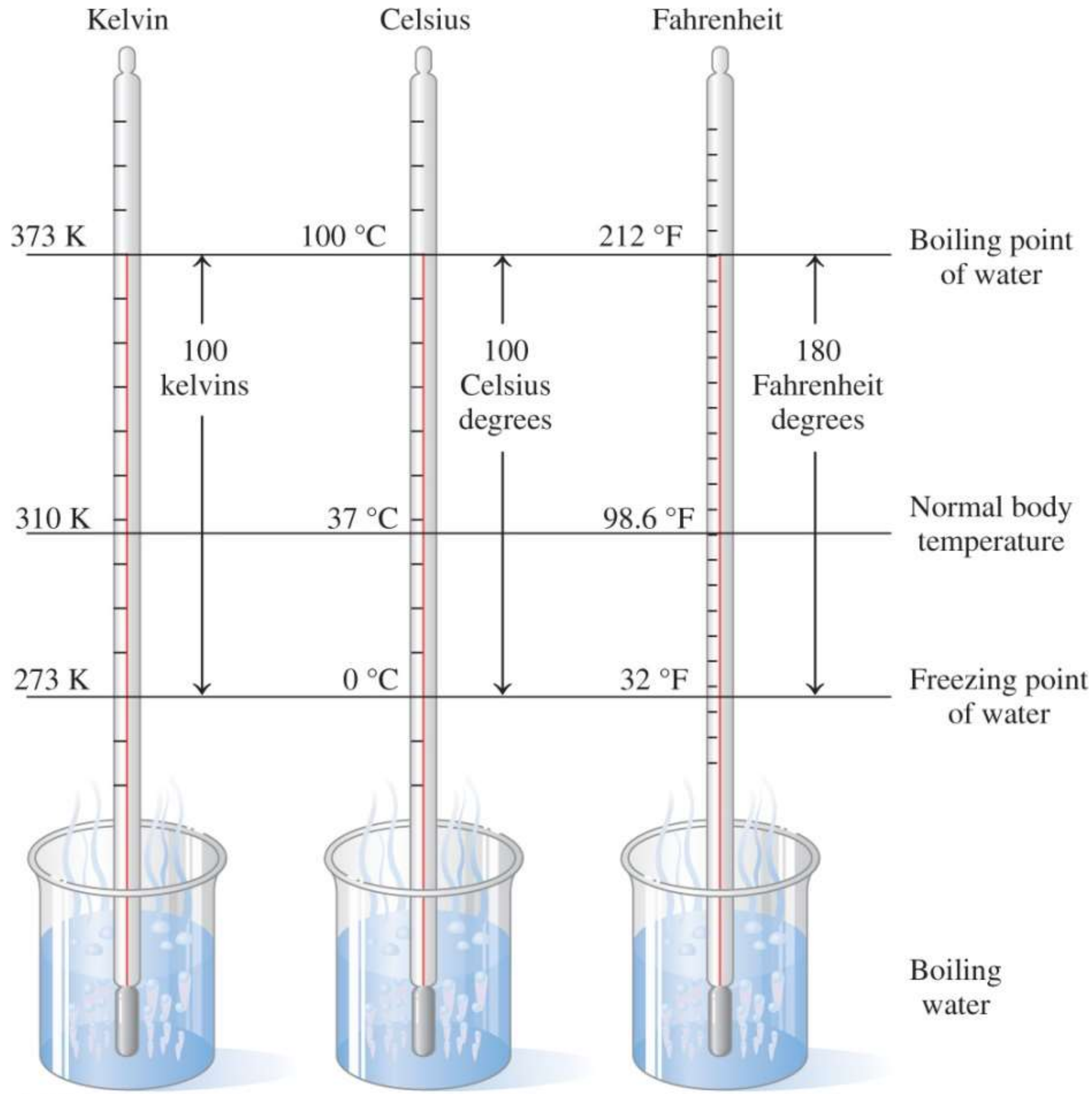


Temperature

Temperature is a measure of how hot or cold an object is compared to another object indicates the heat flow from higher temperature to lower temperature

- **The temperature scales**
  1. Fahrenheit
  2. Celsius
  3. Kelvin
- have reference points for the boiling and freezing points of water





Kelvin
Celsius
Fahrenheit
373 K
100 °C
212 °F
Boiling point of water
100 kelvins
100 Celsius degrees
180 Fahrenheit degrees
310 K
37 °C
98.6 °F
Normal body temperature
273 K
0 °C
32 °F
Freezing point of water
Boiling water

Fahrenheit and Celsius Scales

On the Celsius scale, there are 100 degrees Celsius between the freezing and boiling points of water.

On the Fahrenheit scale, there are 180 degrees Fahrenheit between the freezing and boiling points of water.

180 Fahrenheit degrees = 100 degrees Celsius

$$\frac{180 \text{ Fahrenheit degrees}}{100 \text{ degrees Celsius}} = \frac{1.8\ {}^{\circ}\mathrm{F}}{1\ {}^{\circ}\mathrm{C}}$$

Converting between Degrees Celsius and Degrees Fahrenheit

We can write a temperature equation to convert between Fahrenheit and Celsius temperatures.

$$T_F = 1.8(T_C) + 32 \quad \text{or} \quad T_F = 1.8(T_C) + 32$$

Changes °C to °F

Adjusts freezing point

Temperature equation to obtain degrees Fahrenheit

$$\frac{T_F - 32}{1.8} = \frac{\cancel{1.8}(T_C)}{\cancel{1.8}}$$

$$\frac{T_F - 32}{1.8} = T_C$$

Temperature equation to obtain degrees Celsius

Kelvin Temperature Scale

Scientists have learned that the coldest temperature possible is −273 °C. On the **Kelvin** scale, this is called **absolute zero** and is represented as 0 K.

The Kelvin scale has

A. units called kelvins (K)

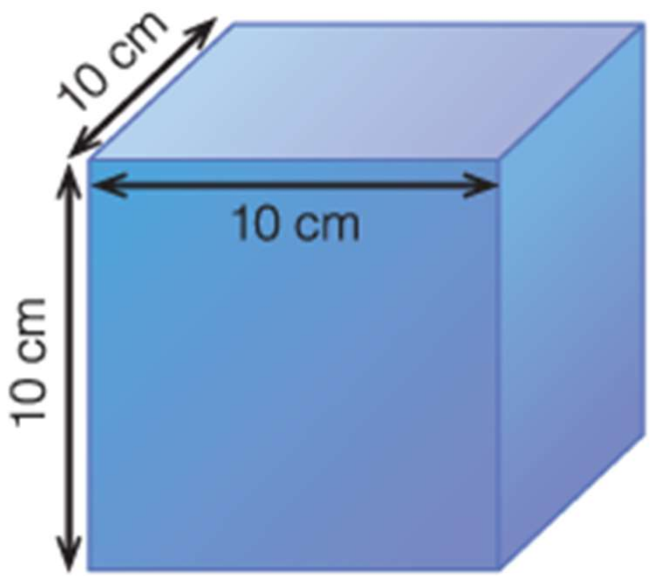
B. no degree symbol in front of K to represent temperature

C. no negative temperatures

D. the same size units as Celsius 1 K = 1 °C

$$T_K = T_C + 273$$ Temperature equation to obtain kelvins

Units of Volume



$10\ cm \times 10\ cm \times 10\ cm = 1{,}000\ cm^3$

Chemists commonly make measurements involving volume (that is, the space something occupies). To measure volume, we use units of length raised to the third power. For example, this cube has a length, width, and height of 10 cm . The volume of this cube is 1,000 centimeters cubed, also called “cubic centimeters”.

Volume Sizes

m^3

dm^3

cm^3

Top: Yuri_Arcurs/Getty Images; bottom left: naito29/Shutterstock; right: Blend Images/Brand X Pictures/Getty Images

Volume Sizes, Continued

liter (L):

$1\ L = 1\ dm^3$

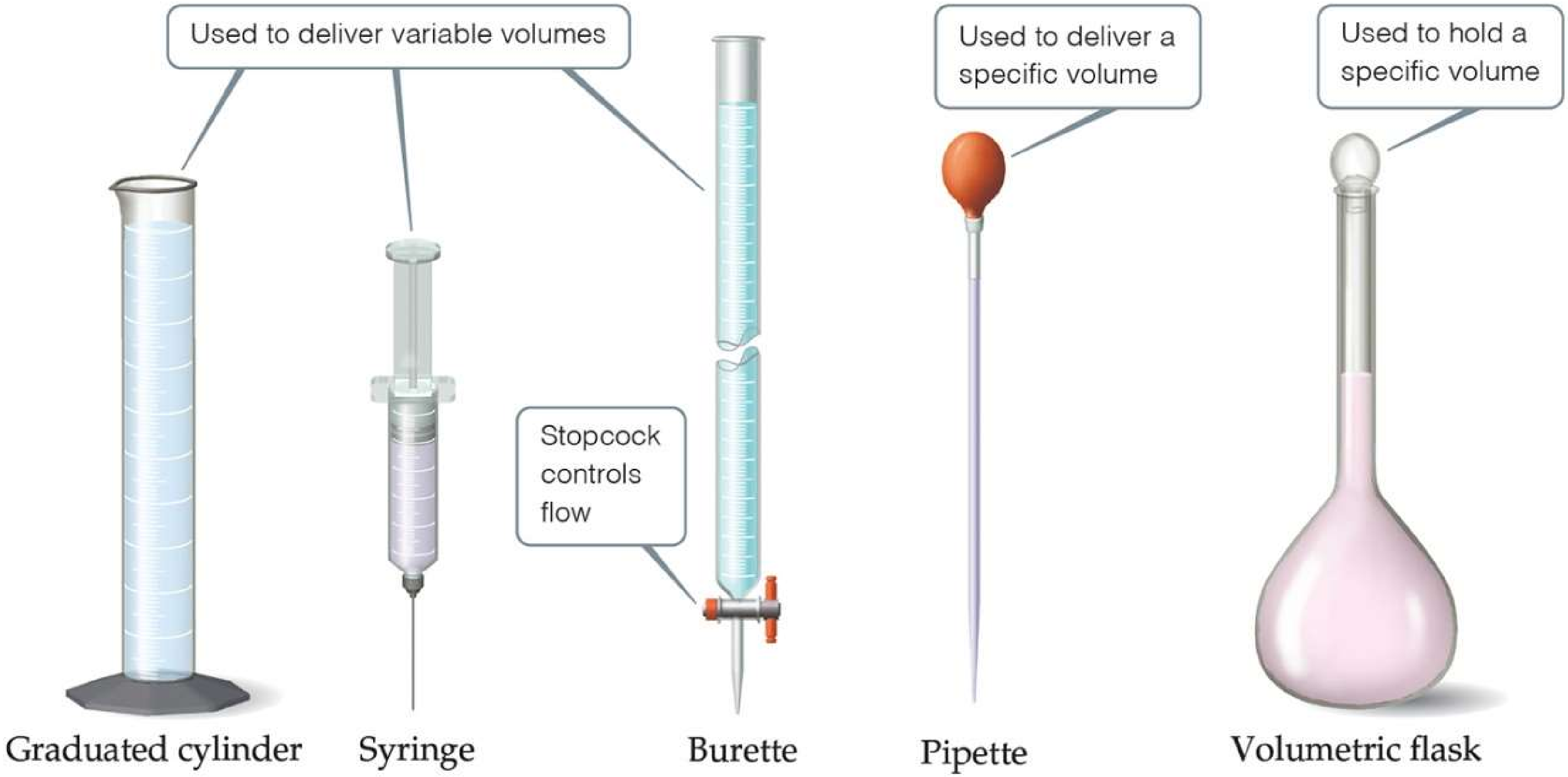
milliliter (mL)

1 milliliter = 1 cubic centimeter

$1\ mL = 1\ cm^3$

The liter is a common derived unit of volume. A liter is defined as one cubic decimeter, which is approximately the volume of a tissue box, or a large soft drink. Chemists often work with volumes of milliliters. One milliliter is the same volume as a cubic centimeter. A milliliter (or cubic centimeter) is about the same volume as a standard game die.

Glassware for Measuring Volume



Density

compares the mass of an object to its volume.

is the mass of a substance divided by its volume.

Density Expression

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{g}{mL} \text{ or } \frac{g}{cm^3}$$

Note: $1\ mL = 1\ cm^3$

Density

- Density is a physical property of a substance.
- It has units that are derived from the units for mass and volume.
- The most common units are g/mL or g/cm^3.

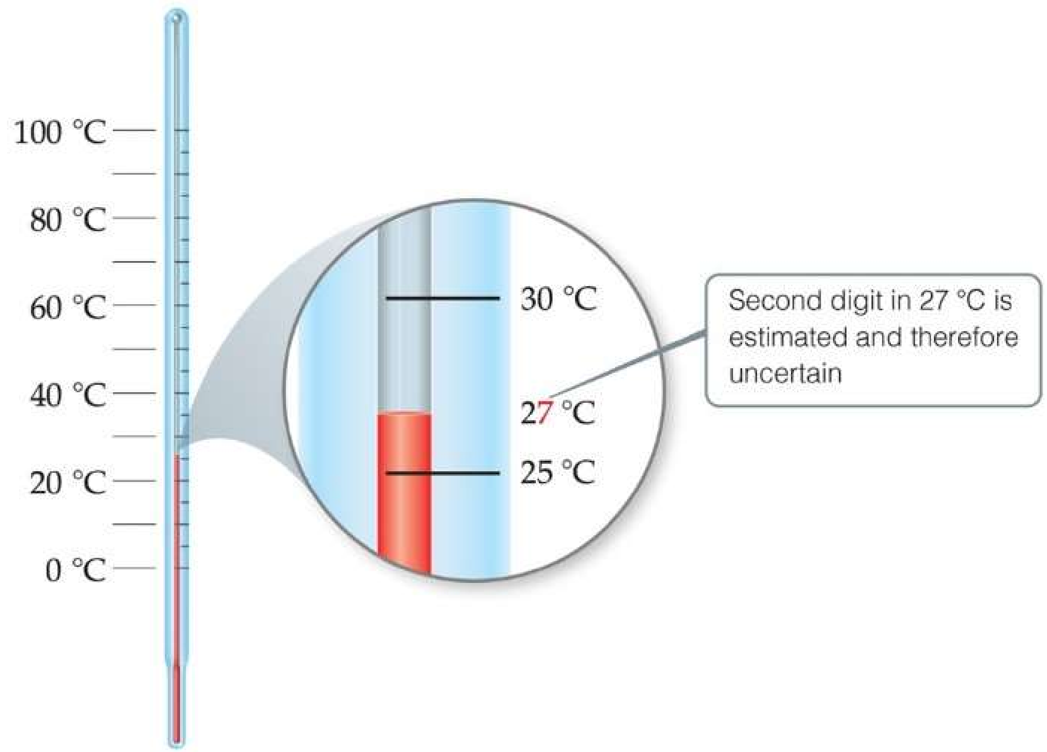
$$D = \frac{m}{V}$$

Table 1.5 Densities of Selected Substances at 25°C

Substance	**Density (g/cm^3)**
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

1.6 Uncertainty in Measurements

- Different measuring devices have different uses and different degrees of accuracy.
- All measured numbers have some degree of inaccuracy.
- The last digit measured is considered **reliable**, but **not exact**.



Numbers Encountered in Science

- **Exact** numbers are known exactly. They are counted or given by definition.
  - Count: there are 12 eggs in 1 dozen.
  - Define: 1 m = 100 cm or 1 kg = 2.2046 lb.
- **Inexact** (or **measured**) numbers depend on how they were determined. Scientific instruments have limitations (**equipment errors**) and individuals can read some instrumentation differently (**human errors**).
  - **Uncertainties always exist.**

Precision and Accuracy

Accuracy

- How reliable are the measurements?
- Do they reflect the true value?

Precision

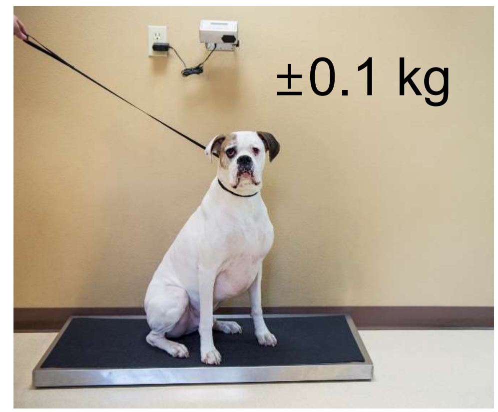
- How finely are the measurements made?
- How closely are they grouped together?

For example, we would say that the analytical balance is more precise than the animal scale. The balance measures to within 0.0001 gram, while the animal scale measures to plus or minus 0.1 kilogram.

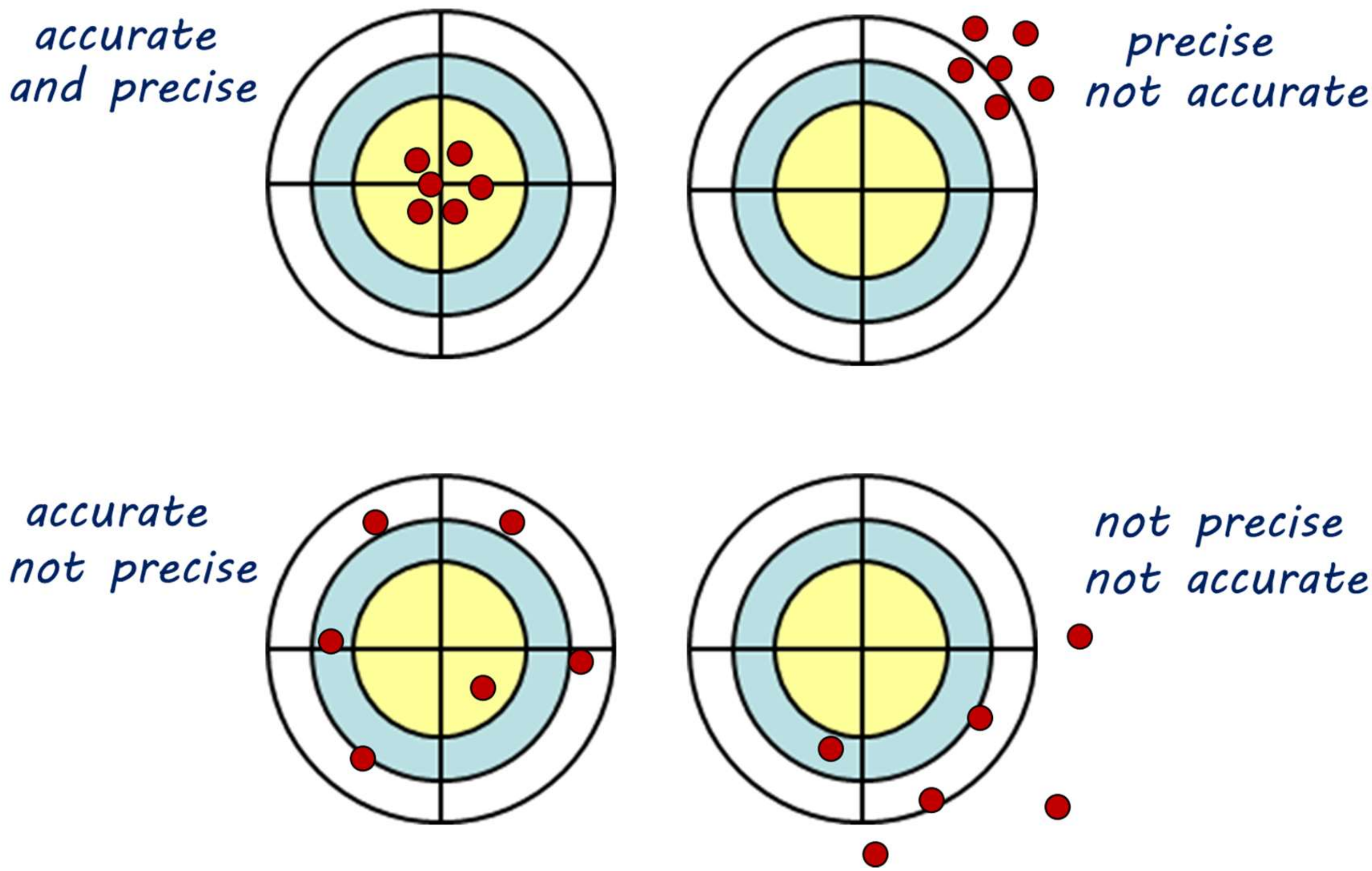
James A. Prince/Science Source

±0.0001 g

Jetta Productions/Getty Images



Precision and Accuracy, Continued

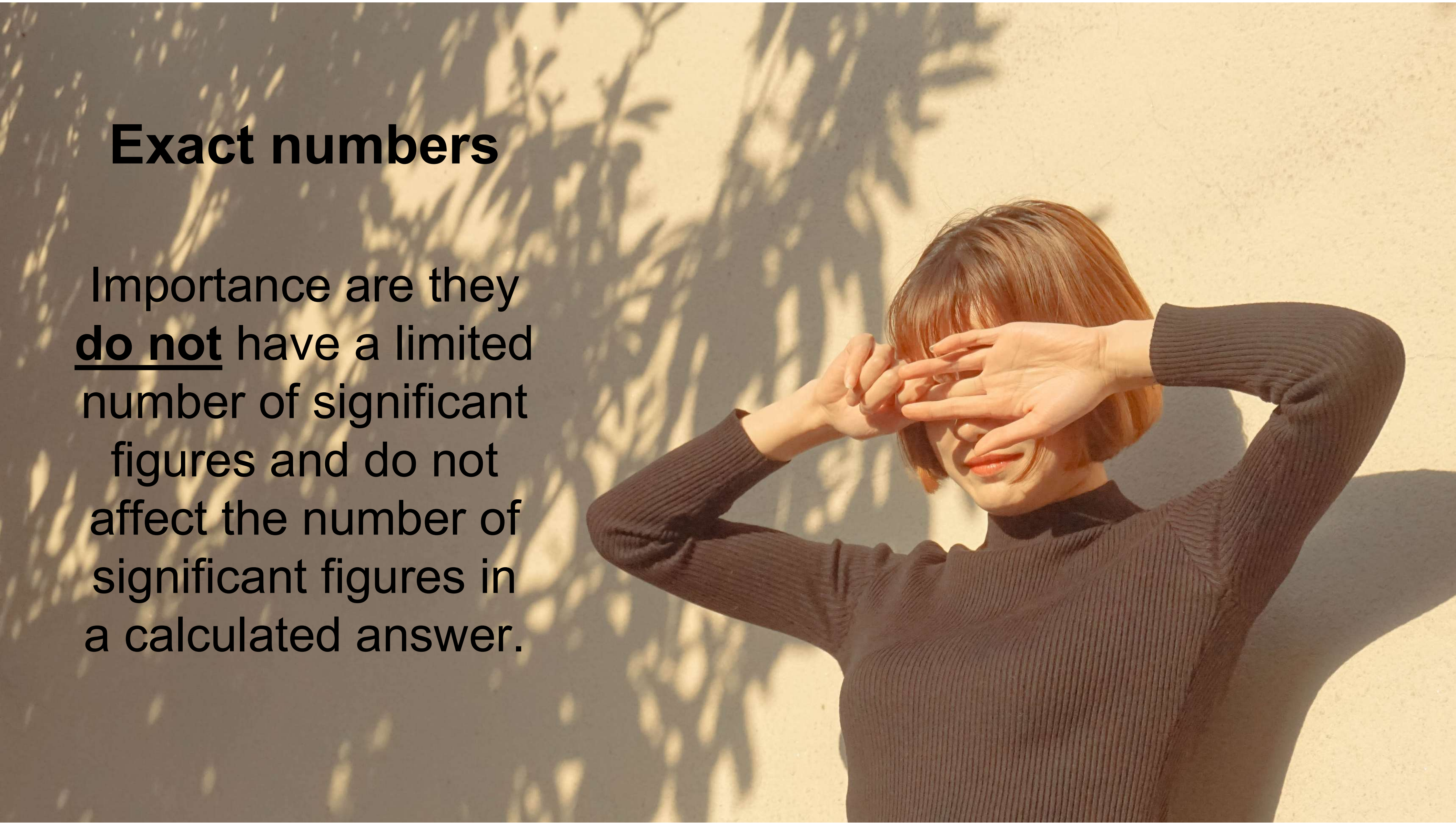


Significant Figures

- All digits of a measured quantity, including the uncertain ones, are called **significant figures**.
- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the accuracy of our answers.
- There is always uncertainty in the last digit reported for any measured quantity. If a balance measures to 0.0001 g$_{\text{ram}}$, mass is reported as 2.2405 ± 0.0001 g.

Rule	Measured Number	Number of Significant Figures
1. A number is a *significant figure* if it is		
a. not a zero	4.5 g	2
	122.35 m	5
b. a zero between nonzero digits	205 m	3
	5.082 kg	4
c. a zero at the end of a decimal number	50. L	2
	25.0 °C	3
	16.00 g	4
d. any digit in the coefficient of a number written in scientific notation	4.0×10^5 m	2
	5.70×10^{-3} g	3
2. A zero is *not significant* if it is		
a. at the beginning of a decimal number	0.0004 lb	1
	0.075 m	2
b. used as a placeholder in a large number without a decimal point	850 000 m	2
	1 250 000 g	3

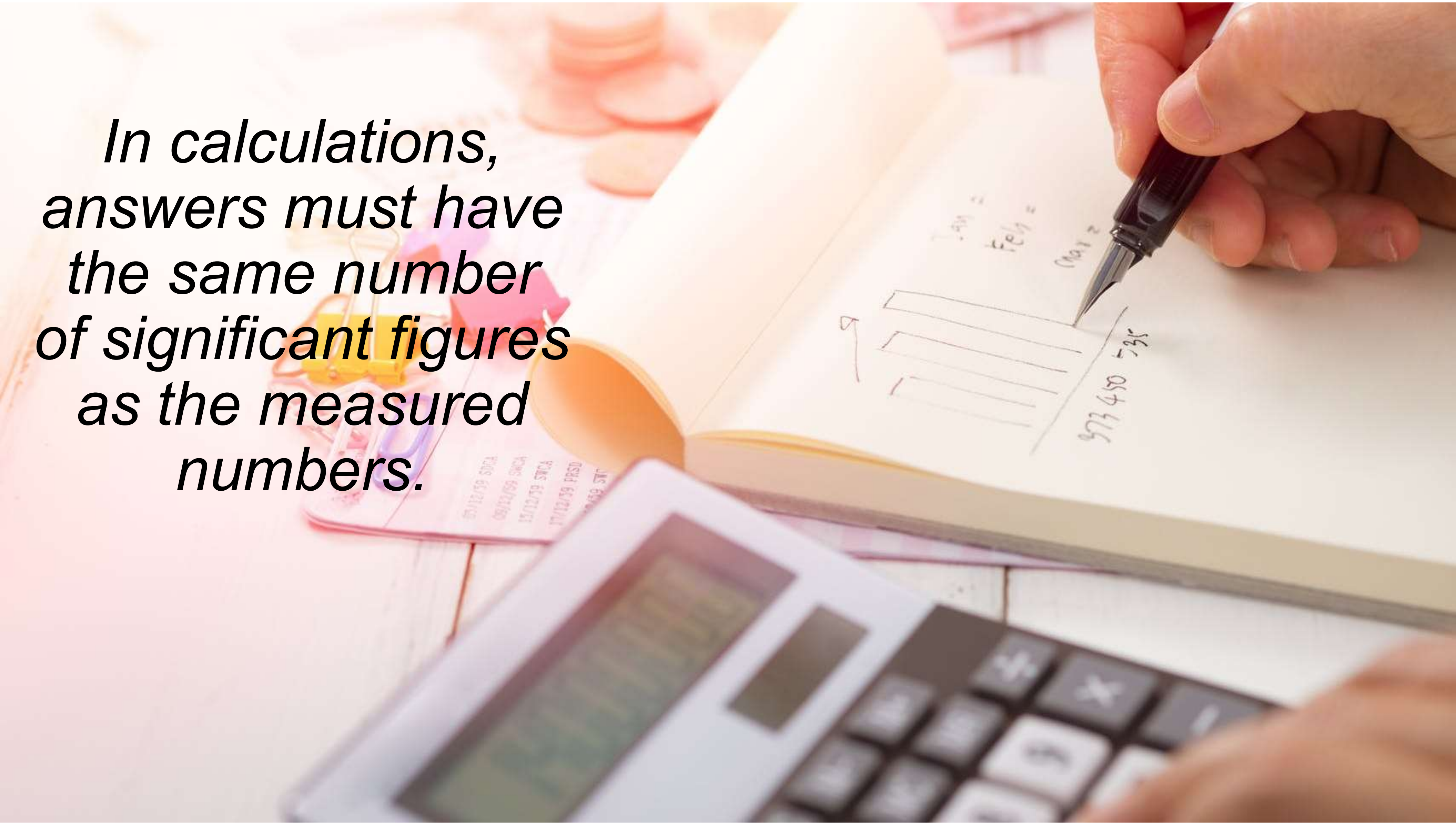
What about non-measured number?



Exact numbers

Importance are they **<u>do not</u>** have a limited number of significant figures and do not affect the number of significant figures in a calculated answer.

	Defined Equalities	
Counted Numbers	U.S. System	Metric System
8 doughnuts	1 ft = 12 in.	1 L = 1000 mL
2 baseballs	1 qt = 4 cups	1 m = 100 cm
5 capsules	1 lb = 16 ounces	1 kg = 1000 g


In calculations,
answers must have
the same number
of significant figures
as the measured
numbers.

Multiplication and Division with Significant Digits

1. When multiplying or dividing, report the same number of digits as are in the least precise **significant figure** (starting measurement.)

A vehicle travels 315.3 miles in the span of 5.2 hours. What is its average speed, in miles per hour?

4 sig. digits

$$\frac{315.3 \text{ miles}}{5.2 \text{ hours}} = 60.\cancel{63461538}$$

2 sig. digits

$$= 61 \text{ miles/hour}$$

Addition and Subtraction with Significant Digits

2. When adding or subtracting, round to the last **decimal place** of the least precise starting measurement.

While training for a triathlon, you swim 0.432 miles, then bike 18.1 miles. What was your total distance traveled?

$$
\begin{array}{lrl}
\text{Swim} & 0.432 & \text{mi.} \\
+\ \text{Bike} & 18.1 & \text{mi.} \\
\hline
 & = 18.5\cancel{32} & \text{mi.} \\
 & & \\
 & = 18.5 & \text{mi.}
\end{array}
$$

If a calculation involves multiple steps, wait until the end to round to significant digits.

Example with Significant Digits

A chemist measures the mass of chloride in three water samples, as shown in the table. Together, the three samples have a volume of 2.31 liters. What is the average mass of chloride per liter of water? Answer to significant digits.

Sample	Mass of Chloride
A	15.21 mg
B	9.33 mg
C	11.329 mg

total mass chloride:

$$\begin{array}{r} 15.21\ \text{mg} \\ 9.33\ \text{mg} \\ 11.329\ \text{mg} \\ \hline 35.869\ \text{mg} \end{array}$$

= 35.87 mg

4 sig. digits

$$\frac{\text{total mass}}{\text{volume}}$$

Use unrounded mass

$$= \frac{35.869\ \text{mg}}{2.31\ \text{L}}$$

4 sig. digits

3 sig. digits

$= 15.5\cancel{2770563}$

= 15.5 mg/L

A food company measures the number of kilocalories (kcal) of energy in a soda. They find that there are 1.50×10^2 kcal in a 350.0 mL sample. Report the energy content of this soda in kcal/mL, to the correct number of significant figures.

A. 4.28 kcal/mL

B. 4.29 kcal/mL

C. 0.428 kcal/mL

D. 0.429 kcal/mL

E. 4.286 kcal/mL

A food company measures the number of kilocalories (kcal) of energy in a soda. They find that there are 1.50×10^2 kcal in a 350.0 mL sample. Report the energy content of this soda in kcal/mL, to the correct number of significant figures. **(Answer)**

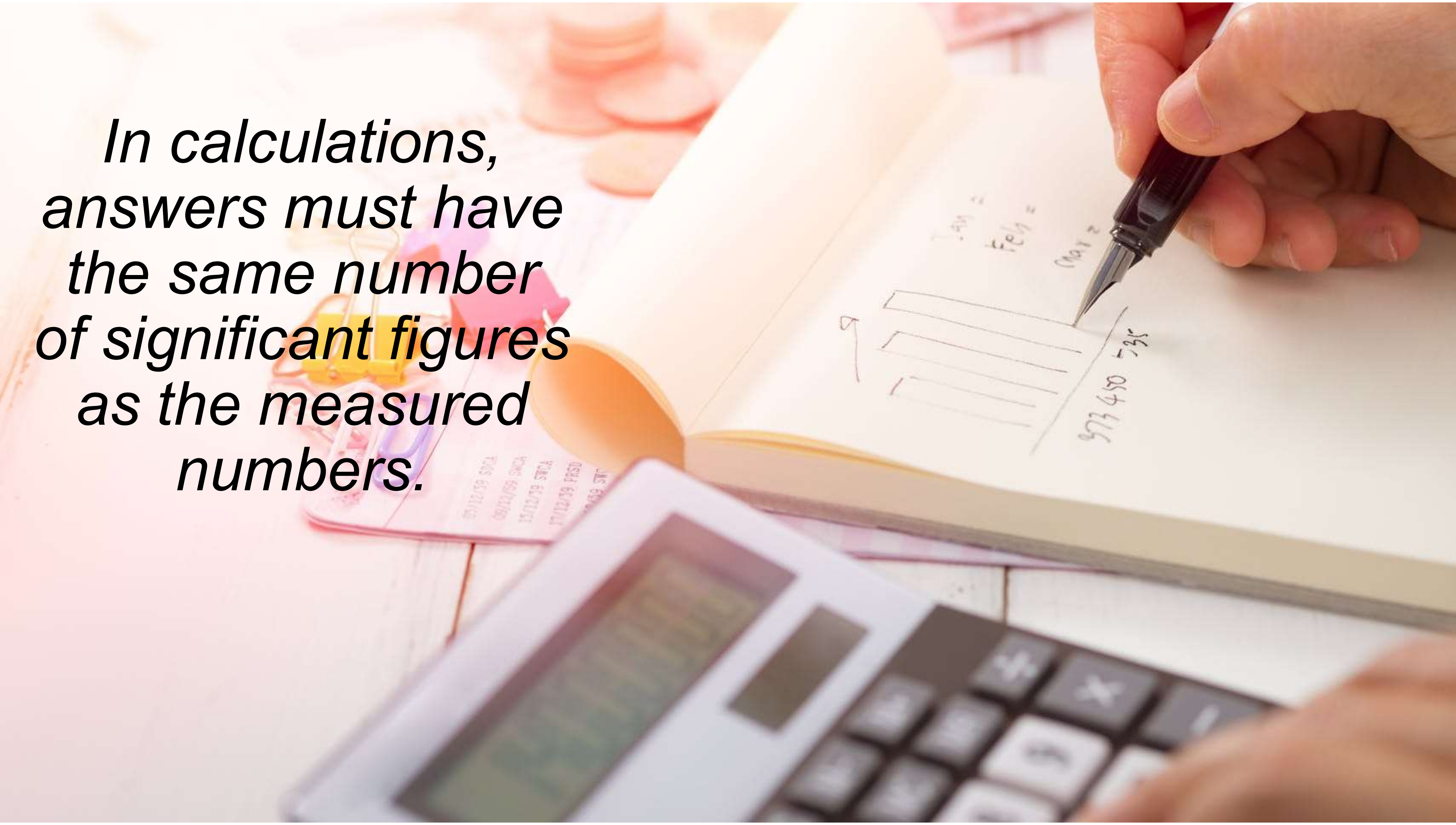
A. 4.28 kcal/mL

B. 4.29 kcal/mL (correct answer)

C. 0.428 kcal/mL

D. 0.429 kcal/mL

E. 4.286 kcal/mL



In calculations, answers must have the same number of significant figures as the measured numbers.

Multiplication and Division with Significant Digits

1. When multiplying or dividing, report the same number of digits as are in the least precise **significant figure** (starting measurement.)

A vehicle travels 315.3 miles in the span of 5.2 hours. What is its average speed, in miles per hour?

4 sig. digits

$$\frac{315.3 \text{ miles}}{5.2 \text{ hours}} = 60.\cancel{63461538}$$

2 sig. digits

$$= 61 \text{ miles/hour}$$

Addition and Subtraction with Significant Digits

2. When adding or subtracting, round to the last **decimal place** of the least precise starting measurement.

While training for a triathlon, you swim 0.432 miles, then bike 18.1 miles. What was your total distance traveled?

$$\begin{array}{lrl} & \text{Swim} \quad 0.432 & \text{mi.} \\ + & \text{Bike} \quad 18.1 & \text{mi.} \\ \hline & = 18.5\cancel{32} & \text{mi.} \end{array}$$

$$= 18.5 \quad \text{mi.}$$

If a calculation involves multiple steps, wait until the end to round to significant digits.

Example with Significant Digits

A chemist measures the mass of chloride in three water samples, as shown in the table. Together, the three samples have a volume of 2.31 liters. What is the average mass of chloride per liter of water? Answer to significant digits.

Sample	Mass of Chloride
A	15.21 mg
B	9.33 mg
C	11.329 mg

total mass chloride:

15.21 mg
9.33 mg
11.329 mg

35.869 mg

= 35.87 mg

4 sig. digits

$$\frac{\text{total mass}}{\text{volume}}$$

Use unrounded mass

$$= \frac{35.869 \text{ mg}}{2.31 \text{ L}}$$

4 sig. digits

3 sig. digits

$= 15.5\cancel{2770563}$

= 15.5 mg/L

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A. 4.28 kcal/mL

B. 4.29 kcal/mL (correct answer)

C. 0.428 kcal/mL

D. 0.429 kcal/mL

E. 4.286 kcal/mL

Unit conversions

Problems in chemistry

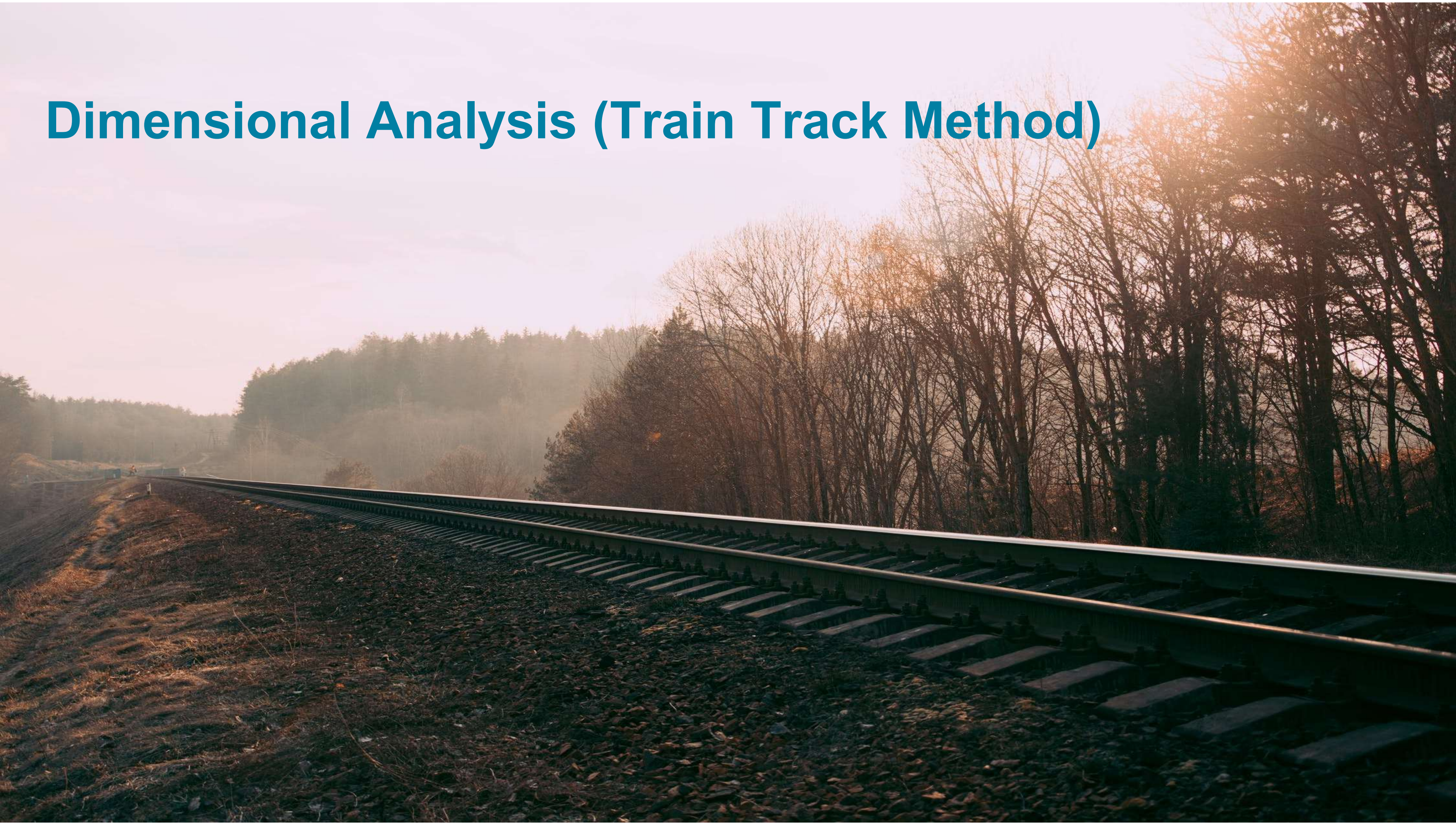
1. **Unit conversion type**:

Many of the problems can be thought of as *unit conversion problems,* in which you are given one or more quantities and asked to convert them into different units.

2. **Specific equation type**:

Other problems require the use of *specific equations* to get to the information you are trying to find.

Dimensional Analysis (Train Track Method)

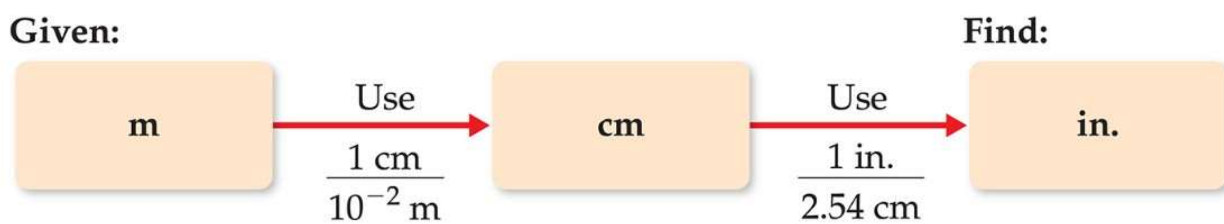


1.7 Dimensional Analysis

- **Dimensional analysis** is used to change units.
- We apply **conversion factors** (e.g., 1 in. = 2.54 cm.), which are equalities.
- We can set up a ratio of comparison for the equality:

  1 in. / 2.54 cm **or** 2.54cm / 1 in.

- We use the ratio which allows us to change units (puts the units we have in the denominator to cancel).
- We can use multiple conversions, as long as each one is an equality.

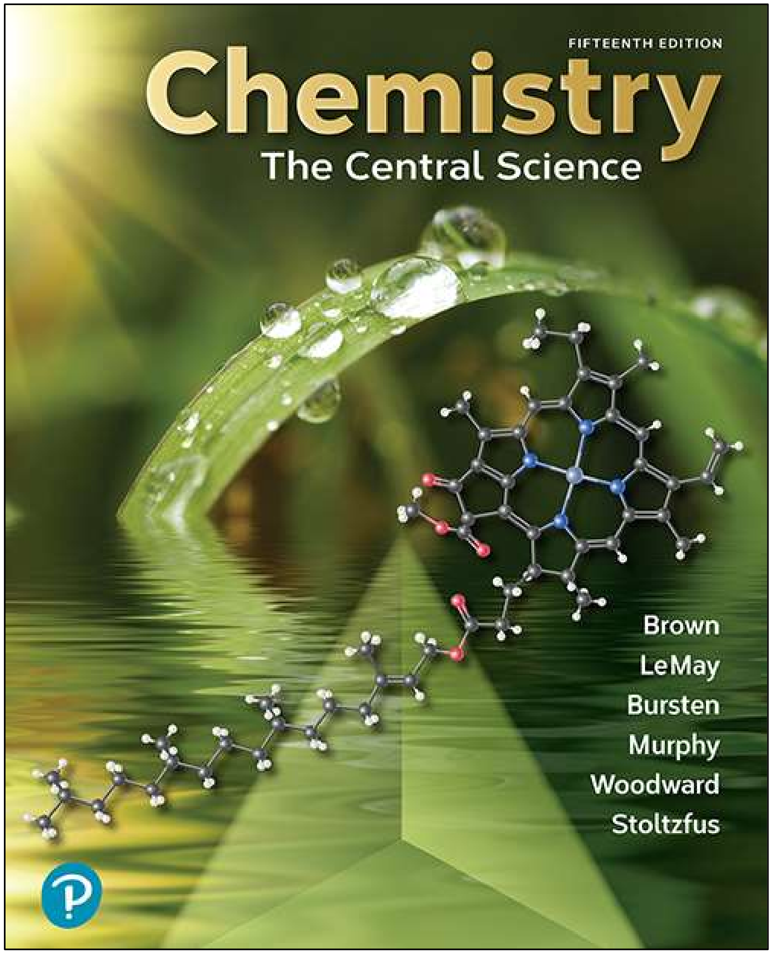


Dimensional Analysis Template

1. State the given and needed units.
2. Write a unit plan to convert the given unit to the final unit.
3. State the equalities and conversion factors to cancel units.
4. Set up problem to cancel units, calculate, and report the answer with the proper number of significant figures.


Questions?

Chemistry: The Central Science



Chapter 2

Atoms, Molecules, and Ions

Atoms: Chemistry Building Blocks

Atoms and molecules are <u>tiny particles that compose all common matter</u>.

The atoms are bound together to form several different types of molecules.

Chemical bonds are the attachments or “glue” that hold atoms together.

2.1 Atomic Theory of Matter

- Experiments in the eighteenth and nineteenth centuries led to an organized atomic theory by John Dalton in the early 1800s:
  - The law of constant composition
  - The law of conservation of mass
  - The law of multiple proportions

Law of Constant Composition

- We introduced this in Chapter 1.
- Compounds have a definite composition. That means that the relative number of atoms of each element in the compound is the same in any sample.
  - H_2O, CO, CO_2
- This law was discovered by Joseph Proust.
- This law was one of the laws on which Dalton’s atomic theory (Postulate 4) was based.

Law of Conservation of Mass

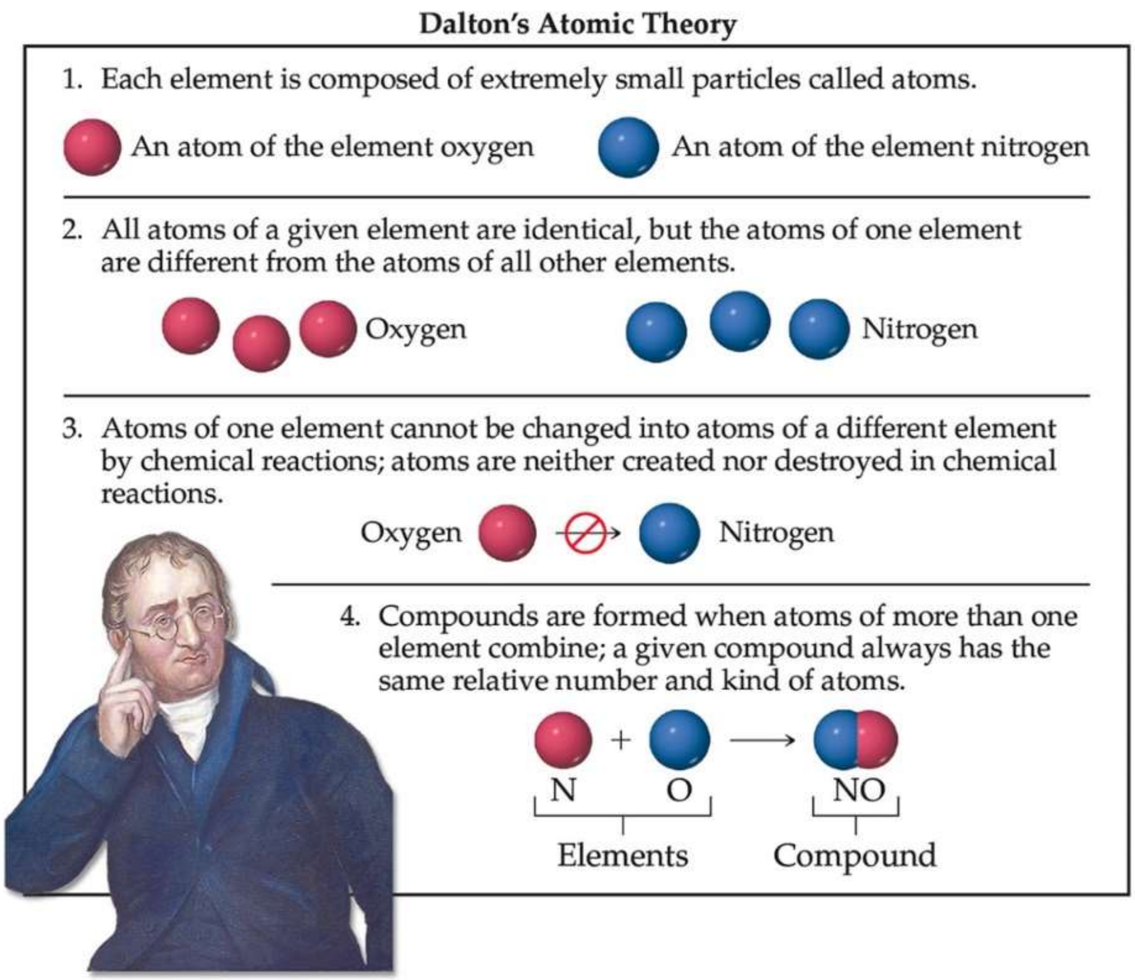
- The **total mass** of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.
- This law is further explained in Chapter 3.
- This law was discovered by Antoine Lavoisier.
- This law was one of the laws on which Dalton's atomic theory (Postulate 3) was based.

Law of Multiple Proportions

- If two elements, A and B, form more than one compound, the masses of B that combine with a given mass of A are in the ratio of small whole numbers.
- When two or more compounds exist from the same elements, they cannot have the same relative number of atoms, i.e. carbon monoxide CO (poisonous gas) versus

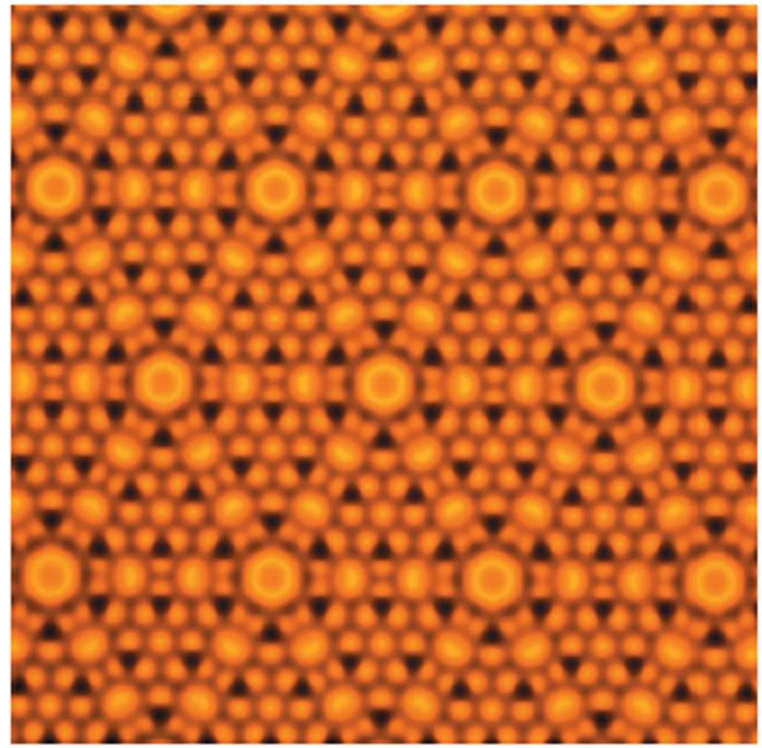
  carbon dioxide CO_2 (what we exhale).
- John Dalton discovered this law while developing his atomic theory.

Postulates of Dalton’s Atomic Theory



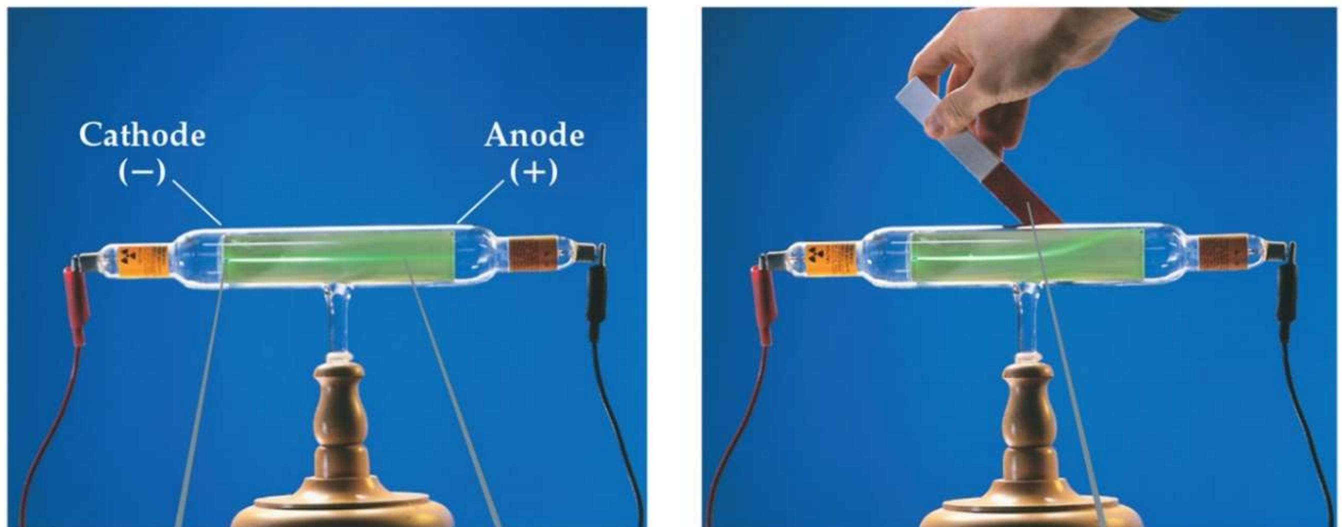
2.2 Discovery of Subatomic Particles

- In Dalton's view, the atom was the smallest particle possible. Many discoveries led to the fact that the atom itself was made up of smaller particles.
  - Electrons and cathode rays
  - Radioactivity
  - Nucleus, protons, and neutrons
  - Today, we can measure the properties of individual atoms and even obtain images of them, that is, silicon



Cathode Rays

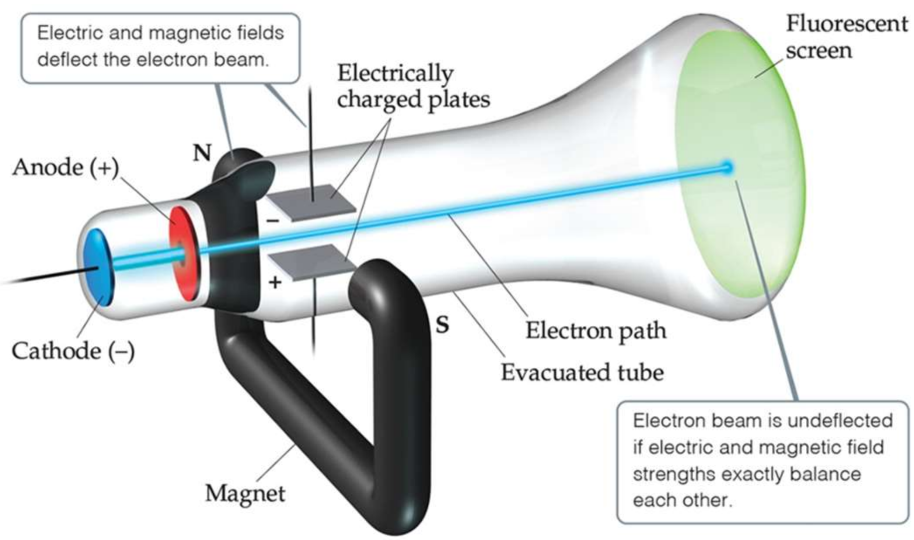
- Streams of negatively charged particles were found to emanate from cathode tubes, causing fluorescence.
- J. J. Thomson is credited with his discovery (1897).



Cathode rays (electrons) move from the negative cathode to the positive anode.

A fluorescent screen is placed in the tube to show the path of the cathode rays. The screen gives off light when a cathode ray strikes it.

The rays are deflected by a magnet.



An English physicist named J. J. Thomson (1856–1940) discovered a smaller and more fundamental particle called the **electron.**

Thomson discovered the following:

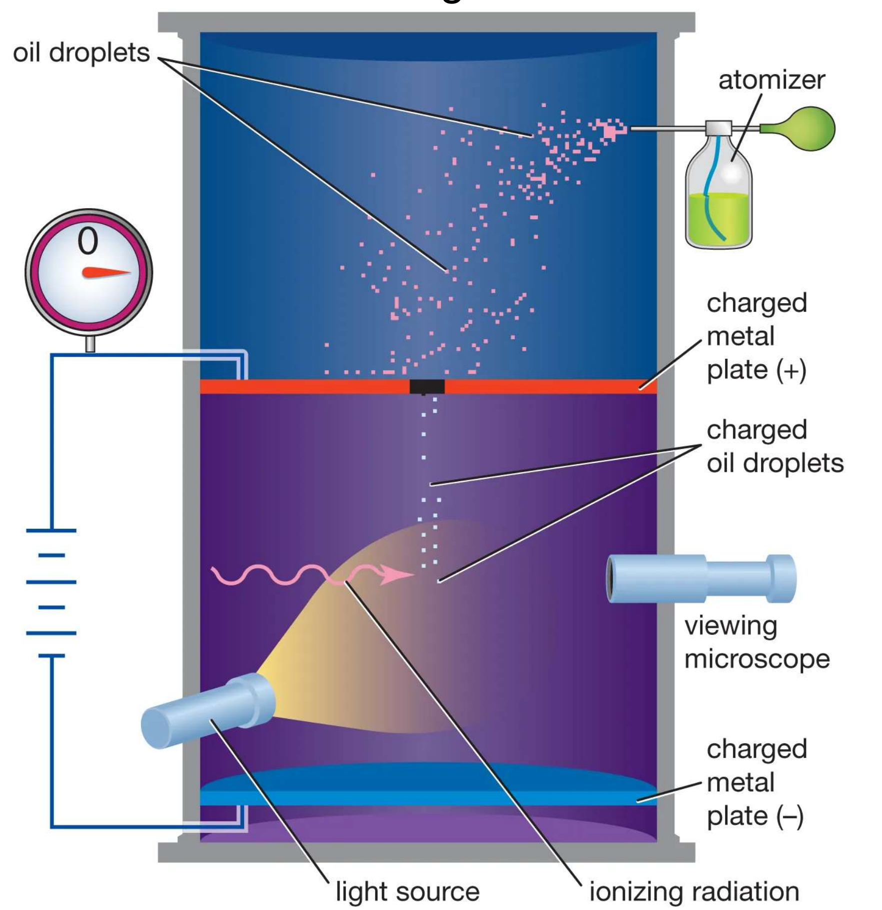
1. Electrons are negatively charged.
2. Electrons are much smaller and lighter than atoms.
3. Electrons are uniformly present in many different kinds of substances.
4. He proposed that atoms must contain positive charge that balances the negative charge of electrons.

Millikan Oil-Drop Experiment

- Once the charge/mass ratio of the electron was known, determination of either the charge or the mass of an electron would yield the other.
- Robert Millikan determined the charge on the electron in 1909 was 1.602×10^{-19} C.

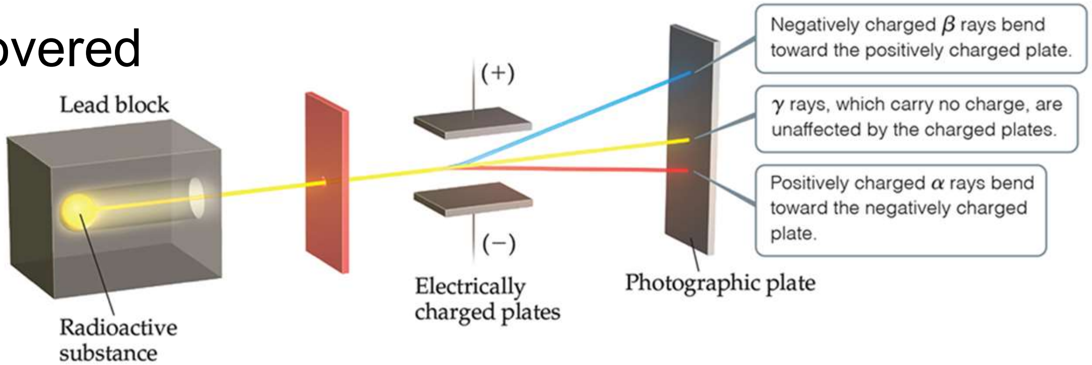
  Electron mass now known.

Ionizing Radiation = electrons

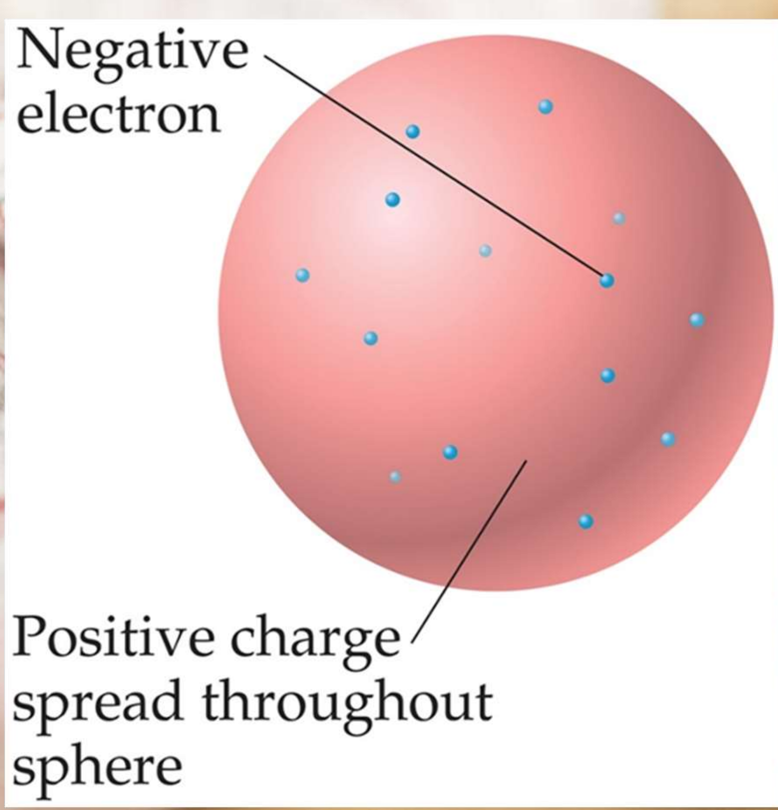


Radioactivity

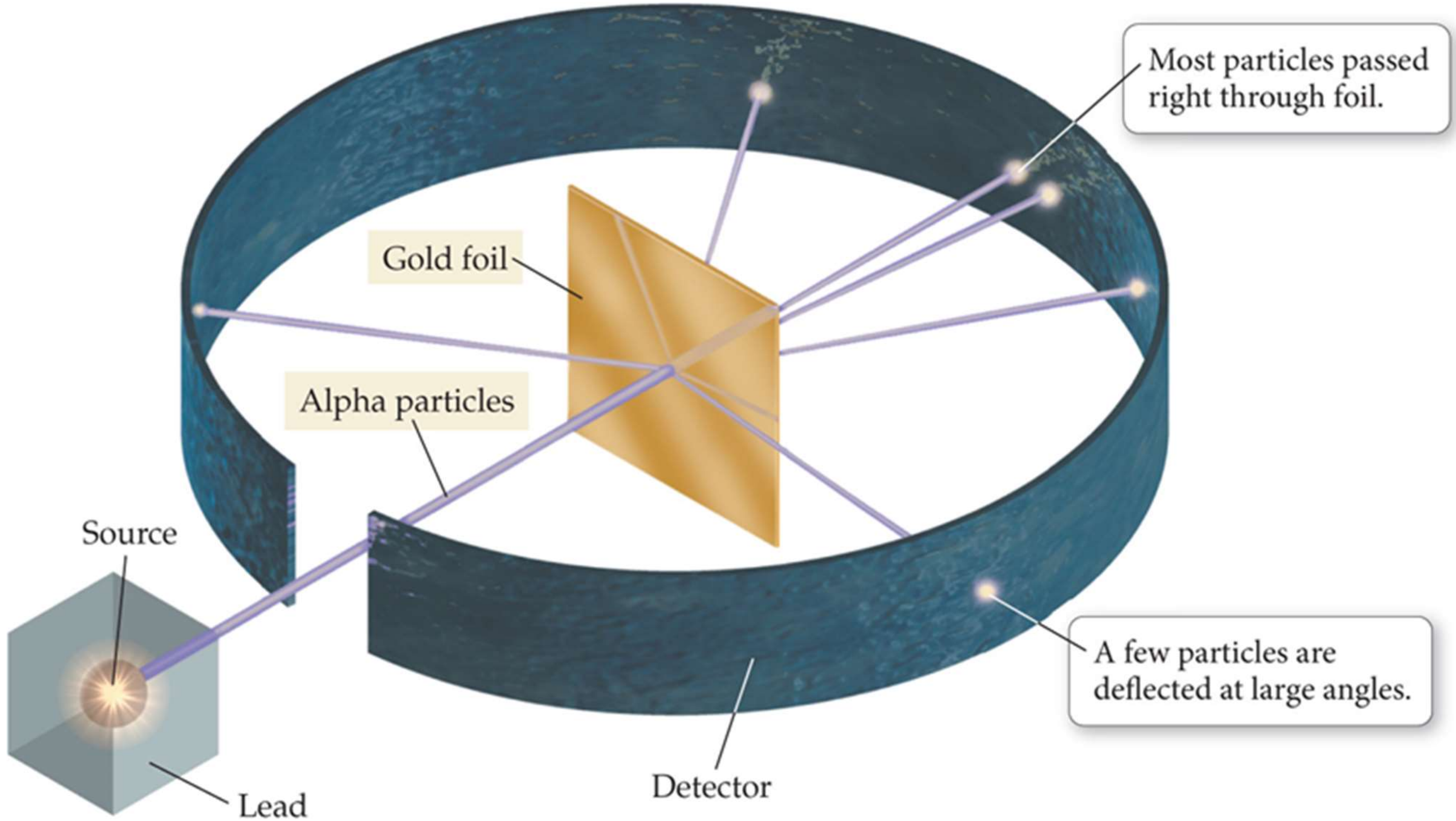
- **Radioactivity** is the spontaneous emission of high-energy radiation by an atom.
- It was first observed by Henri Becquerel.
- Marie and Pierre Curie also studied it.
- Its discovery showed that the atom had more subatomic particles and energy associated with it.
- Three types of radiation were discovered by Ernest Rutherford:
  - α particles (positively charged)
  - β particles (negatively charged, like electrons)
  - γ rays (uncharged)

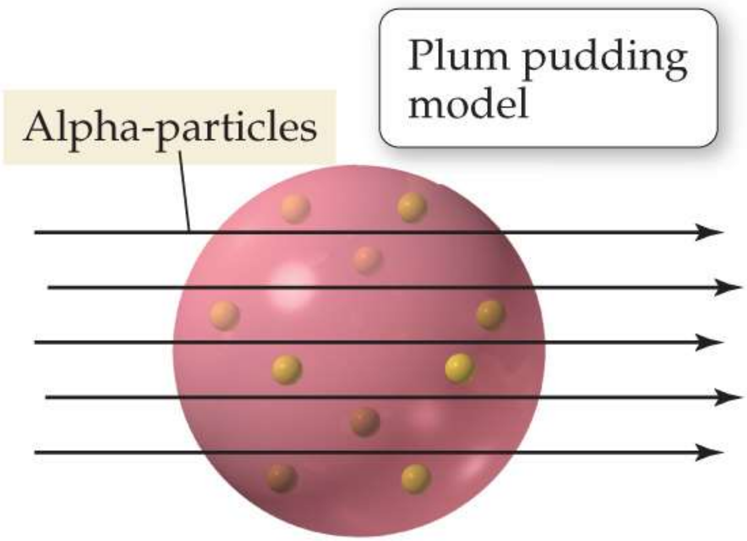


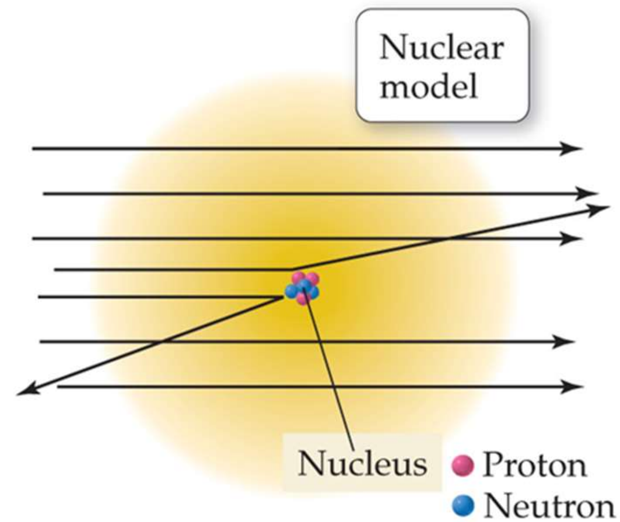
Old theory of atoms



The prevailing theory was that of the “plum pudding” model, put forward by J. J. Thomson.

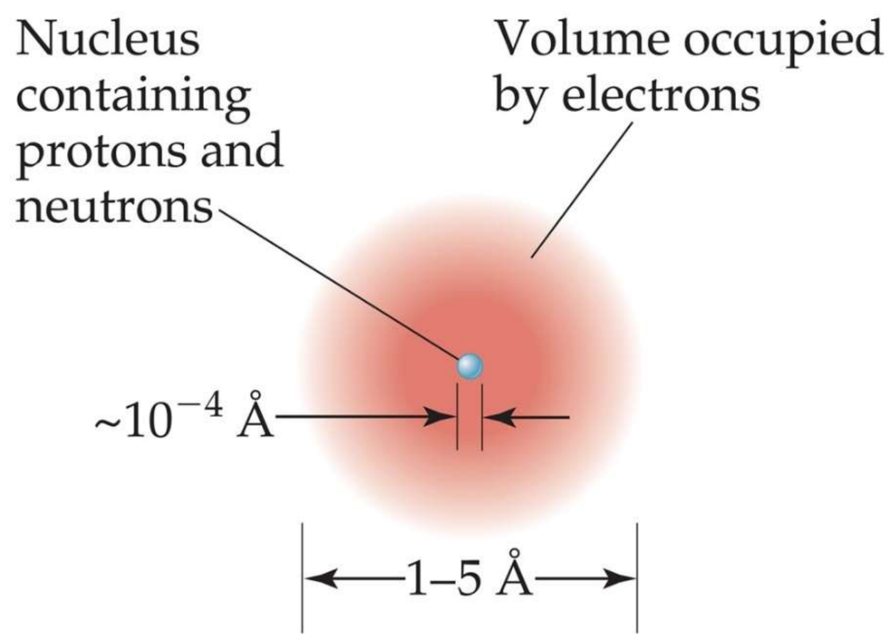

Most particles passed right through foil.
Gold foil
Alpha particles
Source
Lead
Detector
A few particles are deflected at large angles.


Plum pudding model
Alpha-particles


Nuclear model
Nucleus
Proton
Neutron

2.3 Modern View of Atomic Structure

- Rutherford postulated a very small, dense positive center with the electrons around the outside.
- We now know that most of the atom is empty space.
  - Atoms are very small; 1–5 Å or 100–500 pm.
  - Other subatomic particles (protons and neutrons in the nucleus) discovered.



Subatomic Particles

- Protons (+1) and electrons (−1) have a charge; neutrons are neutral.
- Protons and neutrons have essentially the same mass (relative mass 1). The mass of an electron is so small we ignore it (relative mass 0).
- Protons and neutrons are found in the nucleus; electrons travel around the nucleus.

Particle	Symbol	Charge	Mass (amu)	Location in Atom
Proton	p or p^+	1+	1.007	Nucleus
Neutron	n or n^0	0	1.008	Nucleus
Electron	e^-	1−	0.000 55	Outside nucleus

Atoms Are Neutral

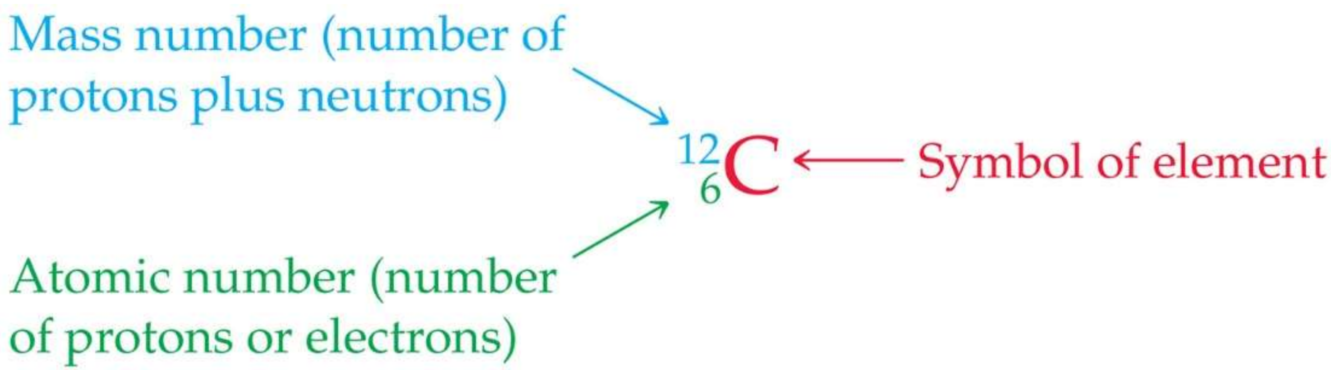
For neutral atoms, the net charge is zero.

number of protons = number of electrons

Aluminum has 13 protons and 13 electrons.
The net (overall)charge is zero.

13 protons (13+)   + 13 electrons (13–) = 0

Atoms of an Element



- Elements are represented by a one or two letter **symbol**, for which the first letter is always capitalized. **C** is the **symbol** for carbon.
- All atoms of the same element have the same number of protons, which is called the **atomic number**. It is written as a **subscript before** the symbol. **6** is the atomic number for carbon.
- The **mass number** is the total number of protons and neutrons in the nucleus of an atom. It is written as a **superscript before** the symbol.

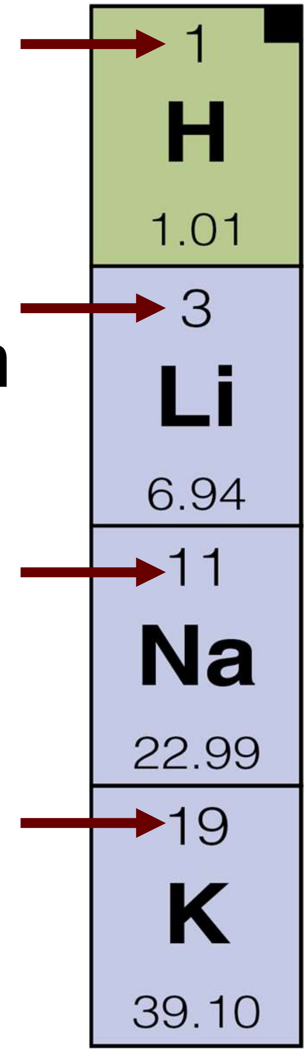
Atomic Number and Mass Number

Atomic number

The number of protons in an atom
Also the number of electrons in a neutral atom

Mass number

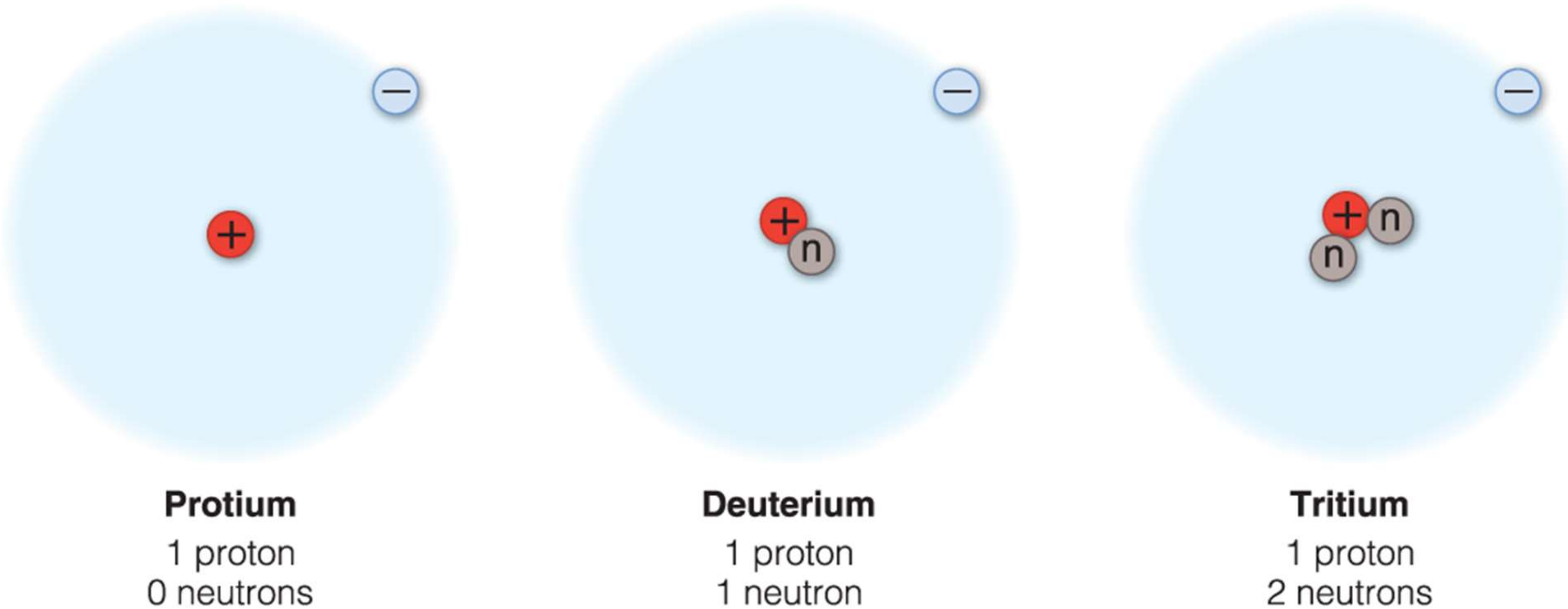
The number of protons + neutrons
Typically, the periodic table does not show the mass number.



Isotopes

Have the same atomic number, but different mass number

Three isotopes of hydrogen:



2.4 Atomic Mass Unit (amu)

- Atoms have extremely small masses.
- In 100 g water, there are 1) 1.1 g of H and 88.9 g of O. and 2) Two H for each O. H was arbitrarily assigned a mass of 1. Masses of all other atoms were assigned relative to H, that is, O = 16.
- Today we can determine the mass to high degree of accuracy and precision.
- A mass scale on the atomic level is used, where an atomic mass unit (amu) is the base unit.

$$1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g}$$

Atomic Weight

- Because in the real world we use large amounts of atoms and molecules, we use average masses in calculations.
- An average mass is found using all isotopes of an element weighted by their relative abundances. This is the element's **atomic weight**.

$$\text{Atomic Weight} = \sum[(\text{isotope mass}) \times (\text{fractional natural abundance})] \text{ for ALL isotopes.}$$

- The masses of any atom is compared to C-12 (6 protons and 6 neutrons) being exactly 12.

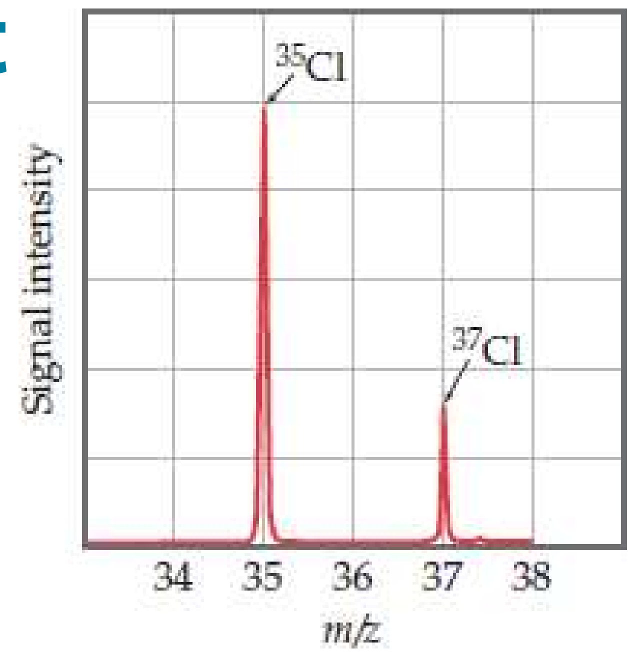
Example Calculation

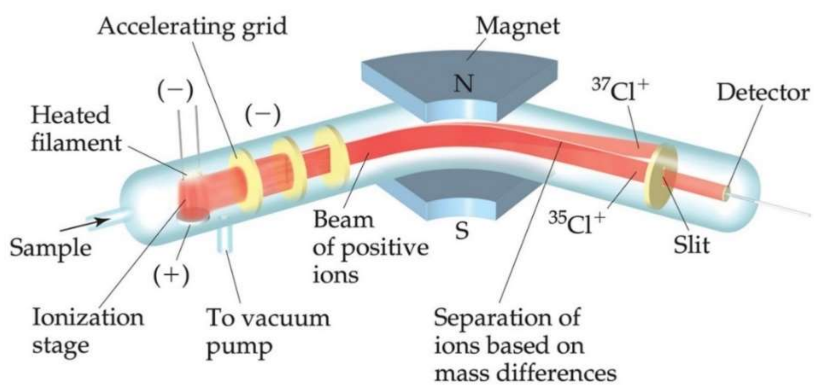
Three isotopes of silicon occur in nature: ^{28}Si (92.23%), atomic mass 27.97693 amu; ^{29}Si (4.68%), atomic mass 28.97649 amu; and ^{30}Si (3.09%), atomic mass 29.97377 amu. Calculate the atomic weight of silicon.

$$\text{Atomic Weight} = \sum[(\text{isotope mass}) \times (\text{fractional natural abundance})] \text{ for ALL isotopes.}$$

Atomic Weight Measurement

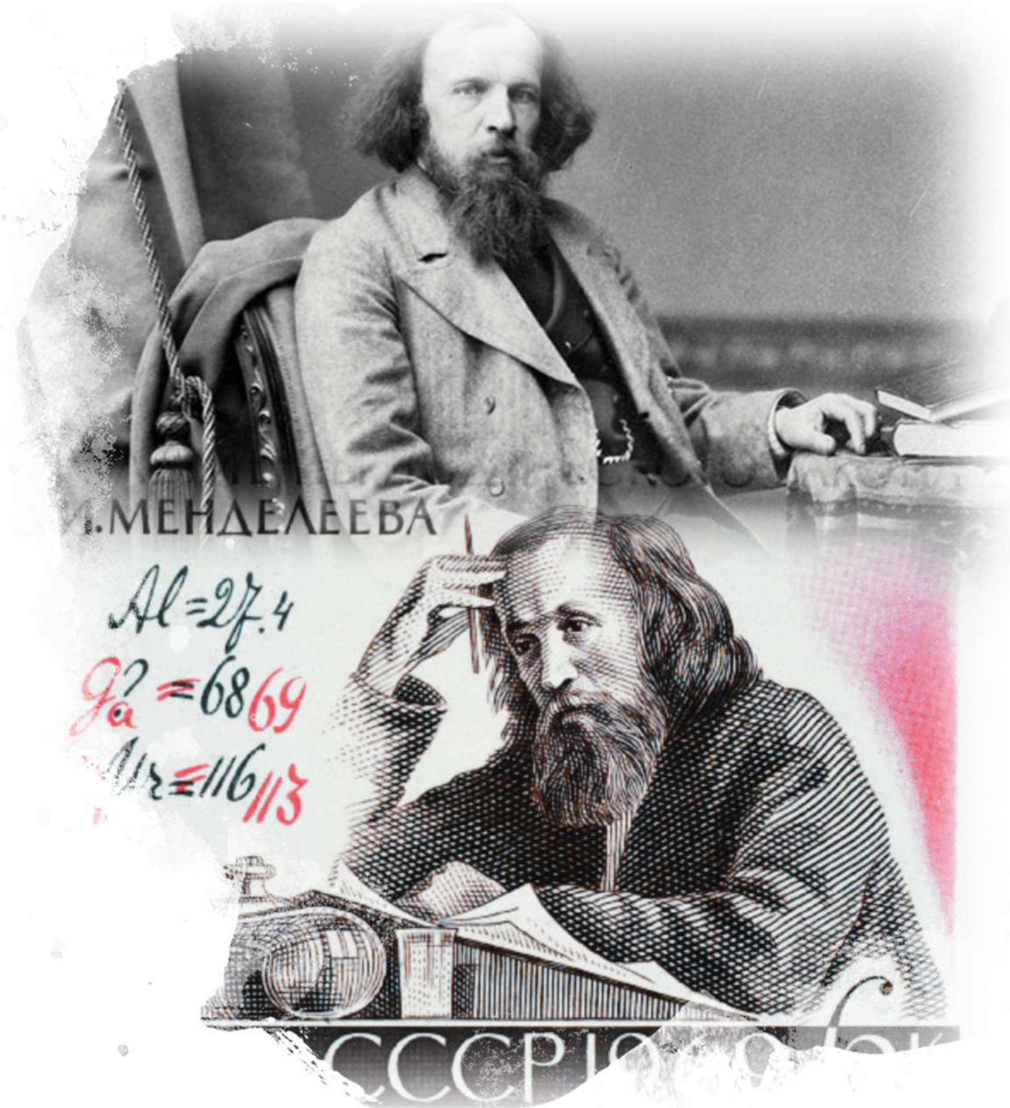
- Atomic and molecular weight can be measured using a mass spectrometer (below).
- The spectrum of chlorine showing two isotopes is seen on the right. Isotope abundance can also be determined this way.





Looking for Patterns: Dmitri Mendeleev

Dmitri Mendeleev, a Russian chemistry professor, proposed from observation that when the elements are arranged in order of increasing relative mass, certain sets of properties recur periodically.



Looking for Patterns: Recurring Properties

1 H	2 He	3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne	11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar	19 K	20 Ca

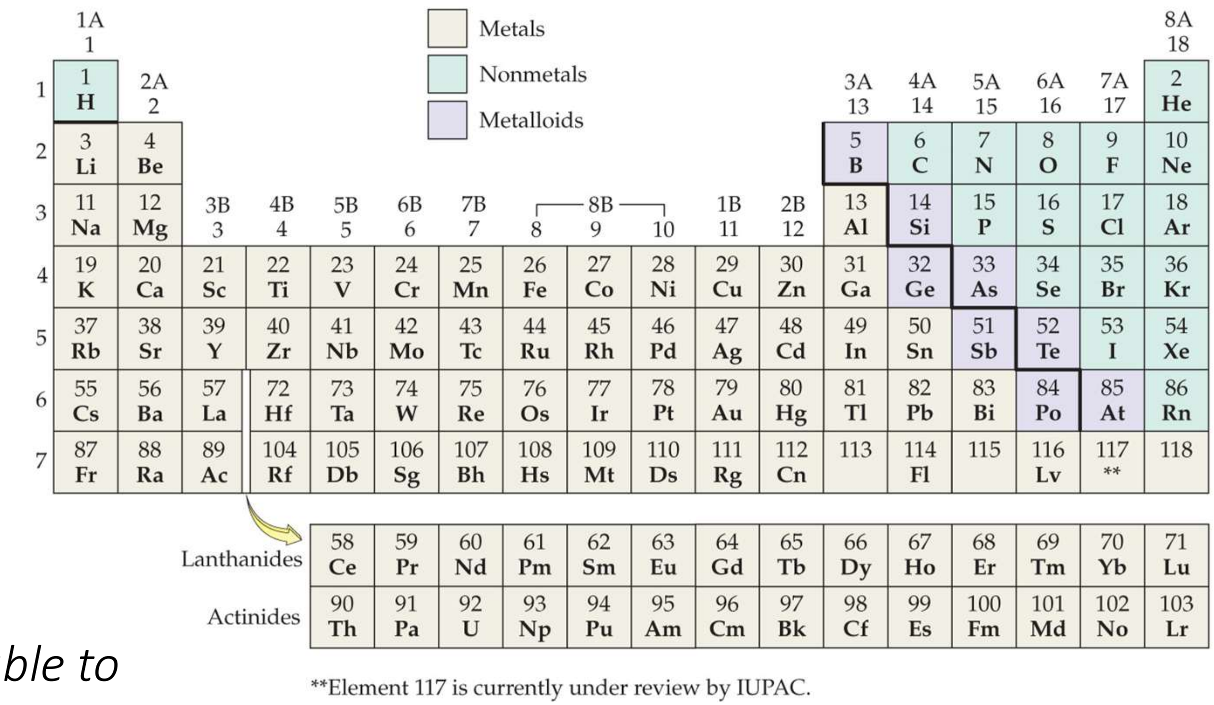
The color of each element represents its properties.

We arrange them in rows so that similar properties align in the same vertical columns. This figure is similar to Mendeleev's first periodic table.

1 H							2 He
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca						

The Periodic Table

*The **periodic table** organizes 118 elements into groups with similar properties and places them in order of increasing atomic mass.*



Learning Goal *Use the periodic table to identify the group and the period of an element; identify the element as a metal, a nonmetal, or a metalloid.*

Metals

Nonmetals

Metalloids

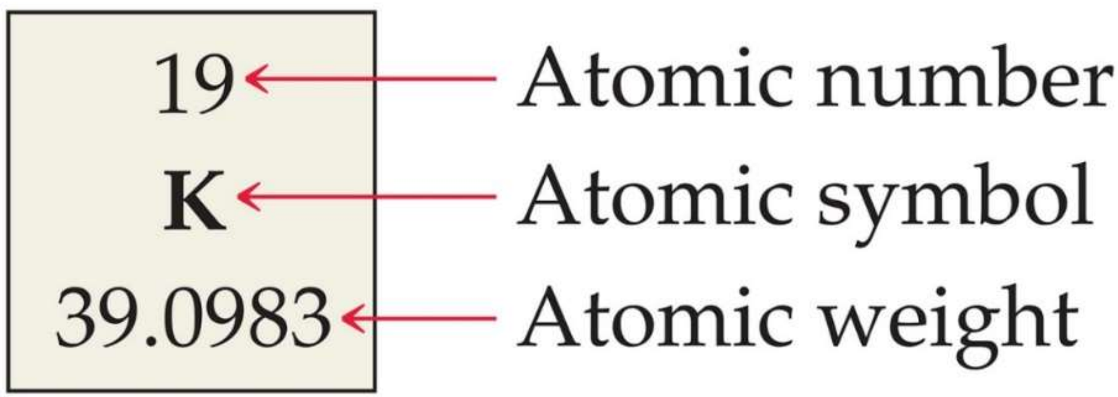
	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H**																	2 **He**
2	3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
3	11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**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

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

**Element 117 is currently under review by IUPAC.

Reading the Periodic Table

- Boxes on the periodic table list the atomic number **Above** the symbol.
- The atomic weight of an element is listed below the symbol on the periodic table.



Groups and Periods

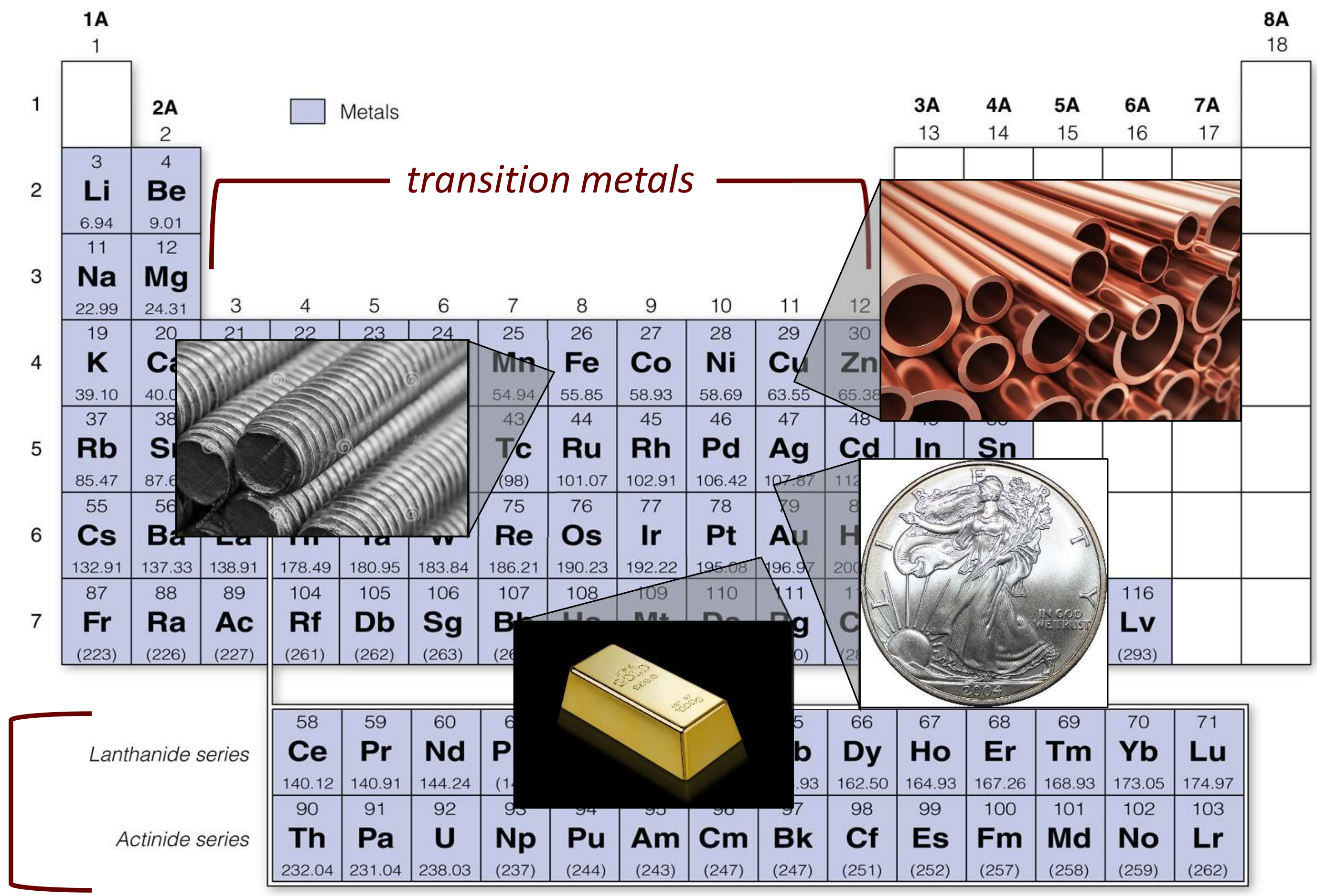
Main-group elements | Transition elements | Main-group elements

Group number

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

**Element 117 is currently under review by IUPAC.

Metals



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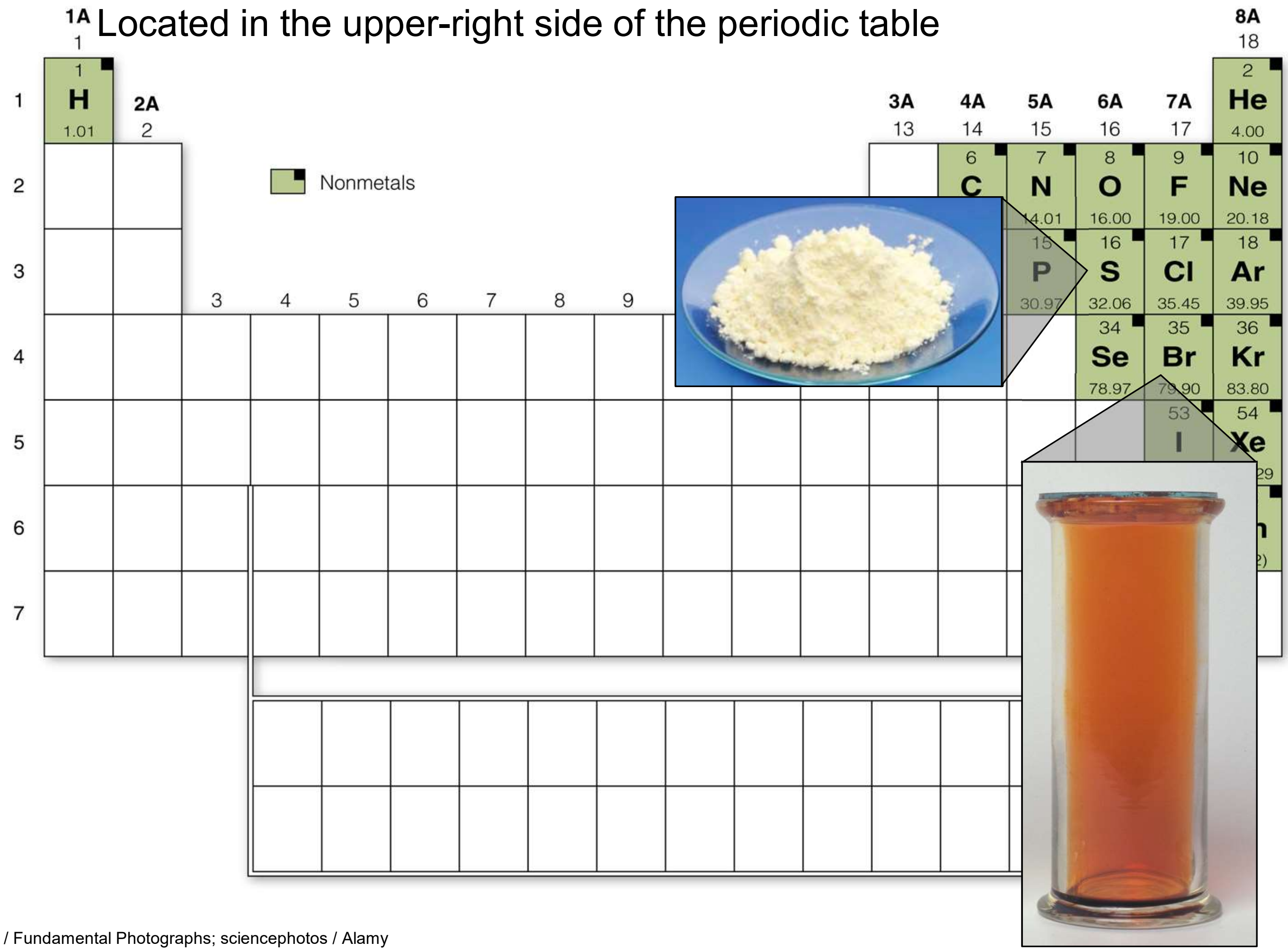
Nonmetals

Located in the upper-right side of the periodic table



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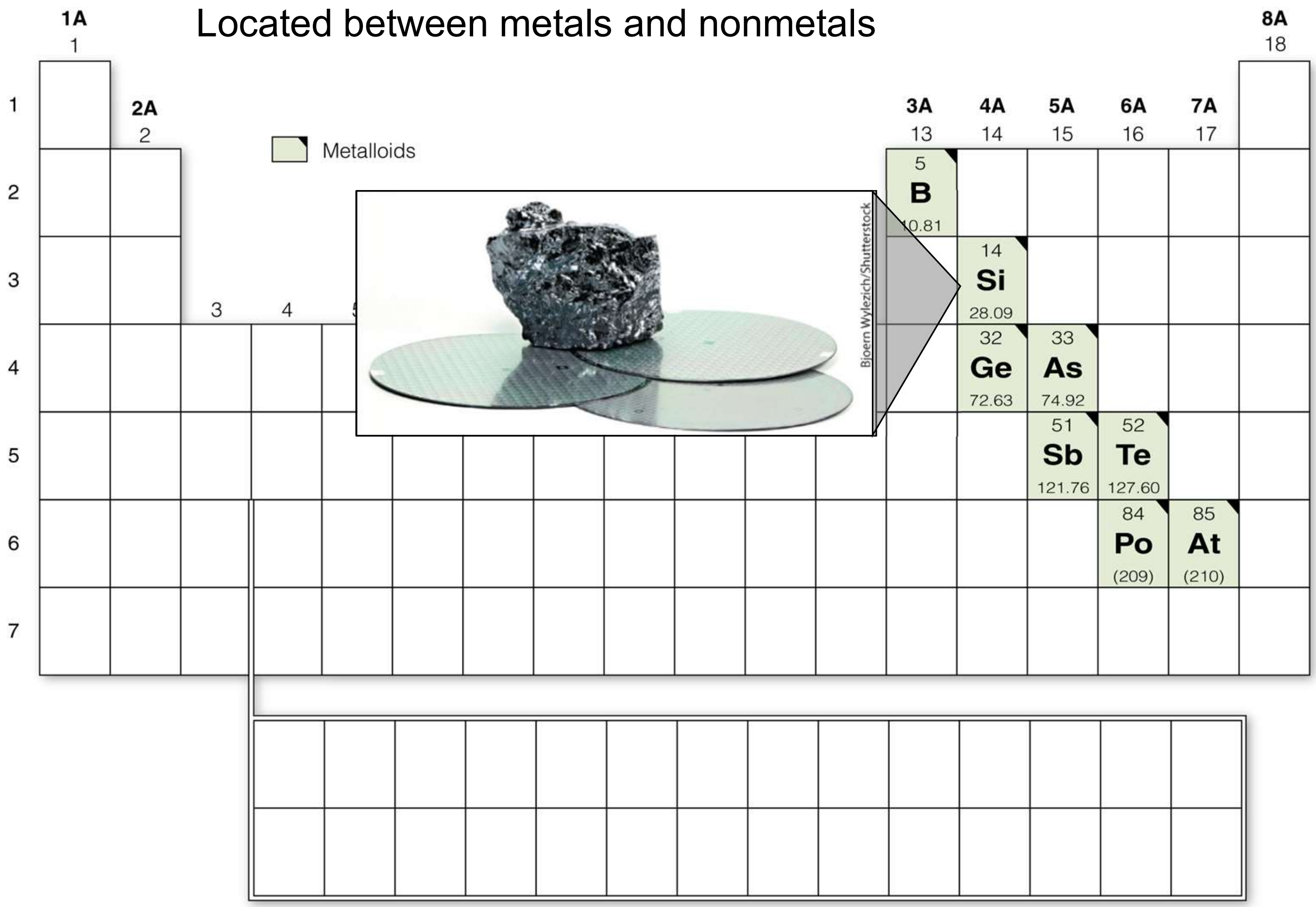
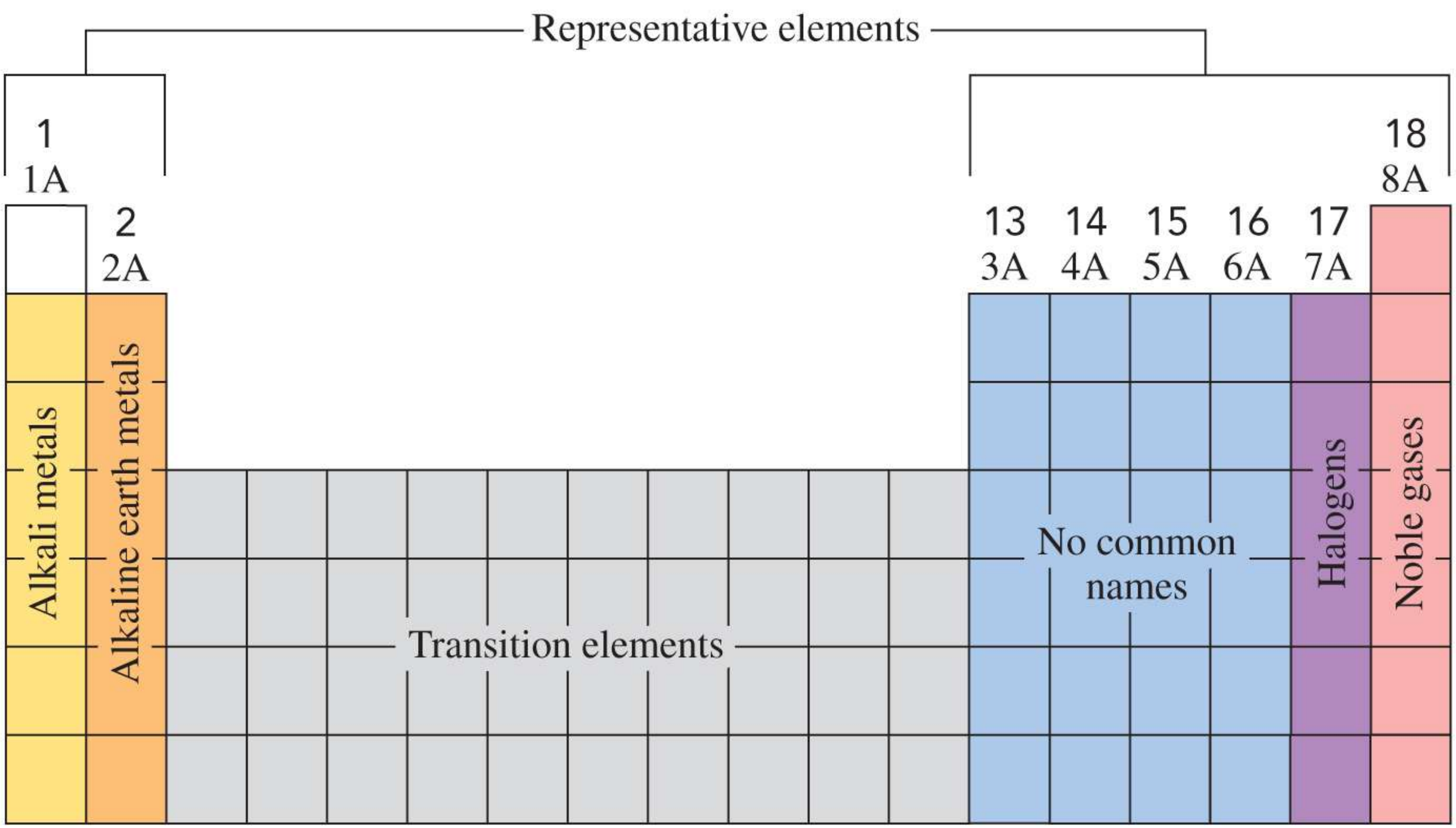
Metalloids

Located between metals and nonmetals

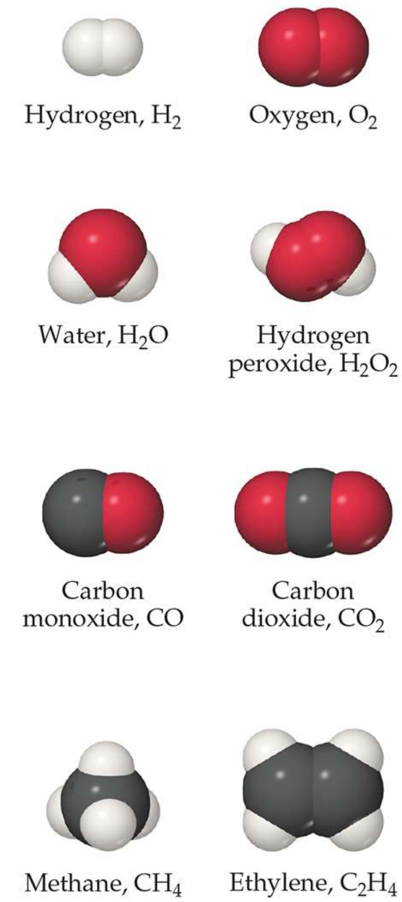


Names of Groups



2.6 Molecules and Molecular Compounds

- **Chemical Formula**: The subscript to the right of the symbol of an element tells the number of atoms of that element in one molecule of the compound.
- **Molecular compounds:** They are composed of molecules and almost always contain only nonmetals.



Diatomic Molecules

- These seven elements occur naturally as molecules containing two atoms:
  - Hydrogen (H_2)
  - Nitrogen (N_2)
  - Fluorine (F_2)
  - Oxygen (O_2)
  - Iodine (I_2)
  - Chlorine (Cl_2)
  - Bromine (Br_2)

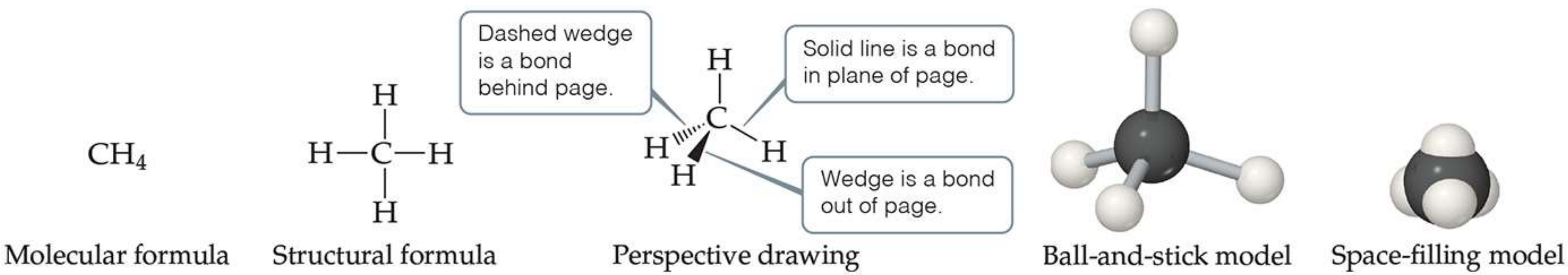
Have
No
Fear
Of
Ice
Cold
Beer

Note their location on the periodic table

Types of Formulas

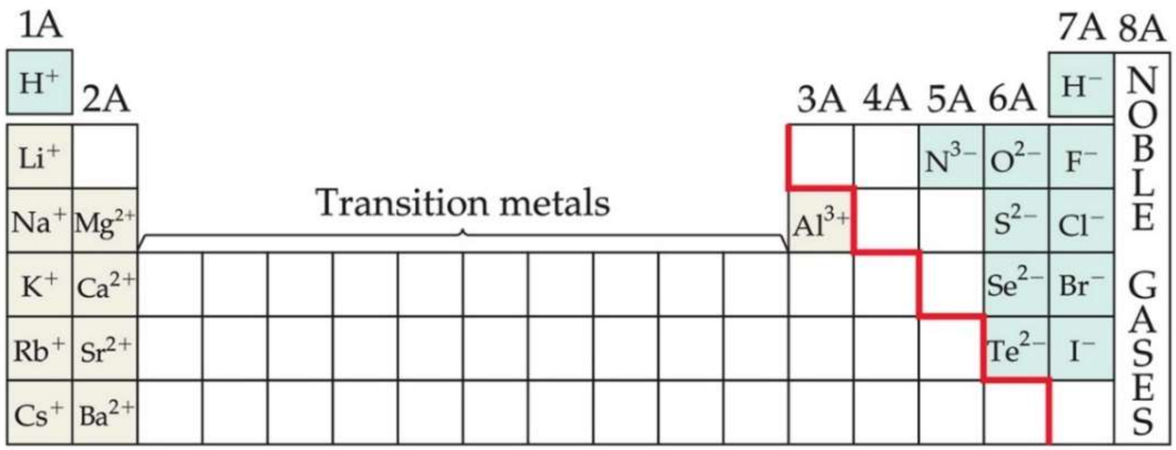
- **Empirical formulas** give the **lowest** whole-number ratio of atoms of each element in a compound.
- **Molecular formulas** give the **exact** number of atoms of each element in a compound.
- If we know the molecular formula of a compound, we can determine its empirical formula. The converse is not true without more information.

Picturing Molecules



- **Structural formulas** (2D) show the order in which atoms are attached. They do **NOT** depict the three-dimensional (3D) shape of molecules.
- **Perspective drawings, ball-and-stick models, and space-filling models** show the three-dimensional (3D) order of the atoms in a compound.

2.7 Ions and Ionic Compounds

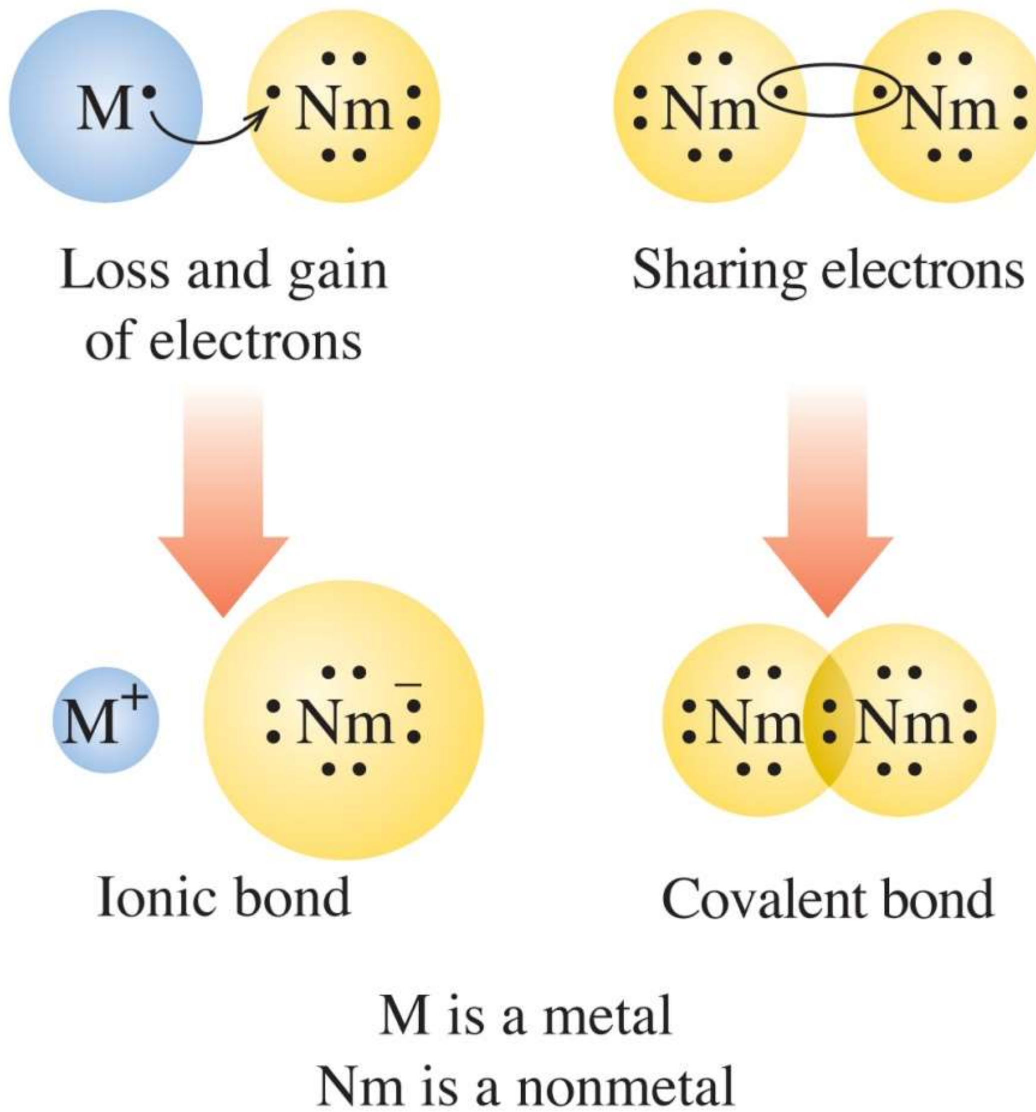


- When an atom of a group of atoms loses or gains electrons, it becomes an **ion**.
- **Cations** are formed when at least one electron is lost. Monatomic cations are formed by metals.
- **Anions** are formed when at least one electron is gained. Monatomic anions are formed by nonmetals, except the noble gases.

Ionic and Covalent Bonds

Atoms that are not noble gases form octets to become more stable.

They do this by losing, gaining, or sharing valence electrons forming ionic bonds or covalent bonds.

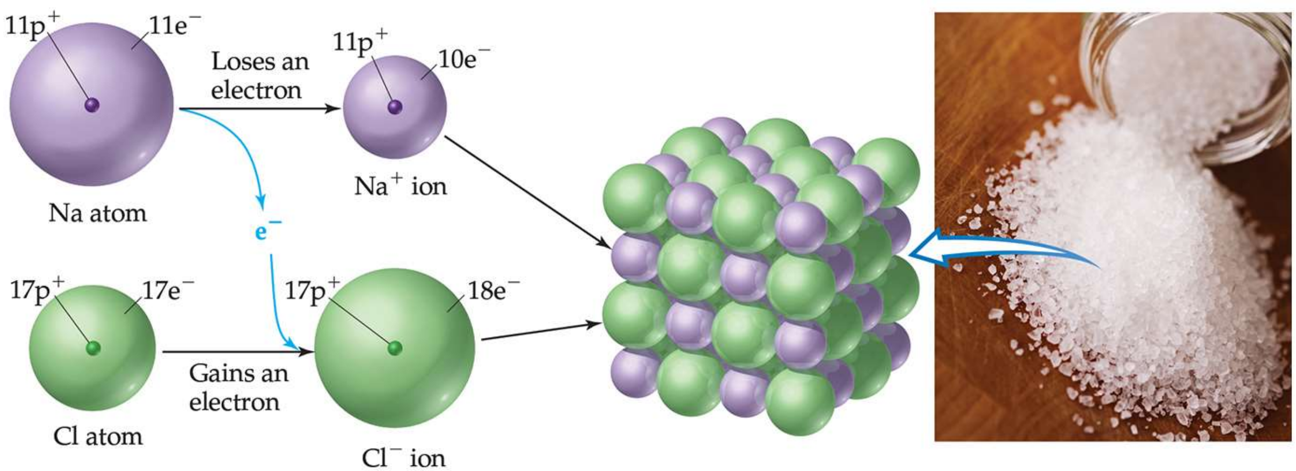


Forming ions (Dr. Drake's way)

		Metals Lose Valence Electrons			Nonmetals Gain Valence Electrons				
Noble Gases	Electron Arrangement	1A (1)	2A (2)	3A (13)	5A (15)	6A (16)	7A (17)	Electron Arrangement	Noble Gases
He	←	Li^{+}							
Ne	←	Na^{+}	Mg^{2+}	Al^{3+}	N^{3-}	O^{2-}	F^{-}	→	**Ne**
Ar	←	K^{+}	Ca^{2+}		P^{3-}	S^{2-}	Cl^{-}	→	**Ar**
Kr	←	Rb^{+}	Sr^{2+}				Br^{-}	→	**Kr**
Xe	←	Cs^{+}	Ba^{2+}				I^{-}	→	**Xe**

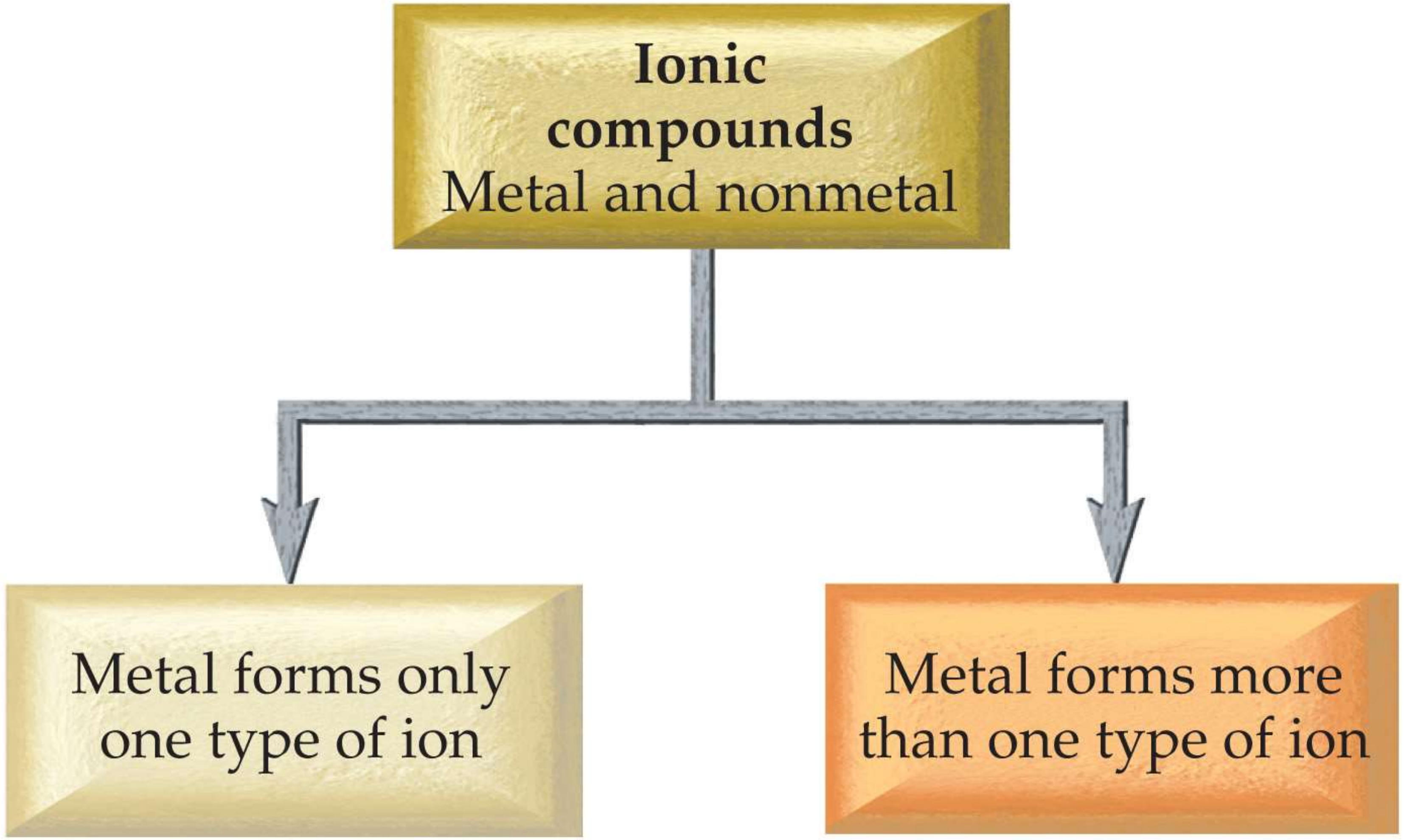
Ionic Compounds

- **Ionic compounds** (such as NaCl) are generally formed between metals and nonmetals.
- Electrons are transferred from the metal to the nonmetal. The oppositely charged ions attract each other. Only empirical formulas are written.



Is everyone comfortable with forming ions?

Categorizing Types of Ionic Compounds



Metals Whose Charge Is Invariant

TABLE 5.4 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
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	*
Ag	Ag^{+}	silver	*

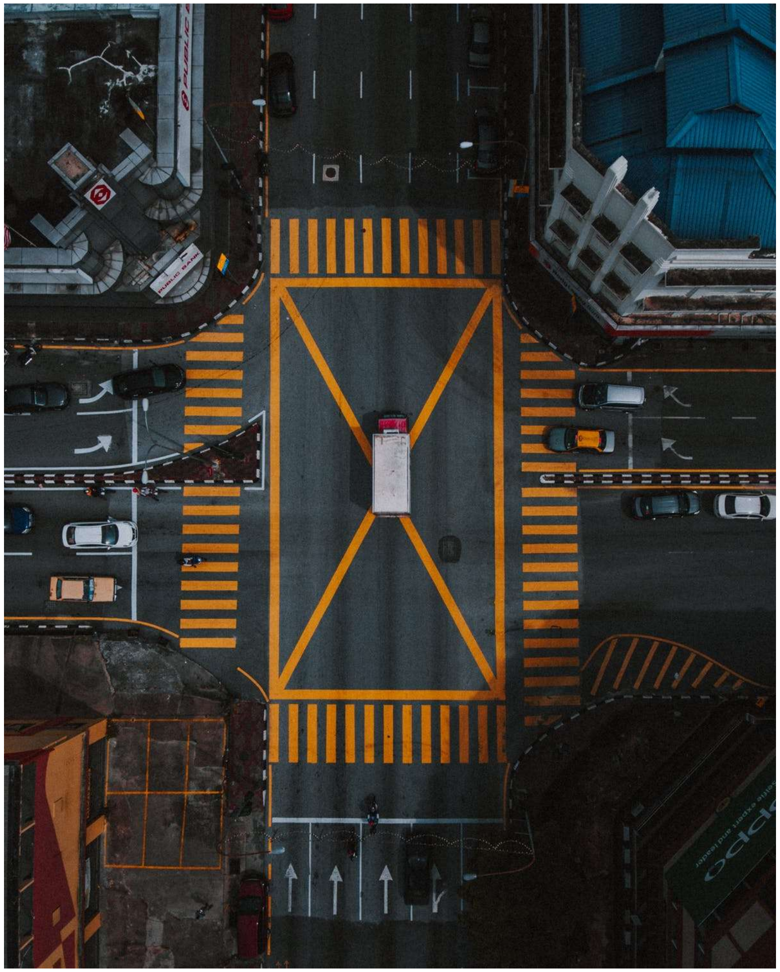
*The charge of these metals cannot be inferred from their group number.

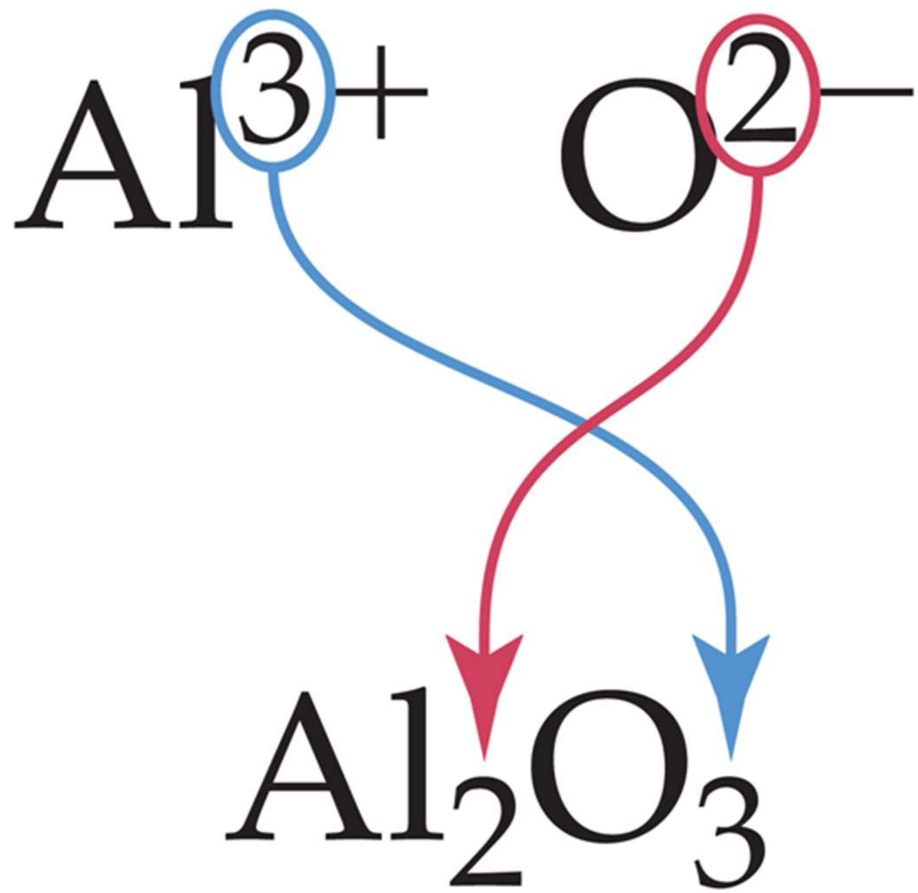
Binary Ionic Compounds with Metal of Invariant Charge

name of cation
(metal)

base name of anion
(nonmetal) + *-ide*

Since the charge of the metal is always the same for these types of compounds, it need not be specified in the compound's name.




Al^{3+} O^{2-}
Al_2O_3

The base names for various nonmetals and their most common charges in ionic compounds

TABLE 5.6 Some Common 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

Binary Ionic Compounds with Metal Whose Charge May Vary (polyvalent)

The full names for these types of compound have the following form:

name of cation (metal) (charge of cation (metal) in roman numerals in parentheses) base name of anion (nonmetal) + *-ide*

Some Metals that Form More than One Type of Ion

TABLE 5.5 Some Metals That Form More Than One Type of Ion and Their Common Charges

Metal	Symbol 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. We do *not* use this older system in this text.

Writing Formulas for Compounds with Polyatomic Ions

Recognize polyatomic ions in a chemical formula by becoming familiar with these common polyatomic ions.

TABLE 5.3 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^-
nitrate	NO_3^-	permanganate	MnO_4^-
nitrite	NO_2^-	sulfate	SO_4^{2-}
chromate	CrO_4^{2-}	sulfite	SO_3^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfite (or bisulfite)	HSO_3^-
phosphate	PO_4^{3-}	hydrogen sulfate (or bisulfate)	HSO_4^-
hydrogen phosphate	HPO_4^{2-}	peroxide	O_2^{2-}
ammonium	NH_4^+	cyanide	CN^-

Elements May Be Atomic or Molecular

Atomic elements are those that exist in nature with single atoms as their basic units. Most elements fall into this category.

Molecular elements do not normally exist in nature with single atoms as their basic units. Instead, these elements exist as diatomic molecules—two atoms of that element bonded together—as their basic units.

Elements that Exist as Diatomic Molecules

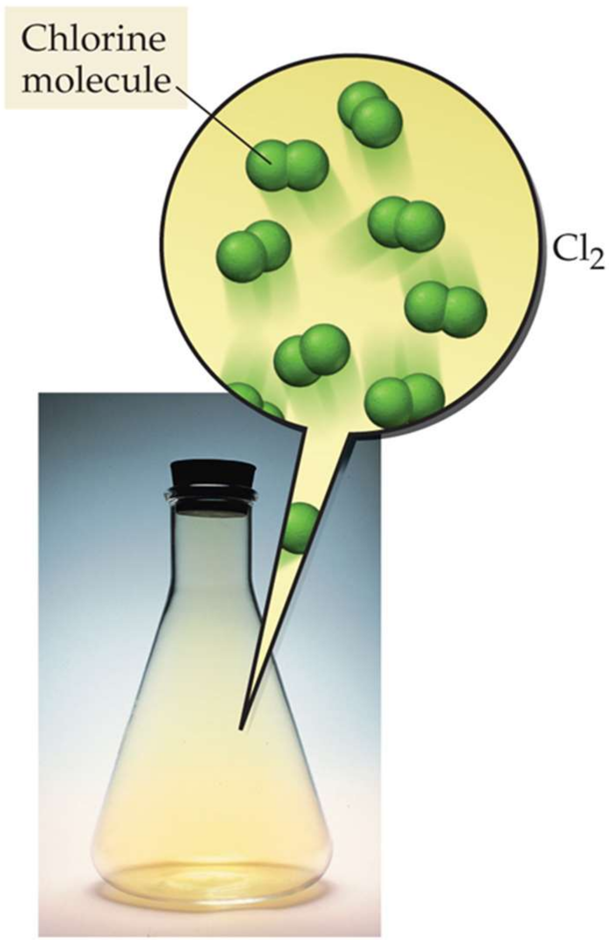
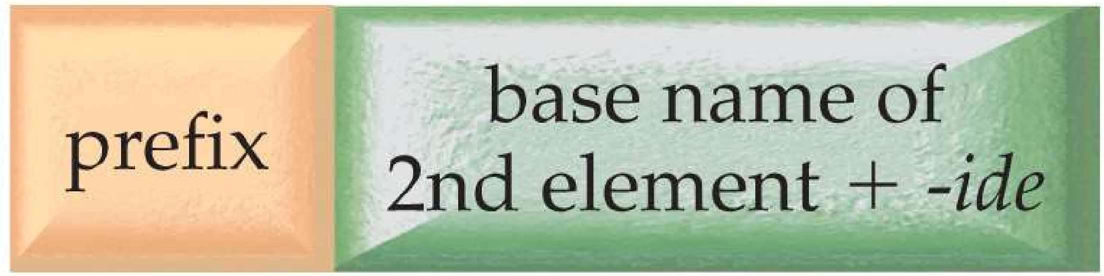


TABLE 5.2 Elements That Occur as Diatomic Molecules

Name of Element	Formula of Basic Unit
hydrogen	H_2
nitrogen	N_2
oxygen	O_2
fluorine	F_2
chlorine	Cl_2
bromine	Br_2
iodine	I_2

Naming Molecular Compounds



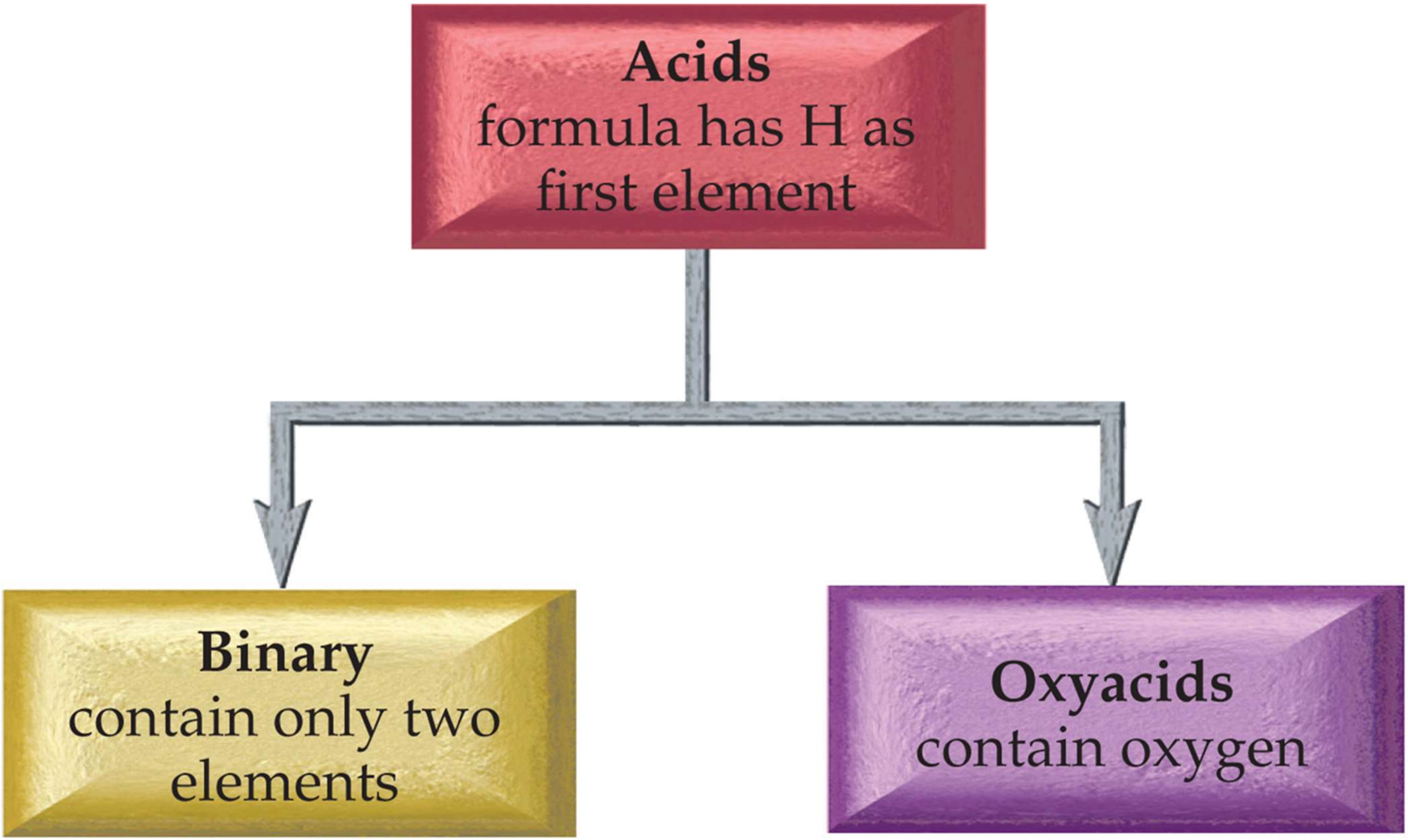
The first step in naming a molecular compound is identifying it as one.

Remember, nearly all molecular compounds form from two or more ***nonmetals.***

We learn how to name binary (two-element) molecular compounds. Their names have the following form:

The ***<u>prefixes</u>*** *given to each element indicate the number of atoms present.*

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


Acids
formula has H as first element
Binary
contain only two elements
Oxyacids
contain oxygen

Binary acids are composed of hydrogen and a nonmetal. The names for binary acids have the following form:

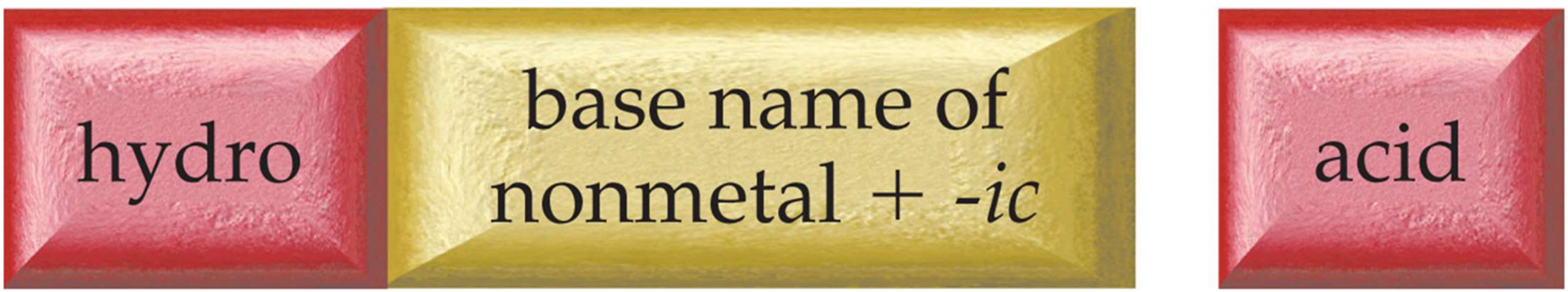
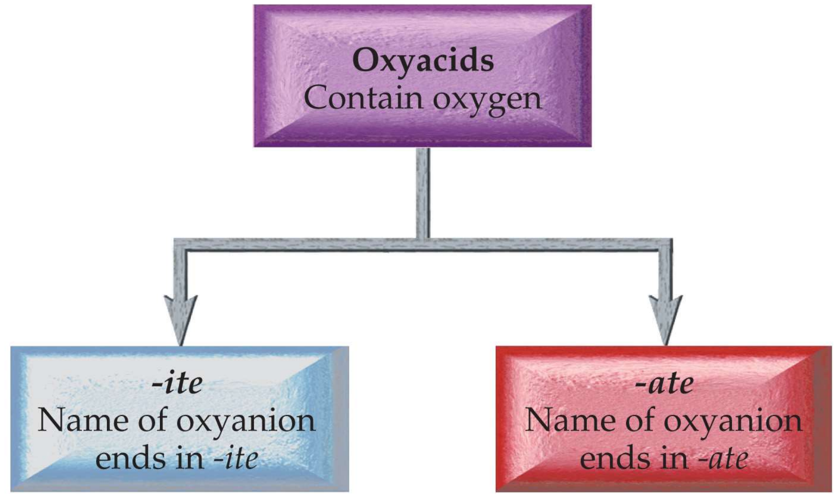


TABLE 5.7 Names of Some Common Oxyacids and Their Oxyanions

Acid Formula	Acid Name	Oxyanion Name	Oxyanion Formula
HNO_2	nitrous acid	nitrite	NO_2^-
HNO_3	nitric acid	nitrate	NO_3^-
H_2SO_3	sulfurous acid	sulfite	SO_3^{2-}
H_2SO_4	sulfuric acid	sulfate	SO_4^{2-}
$HClO_2$	chlorous acid	chlorite	ClO_2^-
$HClO_3$	chloric acid	chlorate	ClO_3^-
$HC_2H_3O_2$	acetic acid	acetate	$C_2H_3O_2^-$
H_2CO_3	carbonic acid	carbonate	CO_3^{2-}



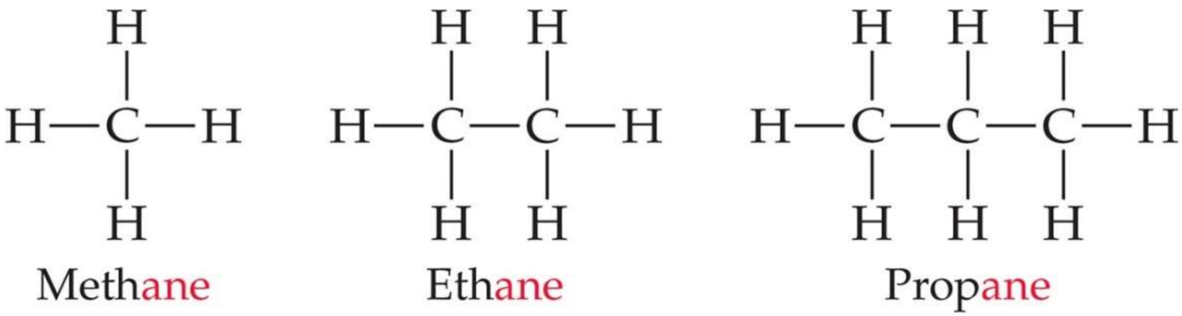
The names of acids containing oxyanions ending with -ite take this form:

base name of oxyanion + *-ous* acid

The names of acids containing oxyanions ending with -ate take this form:

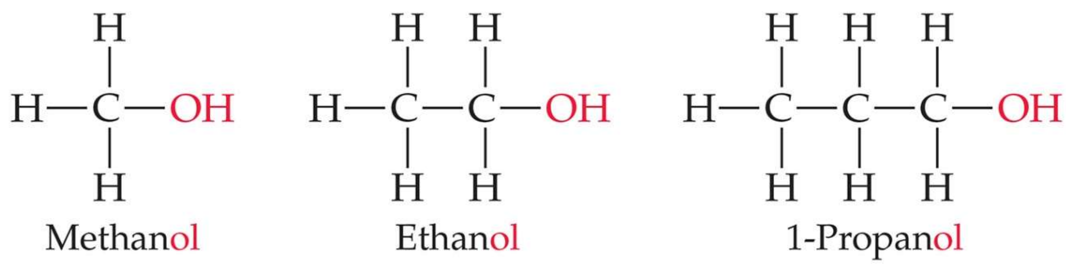
base name of oxyanion + *-ic* acid

2.9 Some Simple Organic Compounds



- **Organic chemistry** is the study of **carbon**.
- Organic chemistry has its own system of nomenclature.
- The simplest hydrocarbons (compounds containing only carbon and hydrogen) are **alkanes**.
- The first part of the names just listed correspond to the number of carbons (**meth-** = 1, **eth-** = 2, **prop-** = 3, etc.).
- It is followed by -**ane**.

Nomenclature of Alcohols



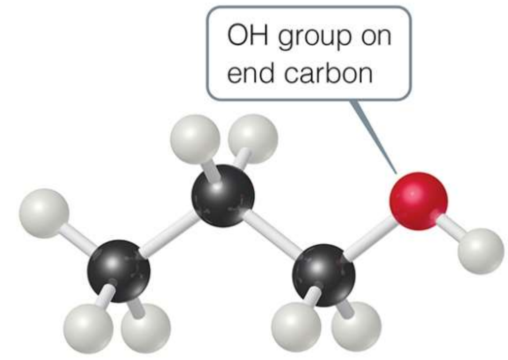
- When a hydrogen in an alkane is replaced with something else (a **functional group**, like $-OH$ in the compounds above), the name is derived from the name of the alkane.
- The ending denotes the type of compound.
  - An **alcohol** ends in **-ol**.

Nomenclature Isomers: Alcohols

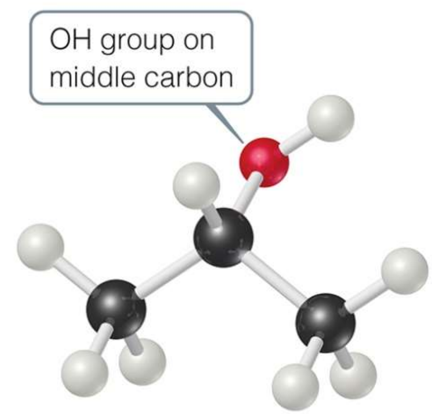
- When two or more molecules have the same chemical formula, but different structures, they are called **isomers**.
- 1-Propanol and 2-propanol have the oxygen atom connected to different carbon atoms. Both have the same empirical and molecular formula

(C_3H_8O). They have different structural formulas:

$$\text{1-Propanol}: CH_3CH_2OH$$
$$\text{2-Propanol}: CH_3CH(OH)CH_3$$



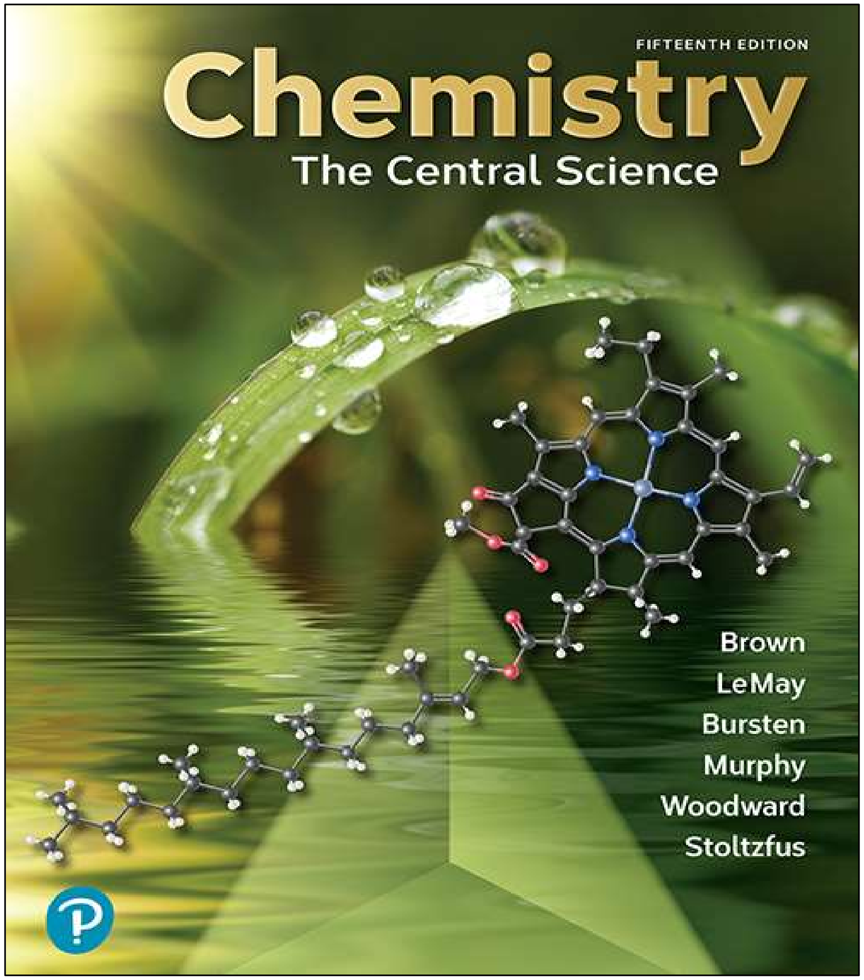
1-Propanol



2-Propanol

Questions?

Chemistry: The Central Science

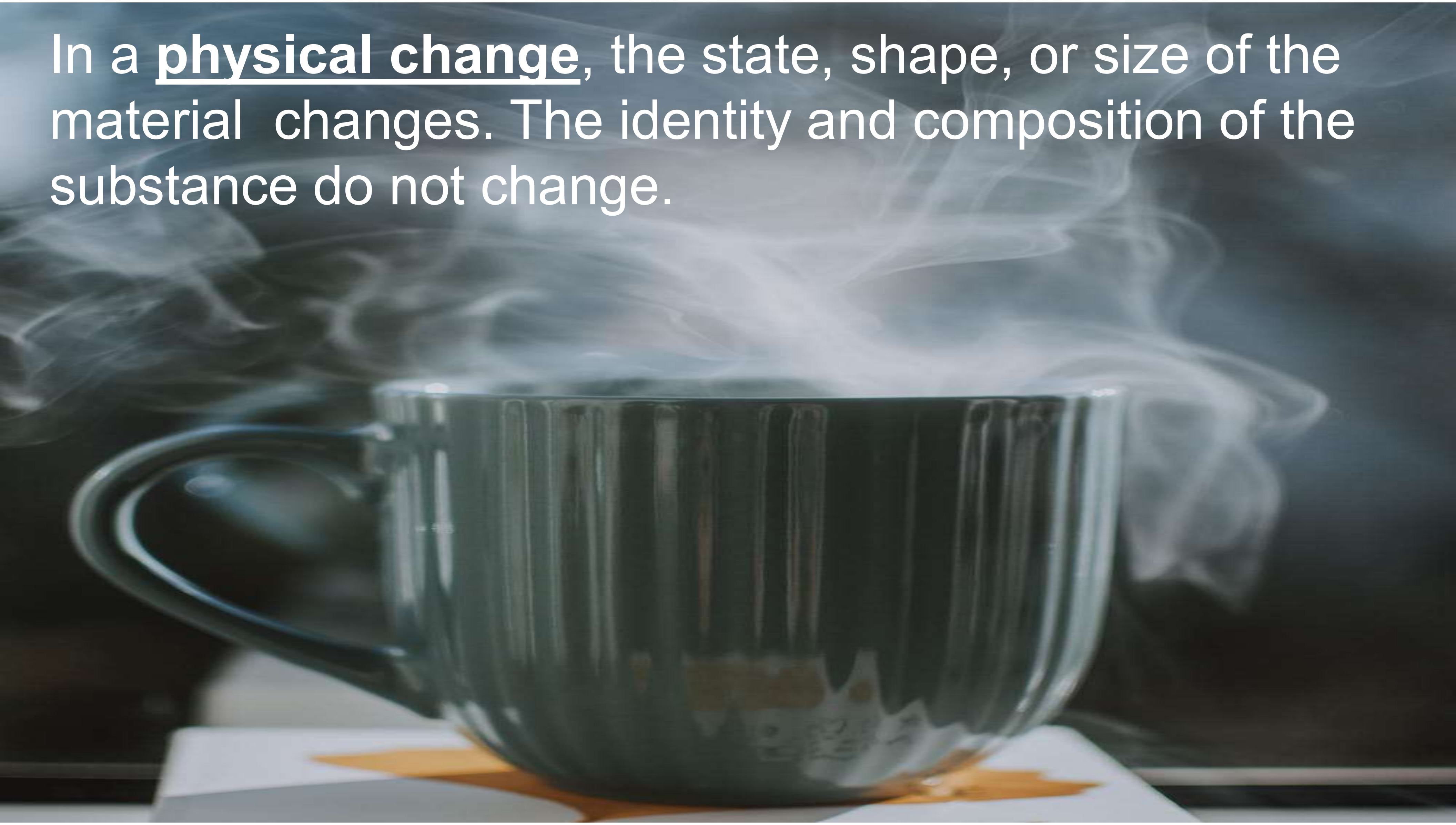


Chapter 3

Chemical Reactions and Reaction Stoichiometry


Let’s Review

In a **<u>physical change</u>**, the state, shape, or size of the material changes. The identity and composition of the substance do not change.



In a **chemical change**, reacting substances form new substances with different compositions and properties.

Chemical Changes	Physical Changes
Rusting nail	Melting ice
Bleaching a stain	Boiling water
Burning a log	Sawing a log in half
Tarnishing silver	Tearing paper
Fermenting grapes	Breaking a glass
Souring of milk	Pouring milk

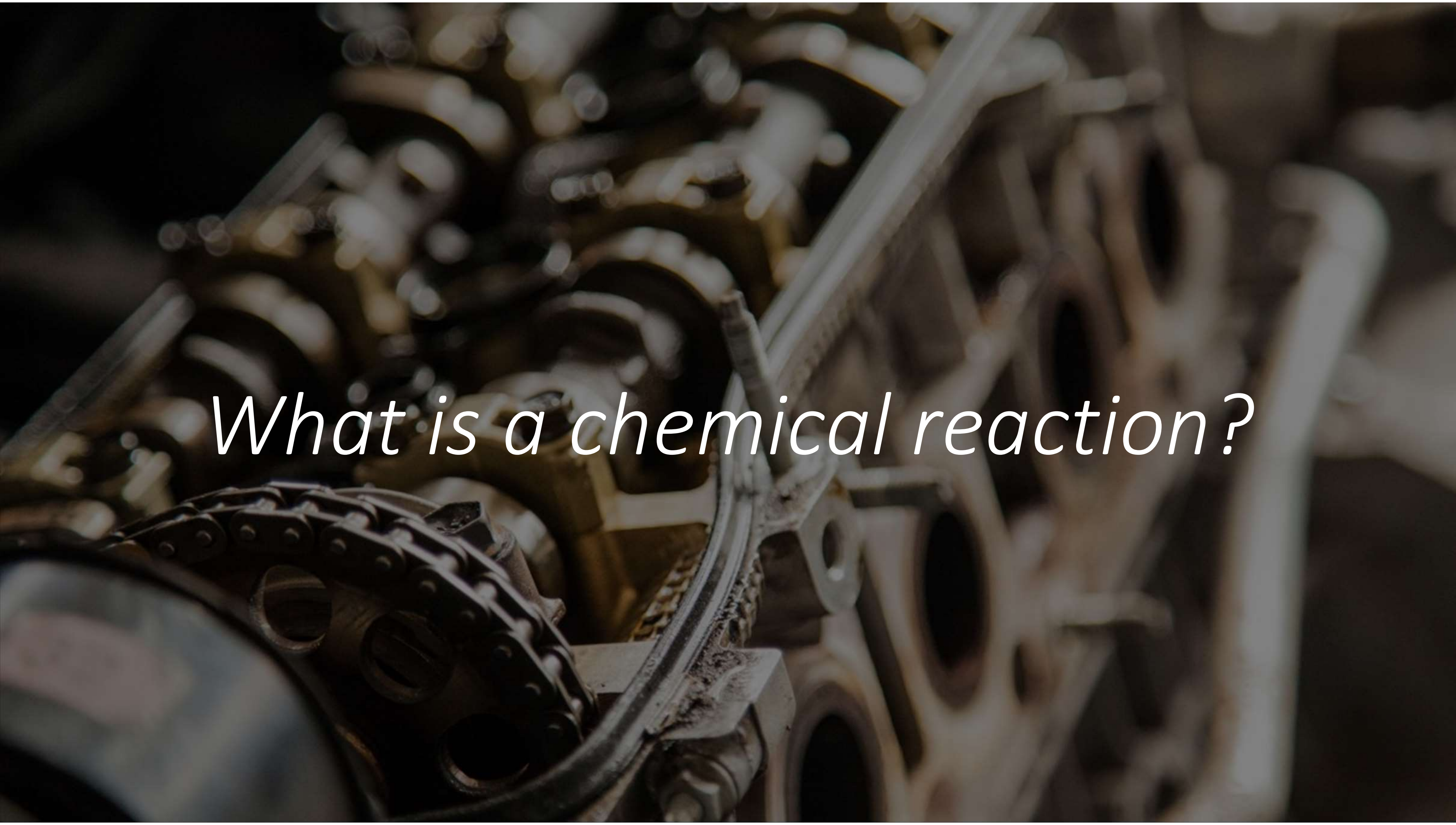
Is there evidence for chemical change?

Evidence of a Chemical Change

Only chemical analysis that shows that the initial substances have changed into other substances conclusively proves that a chemical reaction has occurred.

Chemical changes may occur without any obvious signs, yet chemical analysis may show that a reaction has indeed occurred.

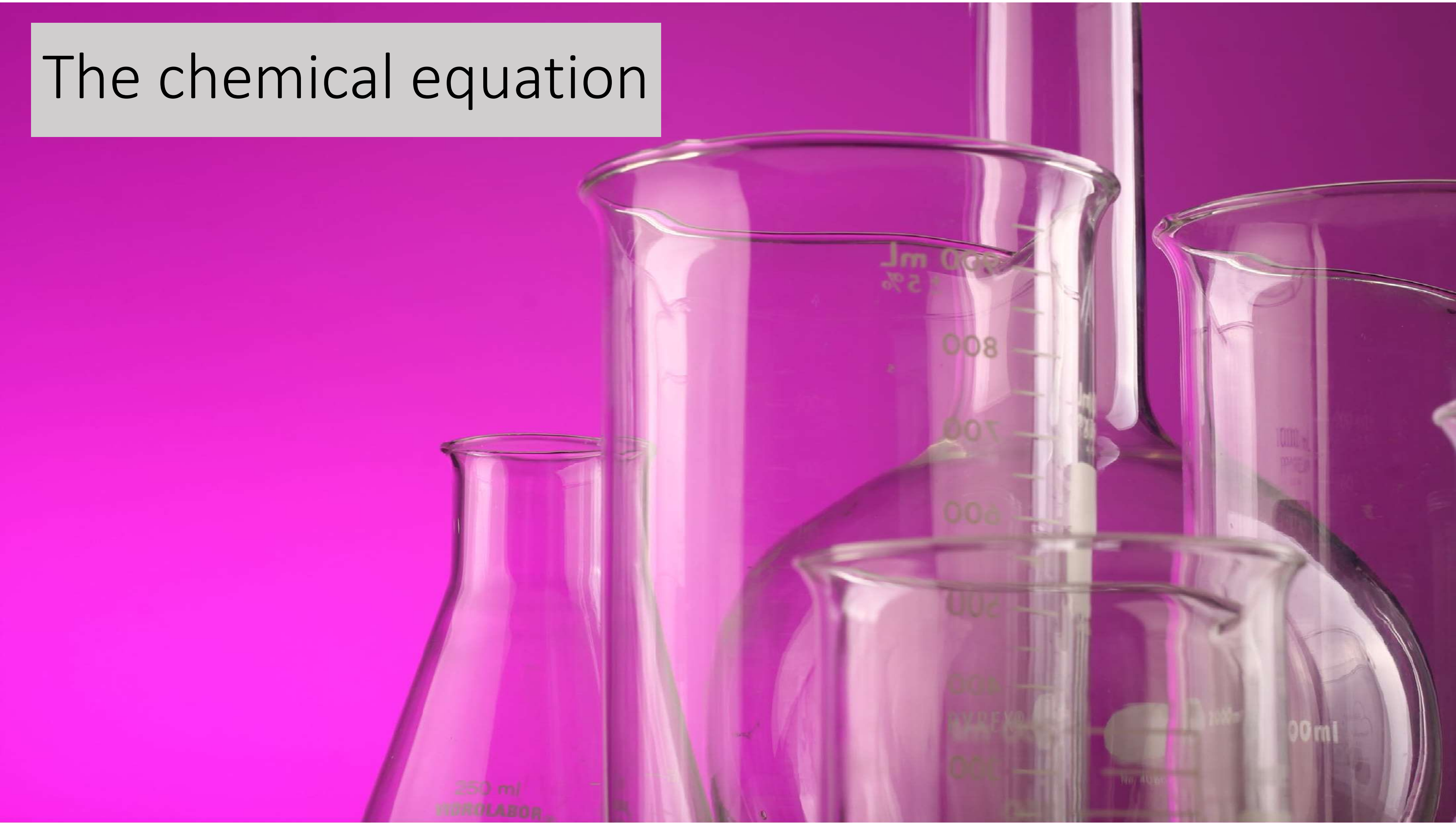
Only then you can state a chemical reaction has occurred.


What is a chemical reaction?

Chemical Reaction

*In a **chemical reaction**, a chemical change produces one or more new substances. There is a change in the composition of one or more substances.*

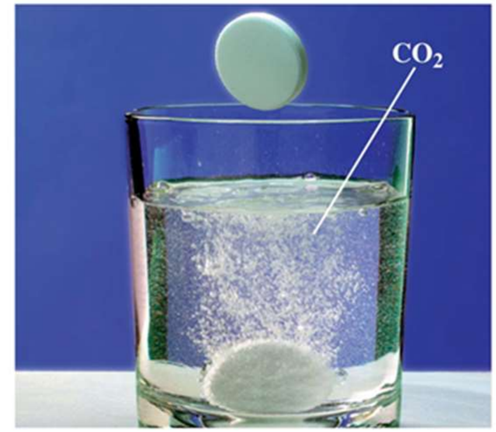
The chemical equation



Deciphering chemical equations

$NaHCO_{3\,(s)}$ + $HCl_{(l)}$ $\longrightarrow$ $CO_{2\,(g)}$ + $H_2O_{(l)}$ + $NaCl_{(aq)}$

Reactants Products



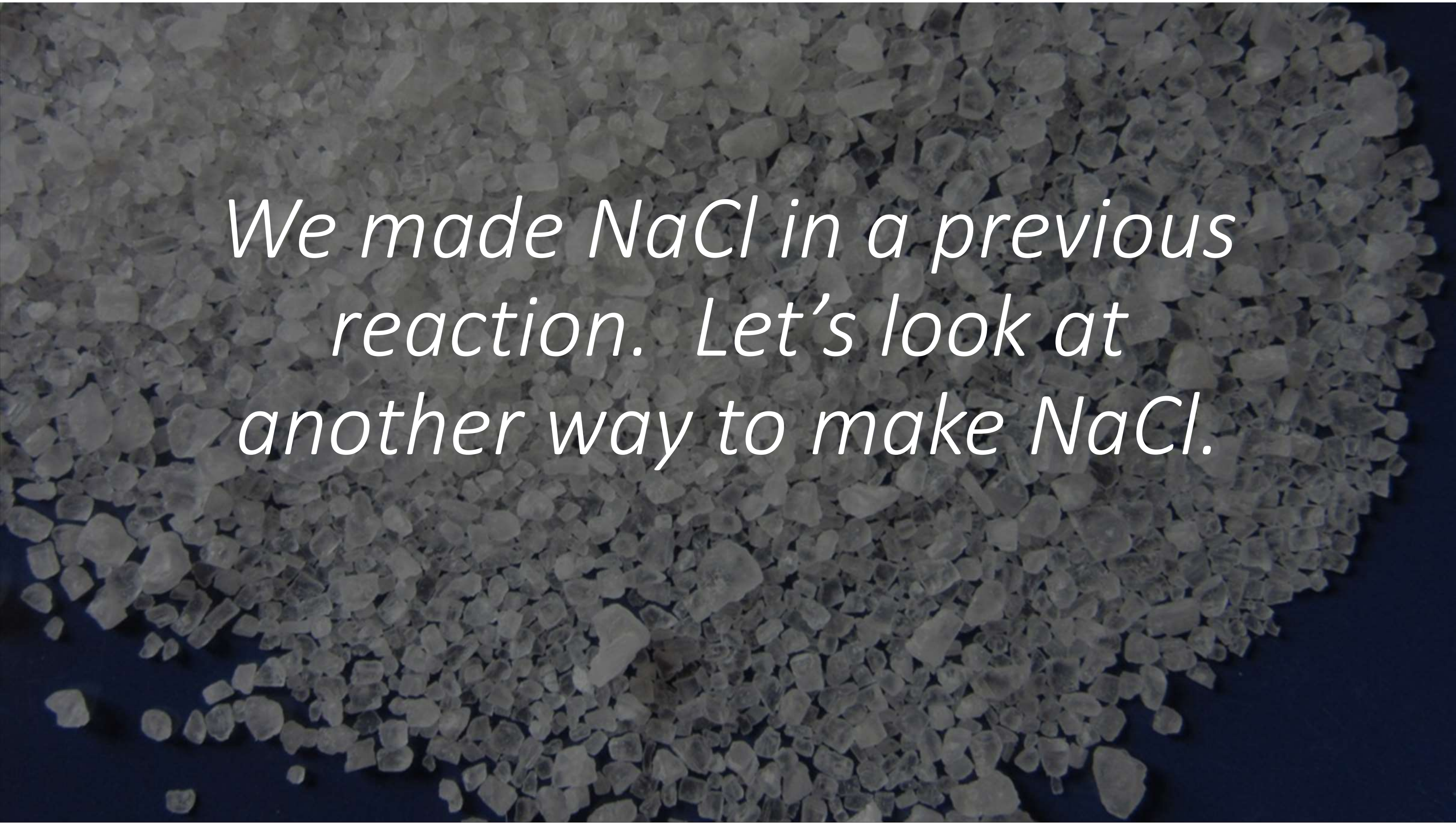
$NaHCO_3$

Symbols in chemical equations show:

1. the states of the reactants.
2. the states of the products.
3. the reaction conditions.

Symbol	Meaning
+	Separates two or more formulas
$\longrightarrow$	Reacts to form products
Δ	The reactants are heated
(*s*)	Solid
(*l*)	Liquid
(*g*)	Gas
(*aq*)	Aqueous

Important -> old bonds are broken, and new bonds are formed


We made NaCl in a previous reaction. Let’s look at another way to make NaCl.

New Reaction

$$Na_{(s)} + Cl_{2\,(g)} \longrightarrow \boxed{NaCl}_{(s)}$$

Reactants Products

Notice different states of matter for the same chemical.

Old Reaction

$$NaHCO_{3\,(s)} + HCl_{(l)} \longrightarrow CO_{2\,(g)} + H_2O_{(l)} + \boxed{NaCl}_{(aq)}$$

Reactants Products

There is a problem with one of the reactions?

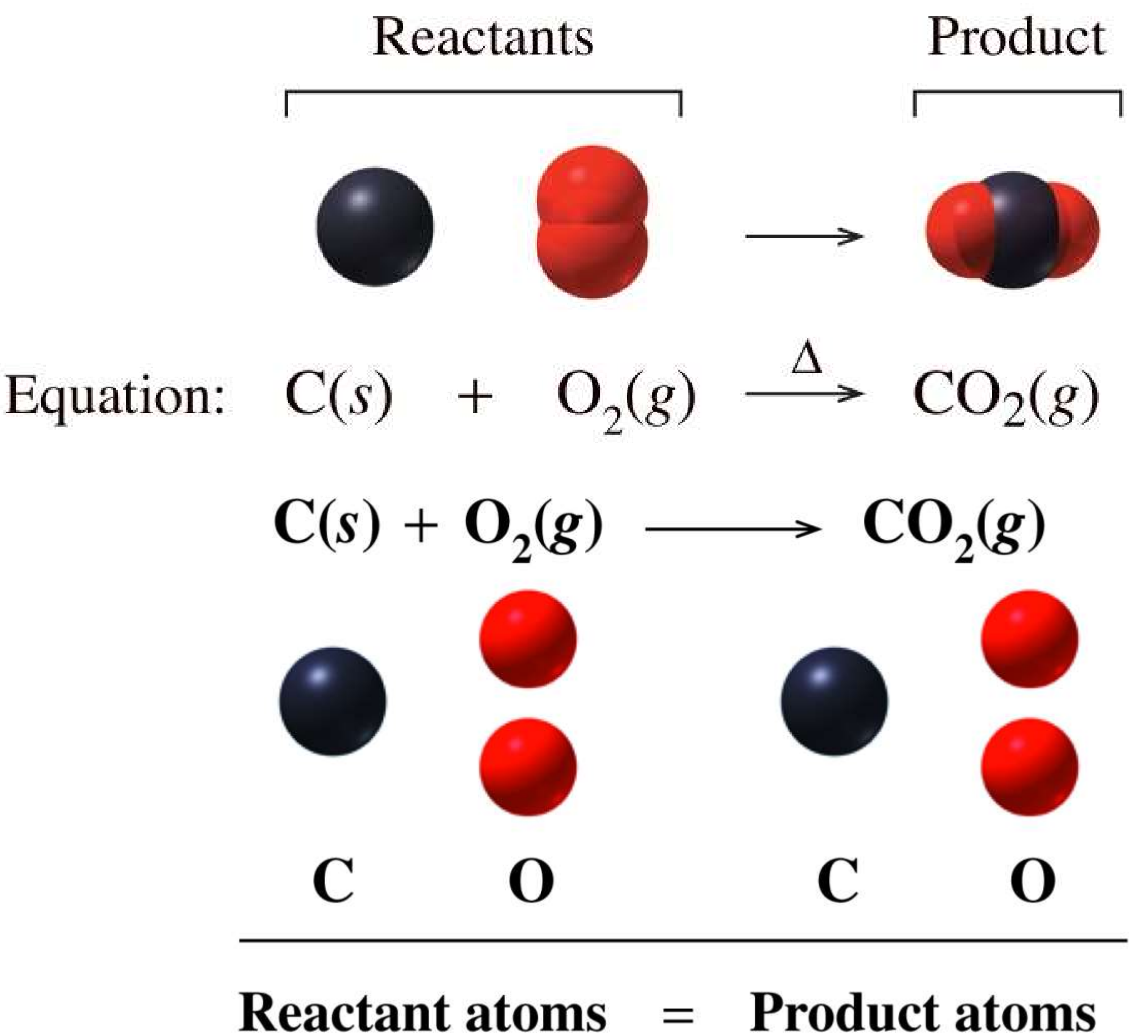
Violation of the law of conservation of matter

$$Na_{(s)} + Cl_{2\,(g)} \longrightarrow NaCl_{(s)}$$

Reactants Products

Matter can not be created or destroyed. Matter ***MUST BE CONSERVED!***


Balancing Equations


Reactants
Product
Equation: $C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)$
$C(s) + O_2(g) \longrightarrow CO_2(g)$
C
O
C
O
Reactant atoms = Product atoms

Violation of the law of conservation of matter

$$Na_{(s)} + Cl_{2\,(g)} \longrightarrow NaCl_{(s)}$$

Reactants Products

The number of chlorines do not equal on both sides.

Balancing Equations

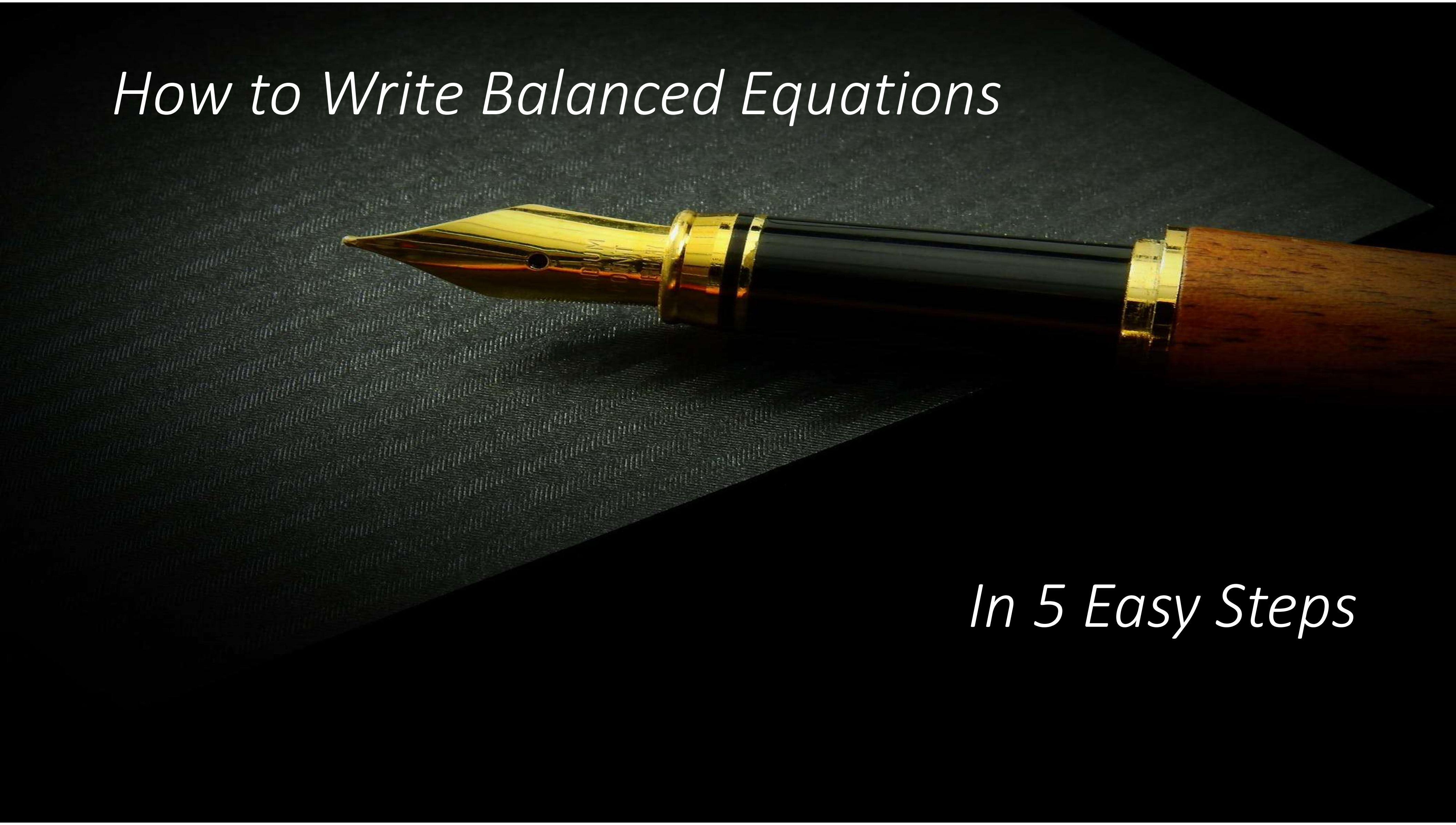
$$\_\_Na_{(s)} + \_\_Cl_{2\,(g)} \longrightarrow \_\_NaCl_{(s)}$$

Reactants Products

In a balanced chemical equation, numbers called coefficients are used in front of one or more formulas to balance the number of atoms.

The total number of each atom type on the reactant side MUST EQUAL to the total number of each atom type on the product side.

How to Write Balanced Equations



In 5 Easy Steps

How to Write Balanced Chemical Equations

1. *Write a skeletal equation by writing correct chemical formulas for each of the reactants and products. Review Chapter 5 for nomenclature rules. (If a skeletal equation is provided, skip this step and go to Step 2.)*
2. *If an element occurs in only one compound on both sides of the equation, balance it first. If there is more than one such element, balance metals before nonmetals.*
3. *If an element occurs as a free element on either side of the chemical equation, balance it last. Always balance free elements by adjusting the coefficient on the free element.*
4. *If the balanced equation contains coefficient fractions, change these into whole numbers by multiplying the entire equation by the appropriate factor.*
5. *Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.*

How to Write Balanced Chemical Equations

1. *Write a skeletal equation by writing correct chemical formulas for each of the reactants and products. Review Chapter 5 for nomenclature rules. (If a skeletal equation is provided, skip this step and go to Step 2.)*

$$\_\_SiO_{2\,(s)} + \_\_C_{(s)} \longrightarrow \_\_SiC_{(s)} + \_\_CO_{(g)}$$

Reactants Products

This could have been written out in words: Silicon dioxide and carbon react to form silicon carbide and carbon monoxide. Balance the following equation.

How to Write Balanced Chemical Equations

2. *If an element occurs in only one compound on both sides of the equation, balance it first. If there is more than one such element, balance metals before nonmetals.*
3. *If an element occurs as a free element on either side of the chemical equation, balance it last. Always balance free elements by adjusting the coefficient on the free element.*

$$\_\_SiO_{2\,(s)} + \_\_C_{(s)} \longrightarrow \_\_SiC_{(s)} + \_\_CO_{(g)}$$

Reactants Products

How to Write Balanced Chemical Equations

4. *If the balanced equation contains coefficient fractions, change these into whole numbers by multiplying the entire equation by the appropriate factor.*

$$\_\_SiO_{2\,(s)} + \_\_C_{(s)} \longrightarrow \_\_SiC_{(s)} + \_\_CO_{(g)}$$

Reactants Products

How to Write Balanced Chemical Equations

5. *Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.*

$$\underline{1}\ SiO_{2\,(s)} + \underline{3}\ C_{(s)} \longrightarrow \underline{1}\ SiC_{(s)} + \underline{2}\ CO_{(g)}$$

Reactants Products

Let's come back to this!

$$\_\_Na_{(s)} + \_\_Cl_{2\,(g)} \longrightarrow \_\_NaCl_{(s)}$$

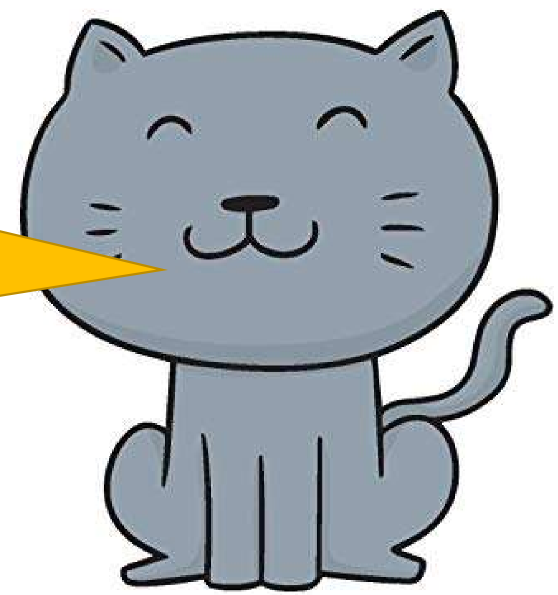
Reactants

Products

Is this correct? I changed the subscript, so it is balanced ☺

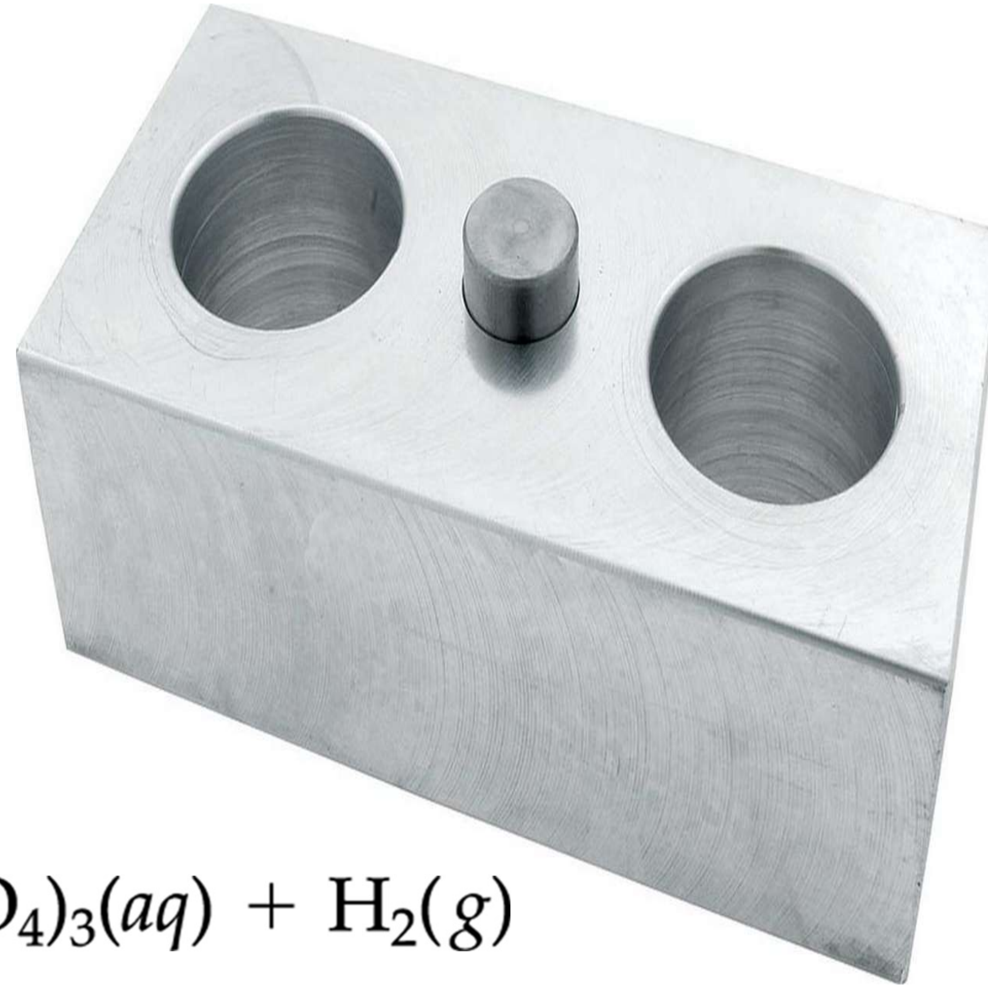
$$\_\_Na_{(s)} + \_\_Cl_{(g)} \longrightarrow \_\_NaCl_{(s)}$$

<u>change only the coefficients to balance</u> a chemical equation; never change the subscripts.



Example

The Reaction of Solid Aluminum with Aqueous Sulfuric Acid to Form Aqueous Aluminum Sulfate and Hydrogen Gas.



$$Al(s) + H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2(g)$$

Easy. Balance polyatomic ions that are the same on both sides of the equation as a unit.

$$\mathrm{Al}(s) + \mathrm{H_2SO_4}(aq) \longrightarrow \mathrm{Al_2(SO_4)_3}(aq) + \mathrm{H_2}(g)$$

What type of reactions are possible?

3.2 Simple Patterns of Chemical Reactivity

- There are many different types of chemical reactions.
- After you master this chapter, three broad classes of reactions can be predicted:
  1. Combination reactions
  2. Decomposition reactions
  3. Combustion reactions

Combination Reactions

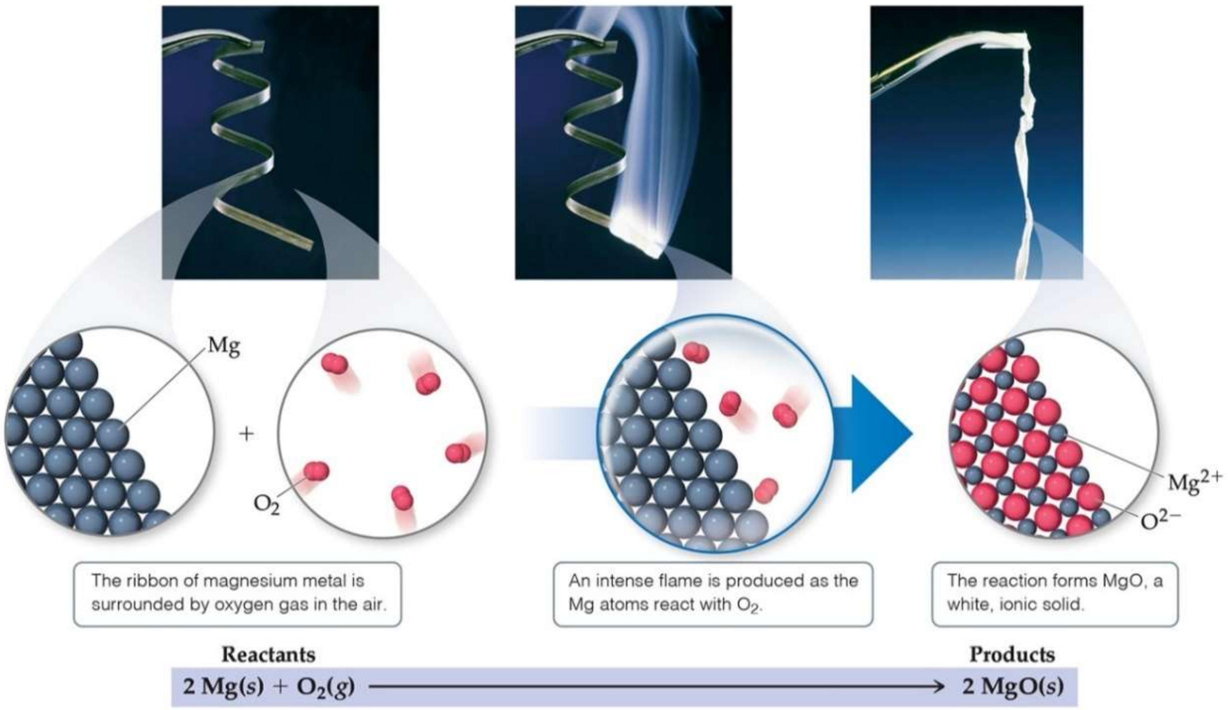
- Two or more substances react to form one product.

Table 3.1: Combination and Decomposition Reactions

Combination Reactions	
$A + B \rightarrow C$ $C(s) + O_2(g) \rightarrow CO_2(g)$ $N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$ $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$	Two or more reactants combine to form a single product. Many elements react with one another in this fashion to form compounds.

Combination Reaction Prediction: A Metal and a Nonmetal

- You should be able to predict the product of a combination reaction between a metal and a nonmetal, like the one below. (**Hint:** Use common charges for Groups)



Decomposition Reactions

- In a **decomposition reaction,** one substance breaks down into two or more substances.
- In the air bag, solid sodium azidegas quickly on impact.
- (NaN_3) releases nitrogen (N_2)

Table 3.1: Combination and Decomposition Reactions

Decomposition Reactions	
$C \rightarrow A + B$ $2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$ $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$ $Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(g)$	A single reactant breaks apart to form two or more substances. Many compounds react this way when heated.

Predicting Decomposition Reactions: Heating a Metal Carbonate

- Metal carbonates decompose when heated to give off carbon dioxide (CO_2) and a metal oxide.
  - CaO is a major raw material for cement production.
- Balancing these equations is based on the charge of the metal.

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

Combustion Reactions

A combustion reaction is characterized by the burning of a carbon-containing compound in the presence of oxygen to form carbon dioxide and water.

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O$$

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

3.3 Formula Weight (FW)

- A **formula weight** is the sum of the atomic weights for the atoms in a chemical formula.
- This is the quantitative significance of a formula.
- For an element like sodium, Na, the formula weight is the atomic weight (23.0 amu). Found on the periodic table.
- For an **ionic compound**, use the empirical formula.
- The formula weight of sulfuric acid, H_2SO_4, would be

$$2(\text{AW of H}) + 1(\text{AW of S}) + 4(\text{AW of O})$$

$$2(1.0\ \text{amu}) + 32.1\ \text{amu} + 4(16.0\ \text{amu})$$

$$\text{FW}(H_2SO_4) = 98.1\ \text{amu}$$

Molecular Weight (MW)

- If the substance is a **molecule**, the formula weight is also called its **molecular weight**.
- A **molecular weight** is the sum of the atomic weights of the atoms in a molecule.
- For glucose, which has a molecular formula of $C_6H_{12}O_6$, the molecular weight is
  - $6(\text{AW of C}) + 12(\text{AW of H}) + 6(\text{AW of O})$
  - $6(12.0\ \text{amu}) + 12(1.0\ \text{amu}) + 6(16.0\ \text{amu})$
  - $\text{MW}(C_6H_{12}O_6) = 180.0\ \text{amu}$

Percent Composition

- One can find the percentage of the mass of each element of a compound from the chemical formula.
- Use each of the elements in the compound with this equation:

$$\%\ \text{Element} = \frac{\begin{pmatrix}\text{number of atoms} \\ \text{of the element}\end{pmatrix}\begin{pmatrix}\text{atomic weight} \\ \text{of the element}\end{pmatrix}}{\text{FW of the substance}} \times 100$$

Percent Composition

- The percentage of carbon in glucose $\left(C_6H_{12}O_6\right)$ is:
  C = 12.0 amu
  6 carbons in glucose
  MW glucose = 180.0 amu
  [6(12) + 12(1) + 6(16)]

$$\%C = \frac{(6)(12.0\text{ amu})}{(180.0\text{ amu})}$$

$$= \frac{72.0\text{ amu}}{180.0\text{ amu}} \times 100$$

$$= 40.0\%$$

A **collection term** states a specific number of items.

Moles

Samples of 1 Mole Quantities

1 mole of C atoms = 6.02×10^{23} C atoms

1 mole of Al atoms = 6.02×10^{23} Al atoms

1 mole of S atoms = 6.02×10^{23} S atoms

1 mole of H_2O molecules = 6.02×10^{23} H_2O molecules

1 mole of CCl_4 molecules = 6.02×10^{23} CCl_4 molecules

One mole on any substance is equal to Avogadro's Number!!!

You can use Avogadro's Number to convert to moles

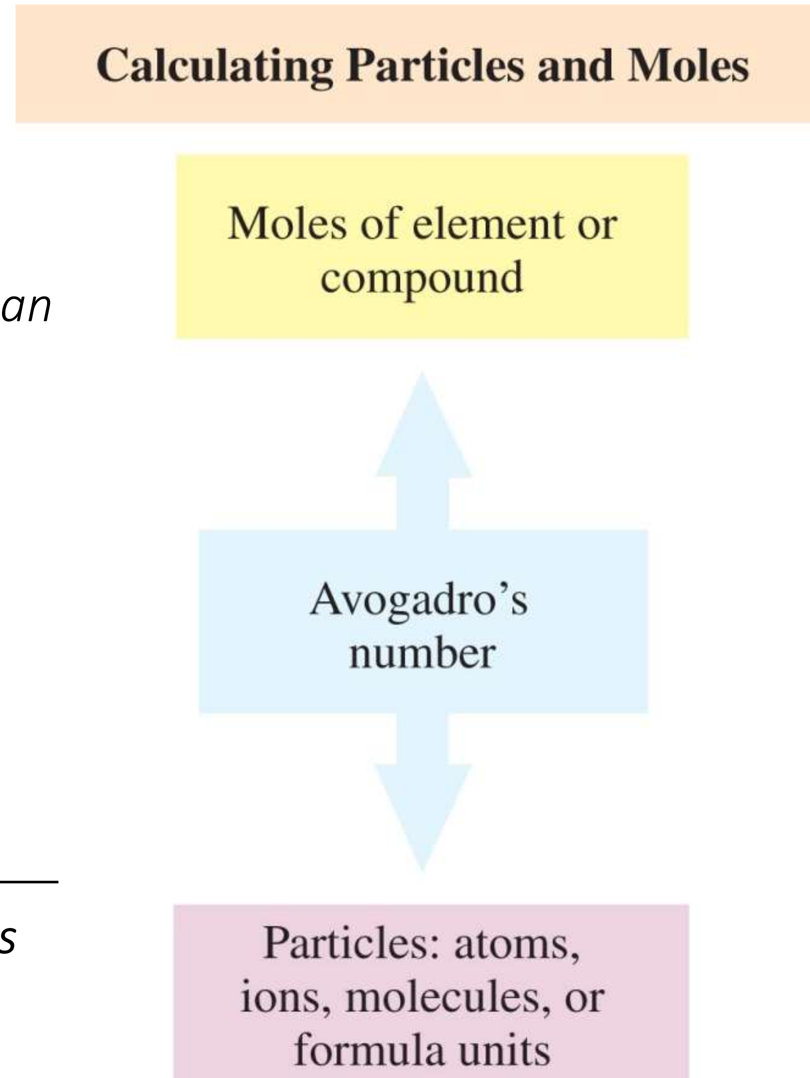
Avogadro's number, 6.02×10^{23}, can be written as an equality and two conversion factors.

Equality:

$$1 \text{ mole of } X = 6.02 \times 10^{23} \text{ particles of } X$$

Conversion Factors:

$$\frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mole}} \quad \text{and} \quad \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ particles}}$$



Molar Mass

Guide to Calculating Molar Mass of a Compound

The molar mass of a compound is the sum of the molar masses of the elements in the formula.

Guide to Calculating Molar Mass

1. Obtain the molar mass of each element.
2. Multiply each molar mass by the number of moles (subscript) in the formula.
3. Calculate the molar mass by adding the masses of the elements.

Build a Table

Example: Calculate the molar mass of $CaCl_2$.

Element	Number of Moles	Atomic Mass	Total Mass
Ca	*1*	*40.1 g/mole*	*40.1 g*
Cl	*2*	*35.5 g/mole*	*71.0 g*
$CaCl_2$			***111.1 g***

Guide to Calculating Molar Mass

1 Obtain the molar mass of each element.

2 Multiply each molar mass by the number of moles (subscript) in the formula.

3 Calculate the molar mass by adding the masses of the elements.

Example: Calculate the molar mass of K_3PO_4.

Element	Number of Moles	Atomic Mass	Total Mass in K_3PO_4
K	3	39.1 g/mole	117.3 g
P	1	31.0 g/mole	31.0 g
O	4	16.0 g/mole	64.0 g
K_3PO_4			**212.3 g**

Guide to Calculating Molar Mass

1. Obtain the molar mass of each element.
2. Multiply each molar mass by the number of moles (subscript) in the formula.
3. Calculate the molar mass by adding the masses of the elements.

Why is molar mass important?

Allows you to covert between mass and moles

Summary of Mole Relationships

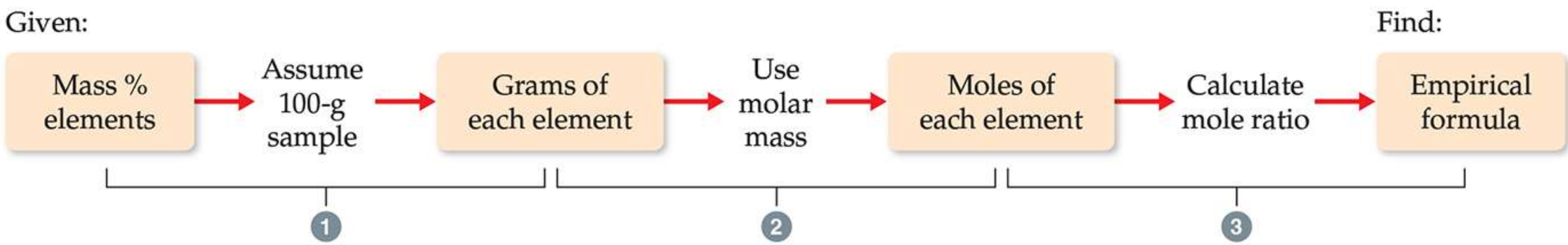
Table 3.2 Mole Relationship

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	6.02×10^{23} N atoms
Molecular nitrogen or "dinitrogen"	N_2	28.0	28.0	$\begin{cases} 6.02 \times 10^{23}\ N_2 \text{ molecules} \\ 2(6.02 \times 10^{23})\ \text{N atoms} \end{cases}$
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	$\begin{cases} 6.02 \times 10^{23}\ BaCl_2 \text{ formula units} \\ 6.02 \times 10^{23}\ Ba^{2+} \text{ ions} \\ 2(6.02 \times 10^{23})\ Cl^- \text{ ions} \end{cases}$

[a]Recall that the mass of an electron is more than 1800 times smaller than the masses of the proton and the neutron; thus, ions and atoms have essentially the same mass

- One mole of atoms, ions, or molecules contains Avogadro's number of those particles.
- The number of atoms of an element in a mole is the subscript in a formula (number of atoms of that element in the formula) times Avogadro's number.

3.5 Empirical Formulas from Analysis



- One can determine the empirical formula from the percent composition by following these three steps. Assume the mass % is based on a 100 g sample.
    1. Convert mass% to grams
    2. Convert grams to moles
    3. Calculate the mole ratio

Determining Empirical Formulas—an Example

- The compound para-aminobenzoic acid (you may have seen it listed as PABA on your bottle of sunscreen) is composed of carbon (61.31%), hydrogen (5.14%), nitrogen (10.21%), and oxygen (23.33%). Find the empirical formula of PABA.
  - The four elements are C, H, N, and O
  - The % become grams
    - C = 61.31 g
    - H = 5.14 g
    - N = 10.21 g
    - O = 23.33 g

Determining Empirical Formulas—an Example

- Convert grams to moles assuming the 100.00 g of para-aminobenzoic acid

$$C: 61.31\,\text{g} \times \frac{1\,\text{mol}}{12.01\,\text{g}} = 5.105\,\text{mol C}$$

$$H: 5.14\,\text{g} \times \frac{1\,\text{mol}}{1.01\,\text{g}} = 5.09\,\text{mol H}$$

$$N: 10.21\,\text{g} \times \frac{1\,\text{mol}}{14.01\,\text{g}} = 0.7288\,\text{mol N}$$

$$O: 23.33\,\text{g} \times \frac{1\,\text{mol}}{16.00\,\text{g}} = 1.456\,\text{mol O}$$

Determining Empirical Formulas—an Example

- Calculate the mole ratio for each element by dividing by the smallest number of moles, i.e., 0.7288 mol:

$$\text{C:}\ \frac{5.105\ \text{mol}}{0.7288\ \text{mol}} = 7.005 \approx 7$$

$$\text{H:}\ \frac{5.09\ \ \text{mol}}{0.7288\ \text{mol}} = 6.984 \approx 7$$

$$\text{N:}\ \frac{0.7288\ \text{mol}}{0.7288\ \text{mol}} = 1.000$$

$$\text{O:}\ \frac{1.458\ \text{mol}}{0.7288\ \text{mol}} = 2.001 \approx 2$$

These become the subscripts for the empirical formula:

$$C_7H_7NO_2$$

Molecular Formulas From Empirical Formulas

- Remember, the number of atoms in a molecular formula is a multiple of the number of atoms in an empirical formula.
- If we find the empirical formula and know a molar mass (molecular weight) for the compound, we can find the molecular formula.

$$\text{Whole number multiple} = \frac{\text{Molecular weight (MW)}}{\text{Empirical formula weight (FW)}}$$

Determining a Molecular Formula—an Example

- The empirical formula of a compound was found to be C H. It has a molar mass of 78 g/mol. What is its molecular formula?
- Solution:

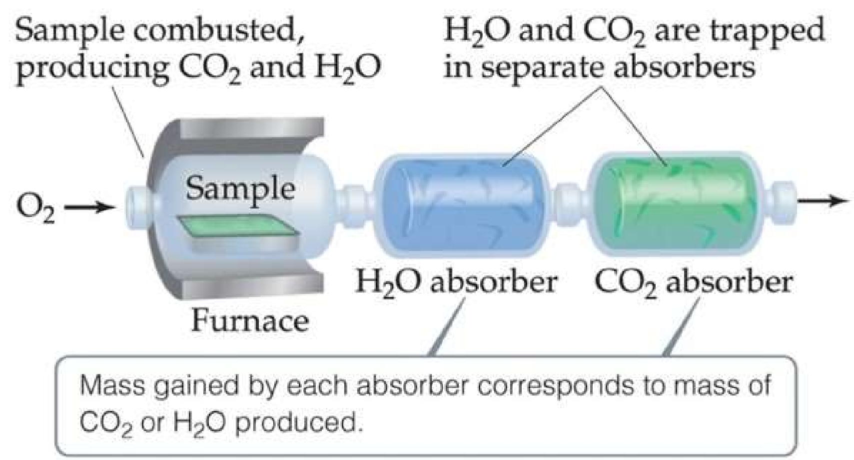
$$C + H = 1(12) + 1(1) = 13 \qquad CH$$

$$\text{Whole-number multiple } = 78/13 = 6$$

$$\text{The molecular formula is } C_6H_6. \qquad (CH)_6$$

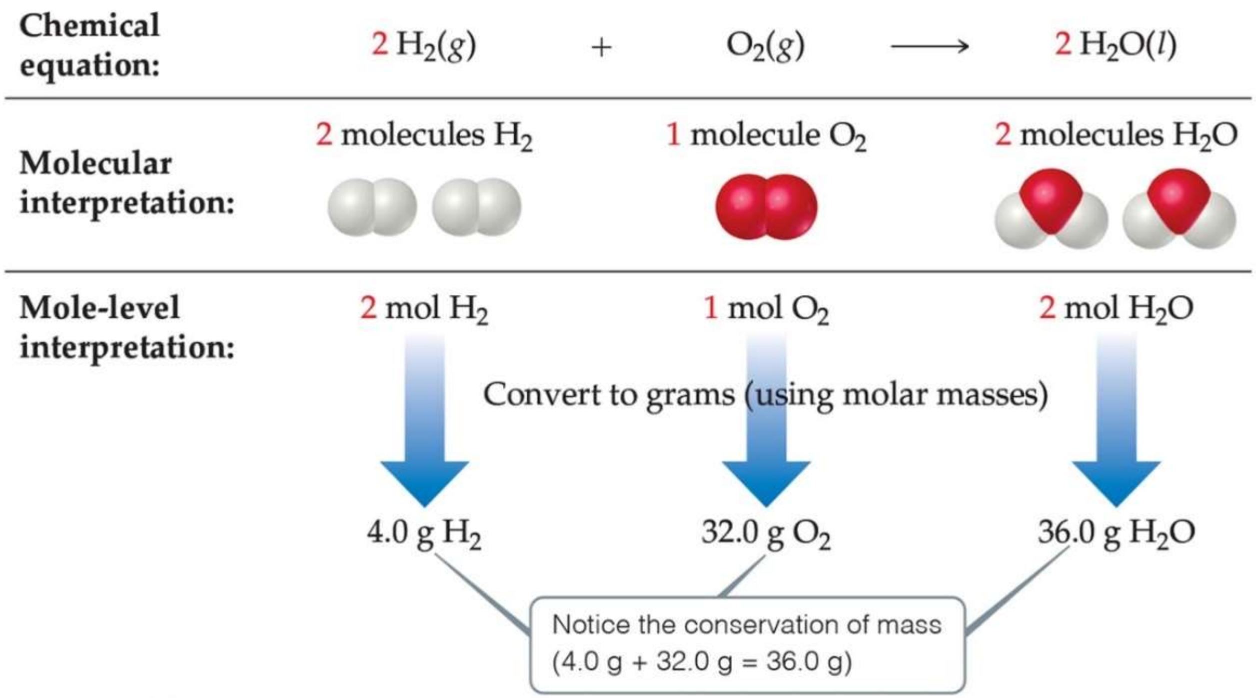
Combustion Analysis

- Compounds containing C, H, and O are routinely analyzed through combustion in a chamber. Once element mole ratio is known, formula can be determined.



1. Mass of C is determined from the mass of CO_2 produced.
2. Mass of H is determined from the mass of H_2O produced.
3. Mass of O is determined by the difference of the mass of the compound used and the total mass of C and H. $\mathbf{g(O) = g(sample) - g(C) - g(H)}$
   - Note: The mass of O in the compound cannot be determined from CO_2 and H_2O because oxygen is added during the combustion.

3.6 Quantitative Information from Balanced Equations

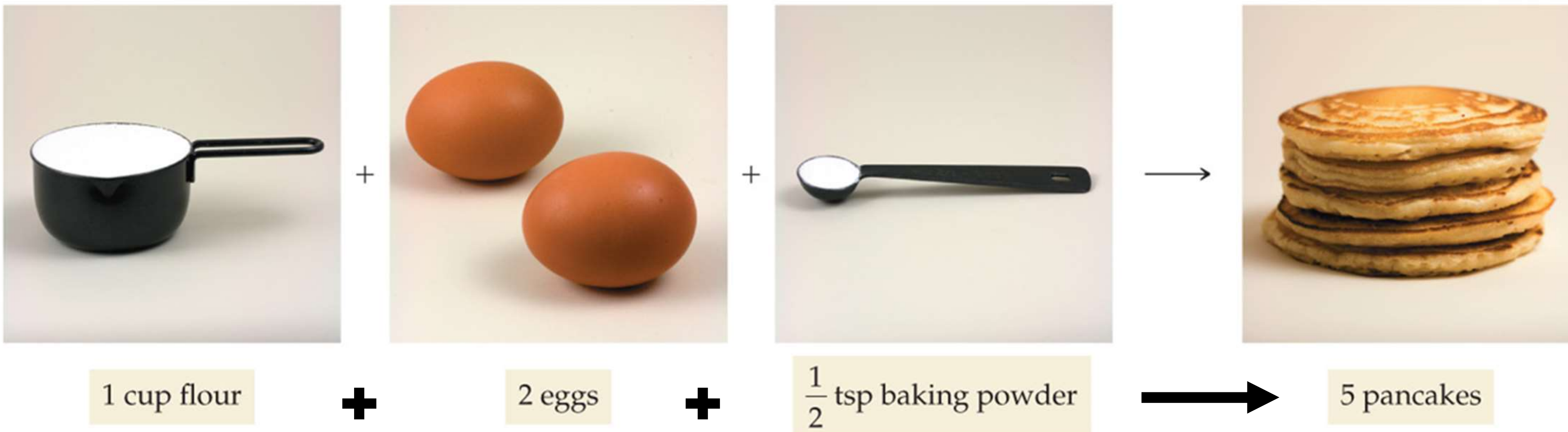


- The coefficients in the balanced equation show
  - Relative numbers of **molecules** of reactants and products
  - Relative numbers of **moles** of reactants and products, which can be converted to **mass**


How many of you cook? Can you make pancakes?

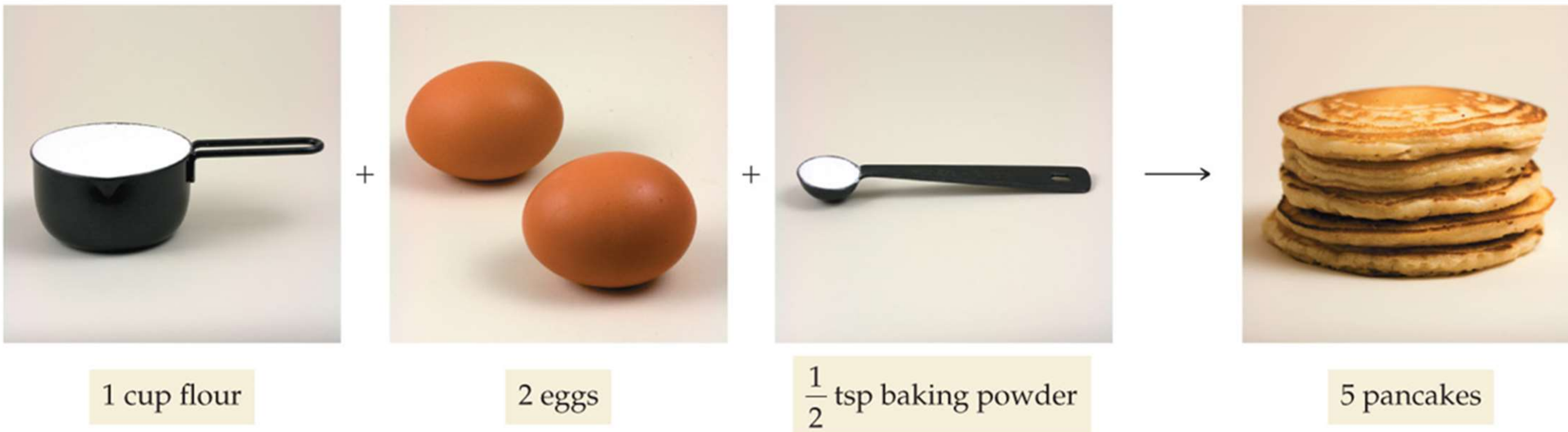
Making Pancakes

1 cup flour + 2 eggs + $\frac{1}{2}$ tsp baking powder $\longrightarrow$ 5 pancakes

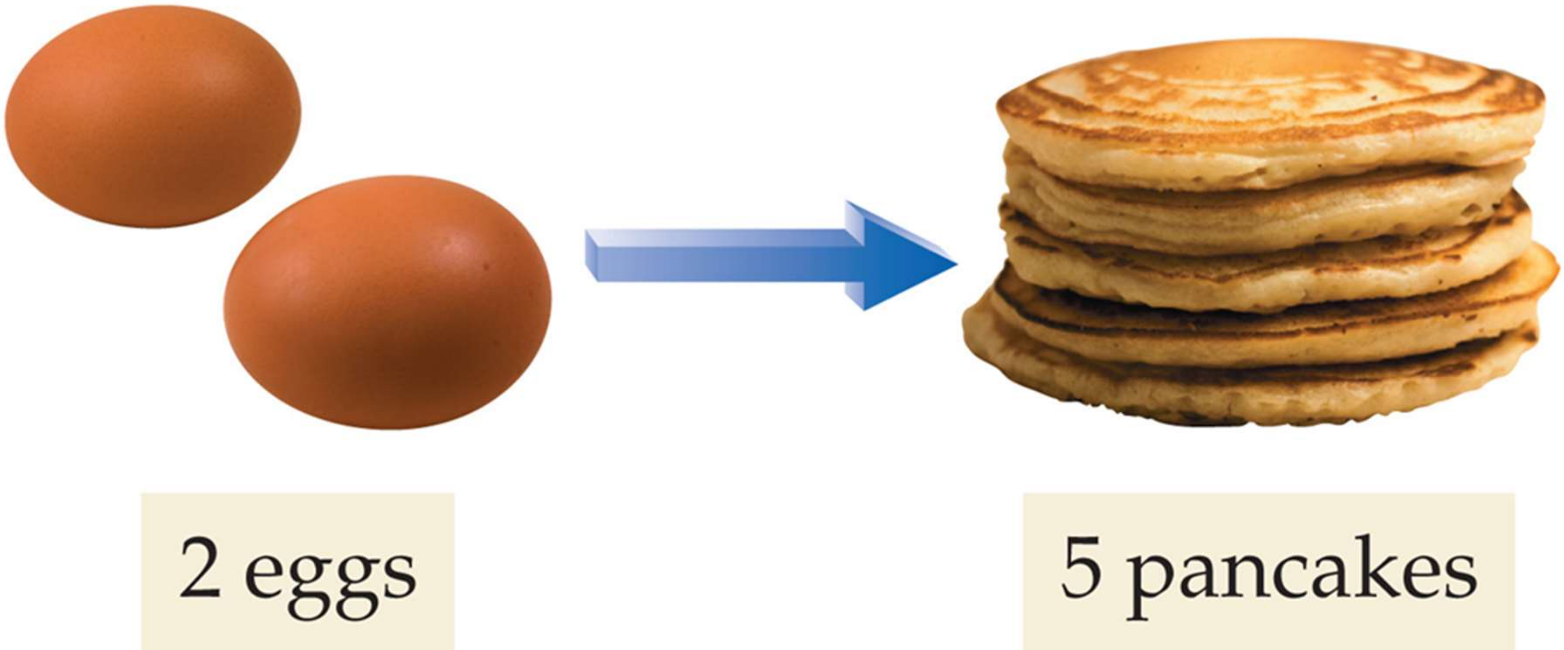


A recipe gives numerical relationships between the ingredients and the number of pancakes.

$$1 \text{ cup flour} + 2 \text{ eggs} + \tfrac{1}{2} \text{ tsp baking powder} \longrightarrow 5 \text{ pancakes}$$



How many eggs are needed? How many pancakes will be produced in this recipe?



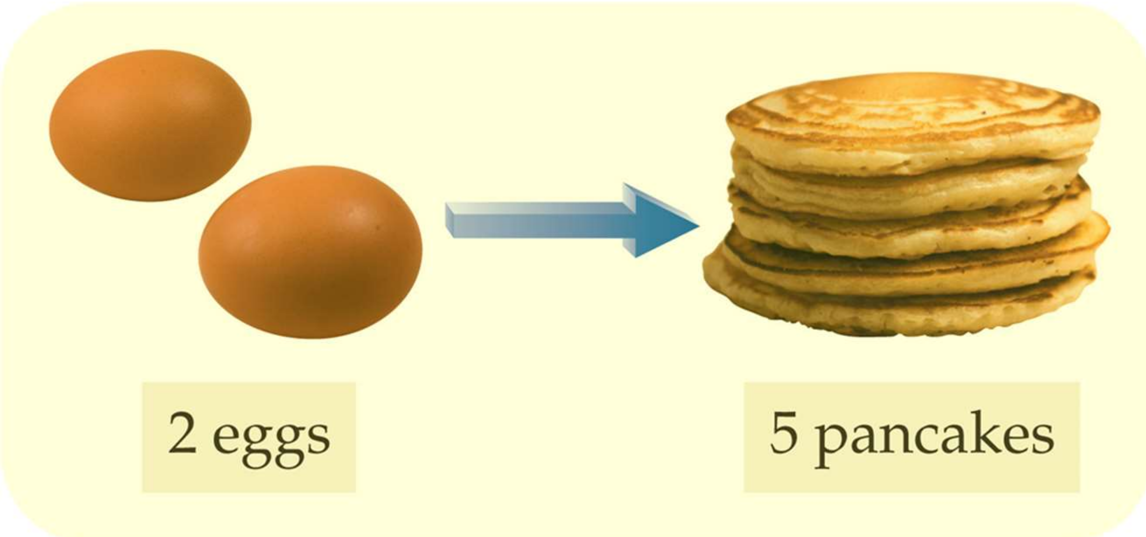
The recipe shows the numerical relationships between the pancake ingredients.

If we have 2 eggs—and enough of everything else—we can make 5 pancakes.

We can write this relationship as an equality or

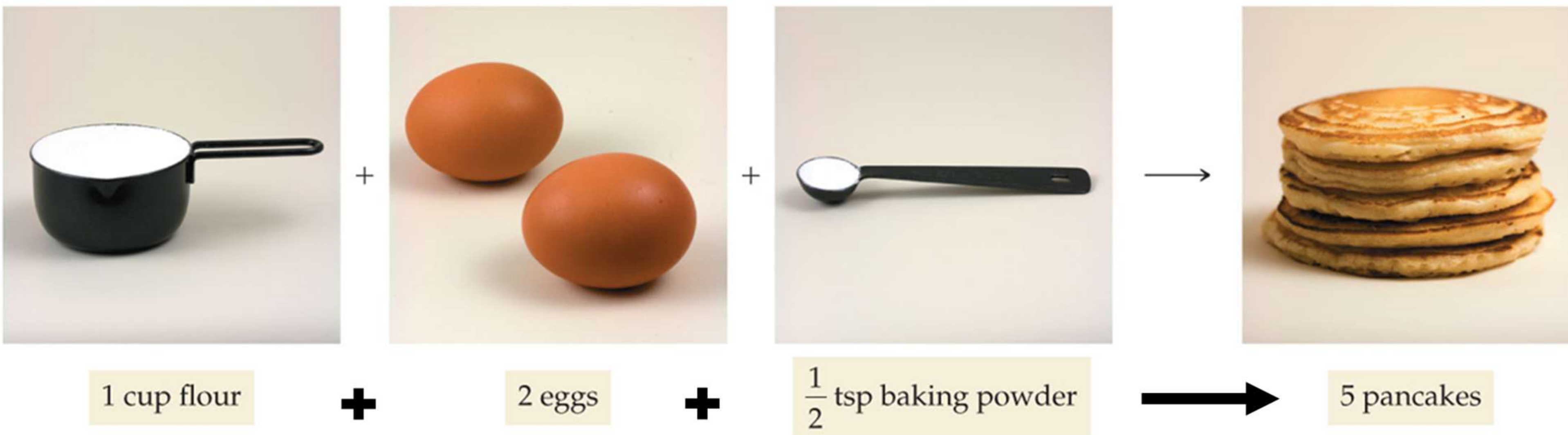
2 eggs=5 pancakes

What if we have 8 eggs? Assuming that we have enough of everything else, how many pancakes can we make?



Knowing this relationship we can use it to figure out how many pancakes we expect.

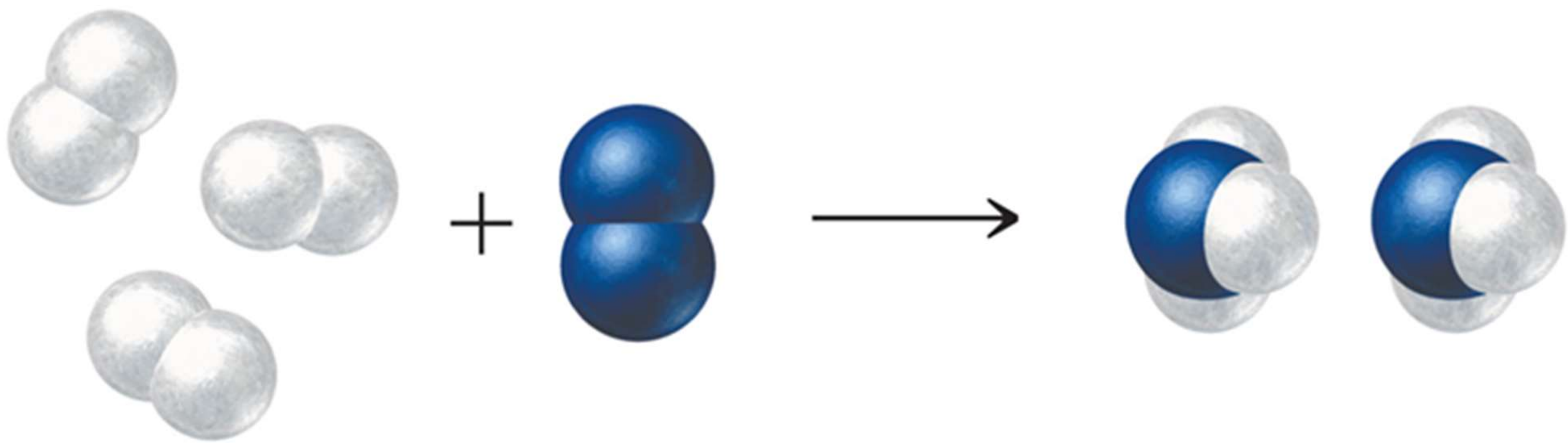
$$8\ \cancel{\text{eggs}} \times \frac{5\ \text{pancakes}}{2\ \cancel{\text{eggs}}} = 20\ \text{pancakes}$$



Just as we make pancakes in a balanced chemical equation, we have a "recipe" for how to combine ingredients to form products.

$$3\ H_{2(g)} + N_{2(g)} \rightarrow 2\ NH_{3(g)}$$

This is chapters concept is known as stoichiometry (stoy·kee·aa·muh·tree) which origin come from the Greek (stoich and English (metry) language and literally translates to the measure of elements



$$3\ H_{2(g)} + N_{2(g)} \rightarrow 2\ NH_{3(g)}$$

3 H_2 molecules : 1 N_2 molecule : 2 NH_3 molecules

Since we do not ordinarily deal with individual molecules, we can express the same ratios in moles.

$$3 \text{ mol } H_2 : 1 \text{ mol } N_2 : 2 \text{ mol } NH_3$$

If we have 3 mol of N_2, and more than enough H_2, how much NH_3 can we make?

$$\text{mol } N_2 \longrightarrow \text{mol } NH_3$$

$$\frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2}$$

A **mole-mole factor** is a ratio of the moles for any two substances in an equation.

$$3 \cancel{\text{mol } N_2} \times \frac{2 \text{ mol } NH_3}{1 \cancel{\text{mol } N_2}} = 6 \text{ mol } NH_3$$

$$1 \text{ cup flour} + 2 \text{ eggs} + \tfrac{1}{2} \text{ tsp baking powder} \longrightarrow 5 \text{ pancakes}$$

Suppose we have 3 cups flour, 10 eggs, and 4 tsp baking powder.

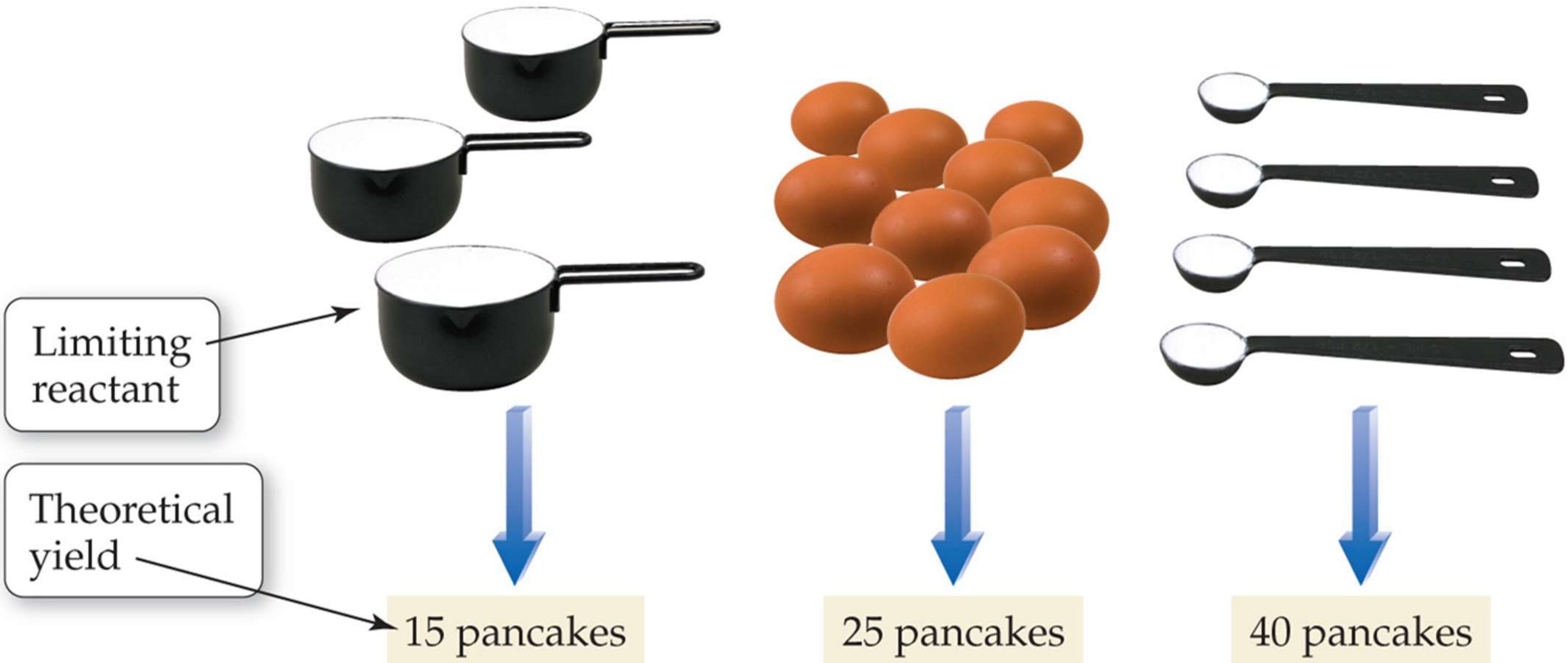
How many pancakes can we make?

$$3 \cancel{\text{cups flour}} \times \frac{5 \text{ pancakes}}{1 \cancel{\text{cup flour}}} = 15 \text{ pancakes} \qquad 10 \cancel{\text{eggs}} \times \frac{5 \text{ pancakes}}{2 \cancel{\text{eggs}}} = 25 \text{ pancakes}$$

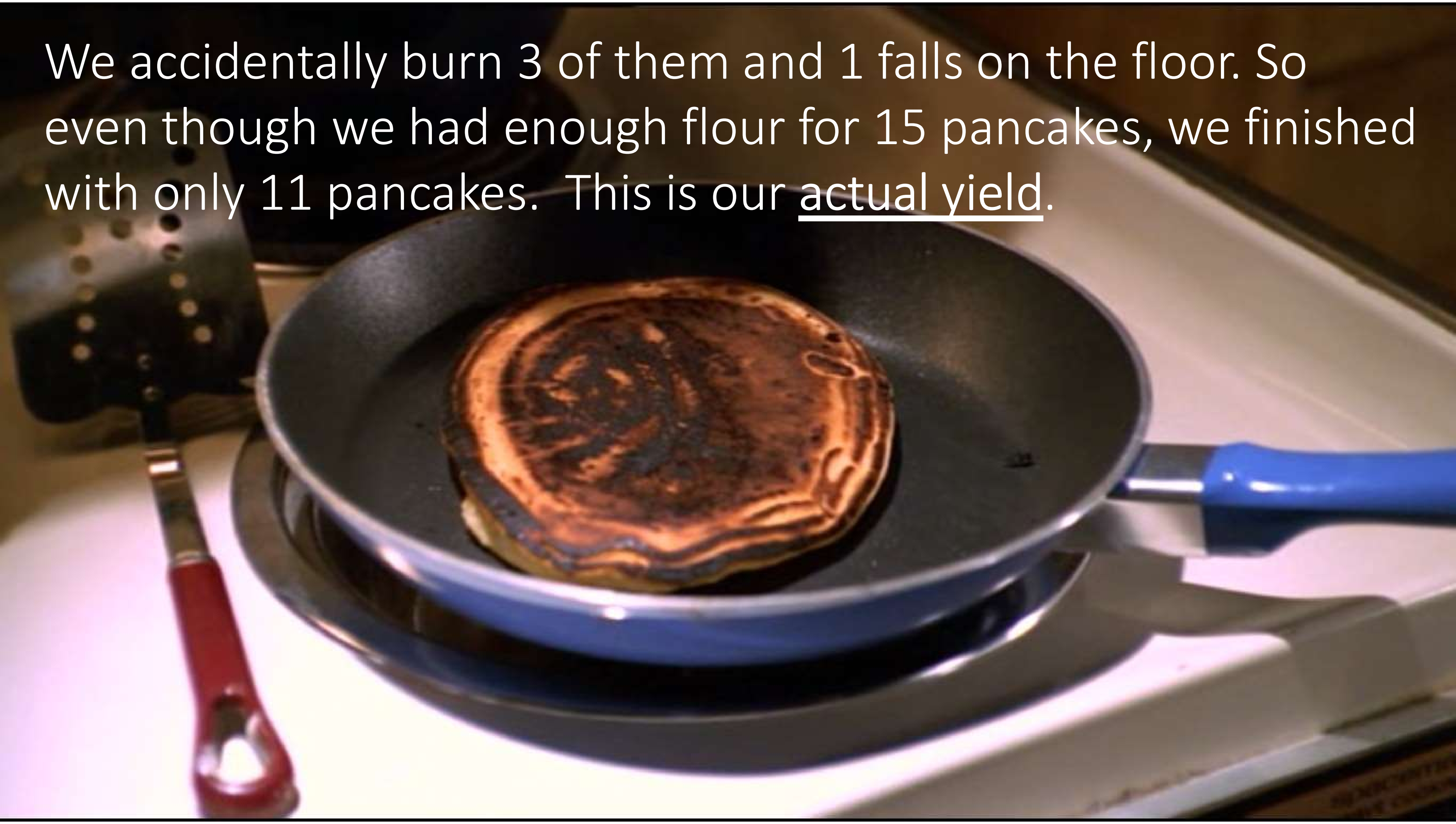
$$4 \cancel{\text{tsp baking powder}} \times \frac{5 \text{ pancakes}}{\frac{1}{2} \cancel{\text{tsp baking powder}}} = 40 \text{ pancakes}$$

We have enough flour for 15 pancakes, enough eggs for 25 pancakes, and enough baking powder for 40 pancakes.

If this were a chemical reaction, the flour would be the limiting reactant and 15 pancakes would be the theoretical yield (ideal amount).



We accidentally burn 3 of them and 1 falls on the floor. So even though we had enough flour for 15 pancakes, we finished with only 11 pancakes. This is our actual yield.



Our **percent yield**, the percentage of the theoretical yield that was attained, is:

$$\text{Percent yield} = \frac{11\ \text{pancakes}}{15\ \text{pancakes}} \times 100\% = 73\%$$

Since 4 of the pancakes were ruined, we got only 73% of our theoretical yield.

To summarize:

Limiting reactant (or limiting reagent)—the reactant that is completely consumed in a chemical reaction

Theoretical yield—the amount of product that can be made in a chemical reaction based on the amount of limiting reactant

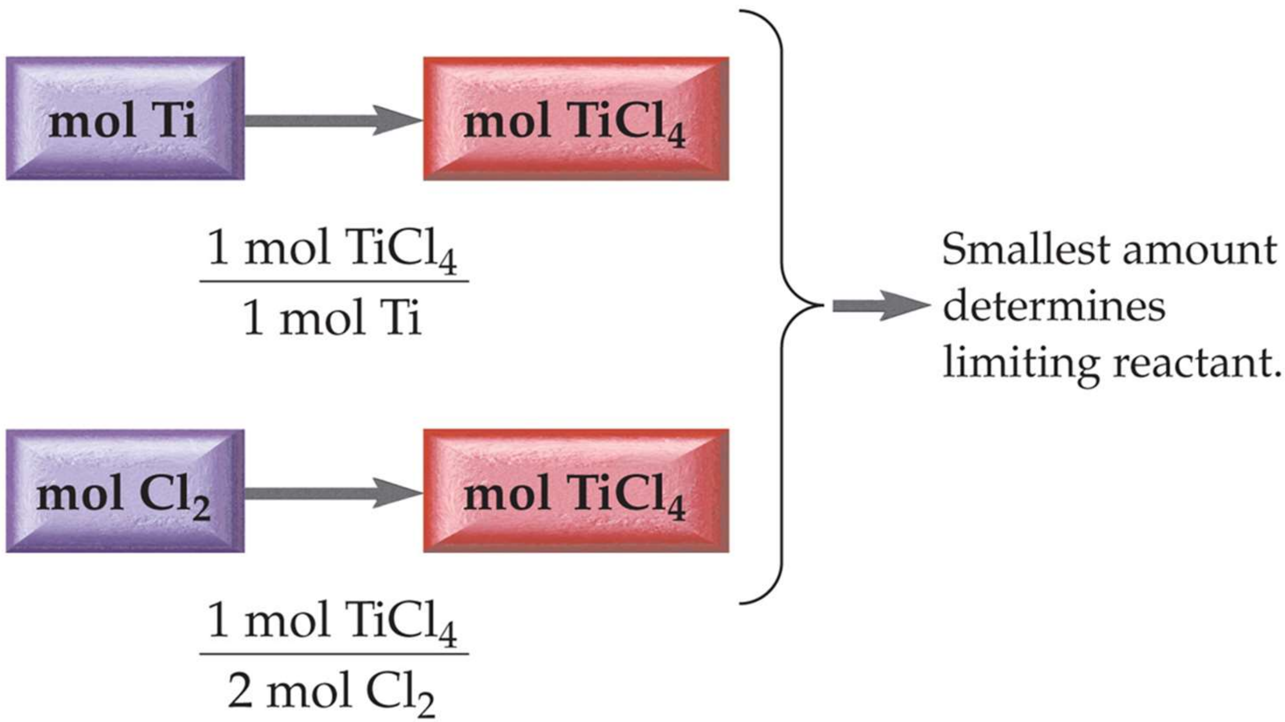
Actual yield—the amount of product produced by a chemical reaction.

Percent yield—(actual yield/theoretical yield) × 100%

$Ti(s) + 2\ Cl_2(g) \rightarrow TiCl_4(s)$

Given (***moles***): 1.8 mol Ti and 3.2 mol Cl_2

Find: limiting reactant and theoretical yield



Steps to solving limiting reactant problems

1. Convert g of each reactant to moles
2. Convert moles of each reactant to moles of the product being asked about.
3. That reactant that produces the least amount of product is the limiting reactant, the other is the excess reactant.
4. Use the least amount of moles attained and convert to mass of product. This is your theoretical yield.
5. If the actual yield is also given in the problem, or an actual yield is obtained in a lab, you can calculate the percent yield.

$Ti(s) + 2\ Cl_2(g) \rightarrow TiCl_4(s)$

Given (***moles***): 1.8 mol Ti and 3.2 mol Cl_2

Find: limiting reactant and theoretical yield

$$1.8\ \cancel{\text{mol Ti}} \times \frac{1\ \text{mol TiCl}_4}{1\ \cancel{\text{mol Ti}}} = 1.8\ \text{mol TiCl}_4$$

$$3.2\ \cancel{\text{mol}\ \boxed{\text{Cl}_2}} \times \frac{1\ \text{mol TiCl}_4}{2\ \cancel{\text{mol Cl}_2}} = \boxed{1.6\ \text{mol TiCl}_4}$$

Limiting reactant

Least amount of product

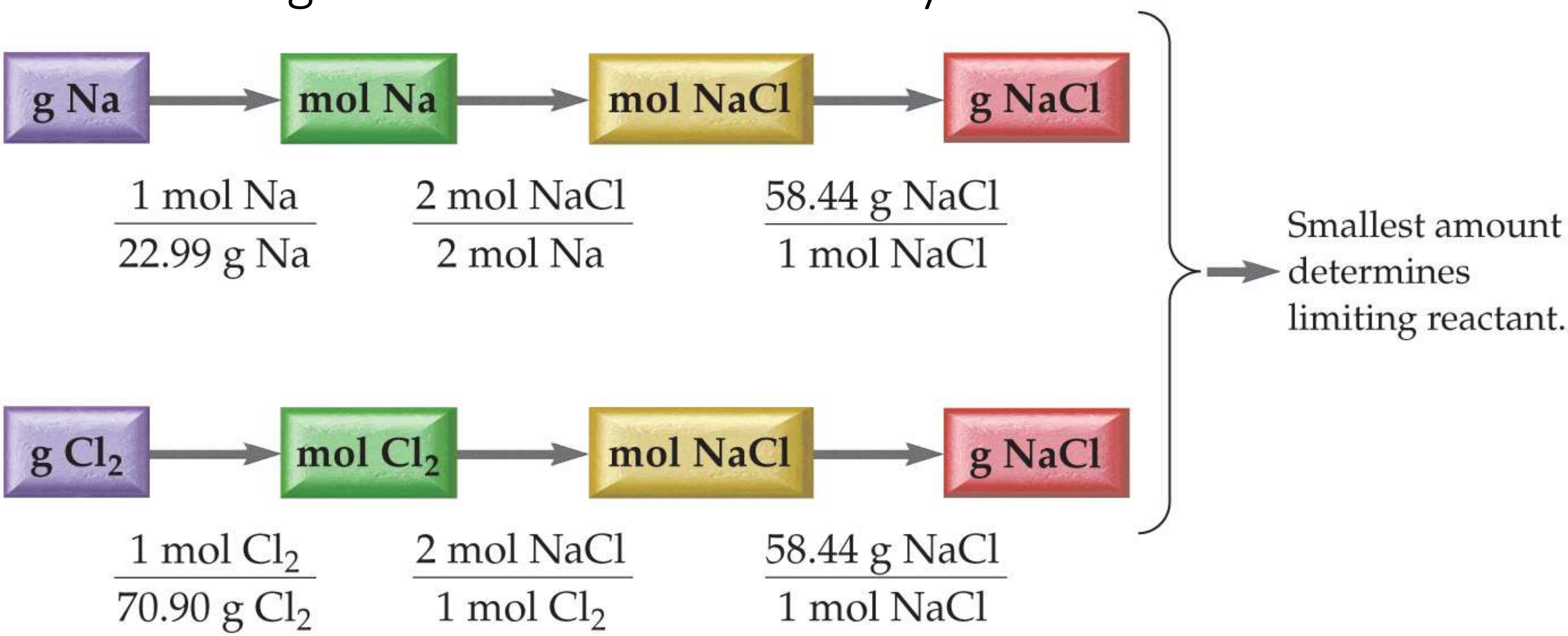
Must report in grams so we can measure out.

Limiting Reactant and Percent Yield: Gram to Gram

Example: $Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

Given (*grams*): 53.2 g Na and 65.8 g Cl_2

Find: limiting reactant and theoretical yield



Limiting Reactant and Percent Yield: Gram to Gram

Example: Na(s) + Cl_2(g) → 2 NaCl(s)
Given (*grams*): 53.2 g Na and 65.8 g Cl_2
Find: limiting reactant and theoretical yield

$$53.2\ \text{g}\ \cancel{\text{Na}} \times \frac{1\ \cancel{\text{mol Na}}}{22.99\ \cancel{\text{g Na}}} \times \frac{2\ \cancel{\text{mol NaCl}}}{2\ \cancel{\text{mol Na}}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = 135\ \text{g NaCl}$$

$$65.8\ \cancel{\text{g}\ \text{Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{2\ \cancel{\text{mol NaCl}}}{1\ \cancel{\text{mol Cl}_2}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = \boxed{108\ \text{g NaCl}}$$

Limiting reactant

Least amount of product

Example: $Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

Given (*grams*): actual yield 86.4 g NaCl

Find: percent yield

The actual yield is usually less than the theoretical yield because at least a small amount of product is lost or does not form during a reaction.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \frac{86.4\ \text{g}}{108\ \text{g}} \times 100\% = 80.0\%$$

Questions?