CHEMISTRY

Chapter 1 ESSENTIAL IDEAS

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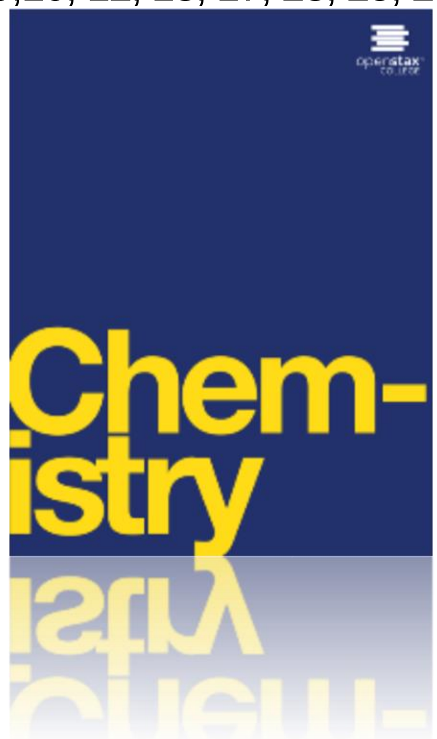
The Cooper Union

HW problems: 1, 2, 8, 9,10, 12, 13, 17, 23, 25, 27, 40, 45, 49, 51, 53, 77



CH. 1 OUTLINE

1.1: Chemistry in Context

1.2: Phases and Classification of Matter

1.3: Physical and Chemical Properties

1.4: Measurements

1.5: Measurement Uncertainty, Accuracy, and Precision

1.6: Mathematical Treatment of Measurement Results

1.1 CHEMISTRY IN CONTEXT

- **Chemistry is the study of the composition, properties, and interactions of matter.**
- **Since the Universe is composed of matter (anything that has mass and takes up space) and energy, and when matter undergoes change it involves energy, Chemistry is the study of the Universe.**
- **Attempts to understand the behavior of matter extend back more than 2500 years.**
  - Greeks: Matter consists of four elements; earth, air, fire, and water. This view persisted for hundreds of years.
  - Alchemists attempted to transform "base metals" into "noble metals". Doomed to fail (that would require a nuclear reaction), they discovered the foundations of modern chemistry in the process.

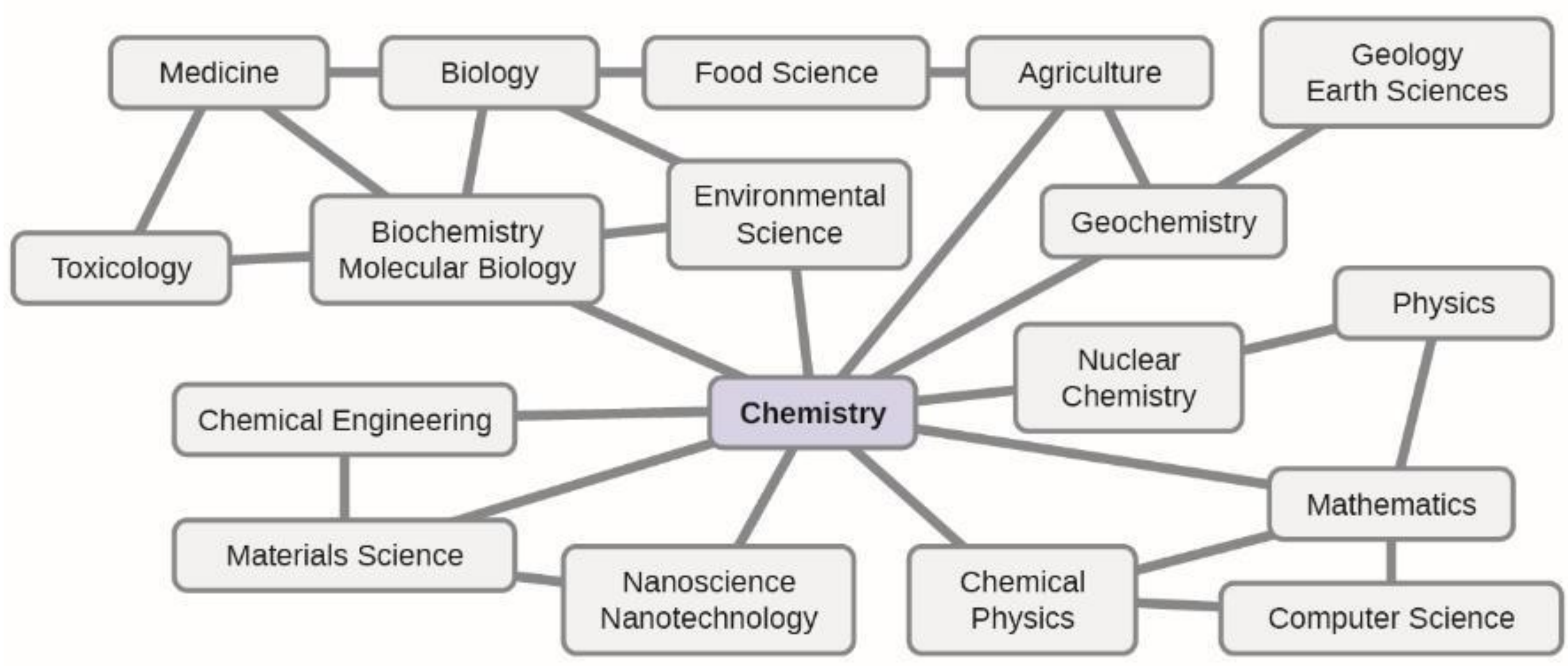
THE BIG SECRET

Treat chemistry as a second language.

IF YOU EVER DON'T UNDERSTAND A WORD YOU READ OR A WORD I SAY IN CLASS, LOOK IT UP OR ASK FOR CLARIFICATION.

CHEMISTRY THE CENTRAL SCIENCE

- Chemistry is interconnected to a vast array of other STEM disciplines.



CHEMISTRY AND EVERYDAY LIFE

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- **Examples of chemistry in everyday life:**
  - Digesting food
  - Synthesizing polymers for clothing, cookware, and credit cards
  - Refining crude oil into gasoline and other products
  - And much more…

- **As you proceed through this course, you will discover:**
  - Many different examples of changes in the composition and structure of matter.
  - How to classify these changes in matter and understand how they occur.
  - The changes in energy that accompany these changes in matter.

THE SCIENTIFIC METHOD

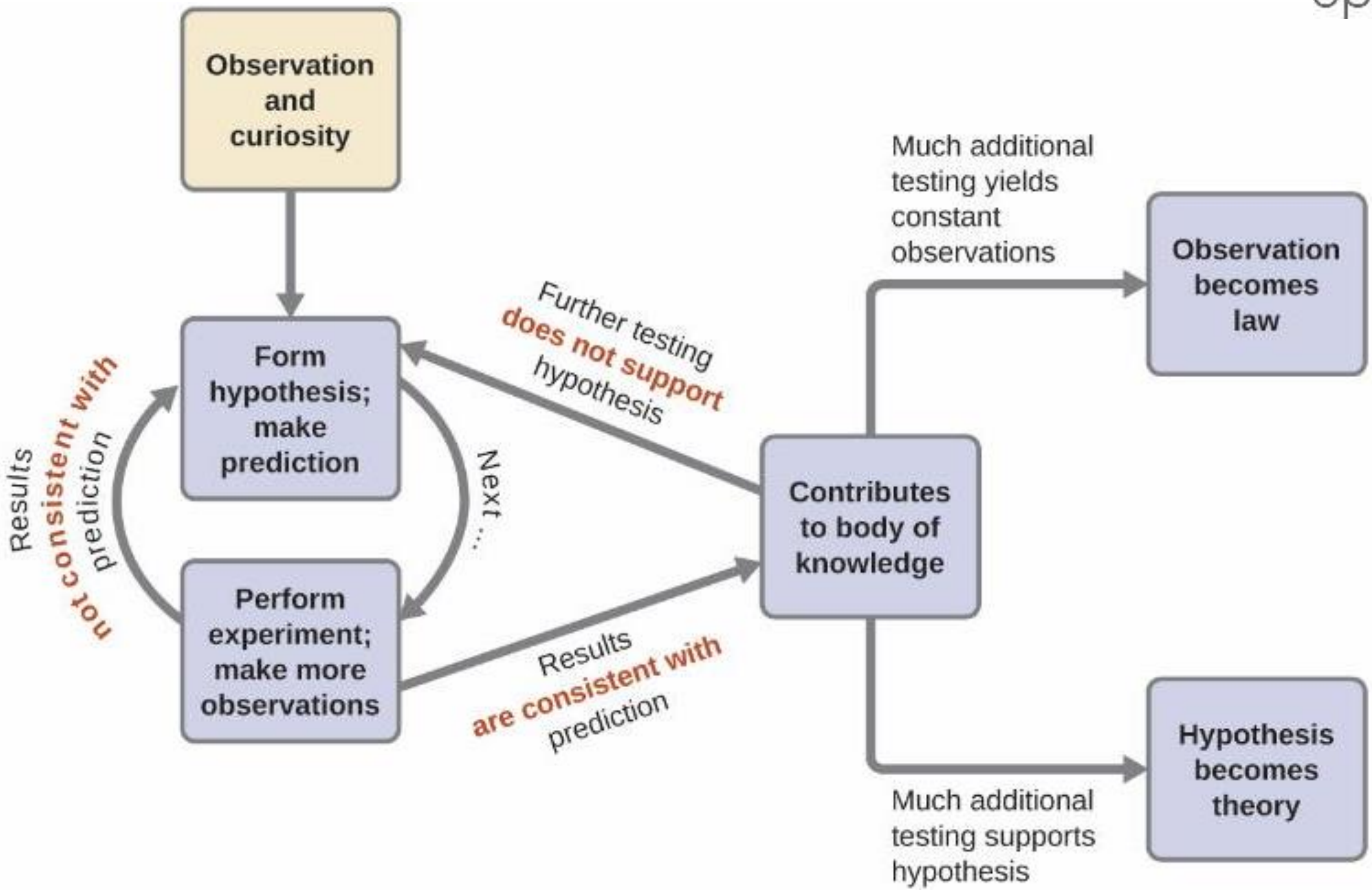
- Chemistry is a science based on observation and experimentation.

- Chemists often formulate a **hypothesis – a tentative explanation of observations**. **(An educated guess.)**

- The **laws** of science summarize a vast number of experimental observations, and describe or predict some facet of the natural world.

- **Theory -** A well-substantiated, comprehensive, testable explanation of a particular aspect of nature.

THE SCIENTIFIC METHOD



The scientific method follows a process similar to the one shown in this diagram. All the key components are shown, in roughly the right order. Scientific progress is seldom neat and clean: It requires open inquiry and the reworking of questions and ideas in response to findings. It involves significant creativity in making hypotheses and designing experiments. And you can still simply get lucky!

The loop is the key to the method… if it doesn't pass the test, revise your guess!

THE DOMAINS OF CHEMISTRY

Chemists study and describe the behavior of matter and energy in three different domains.

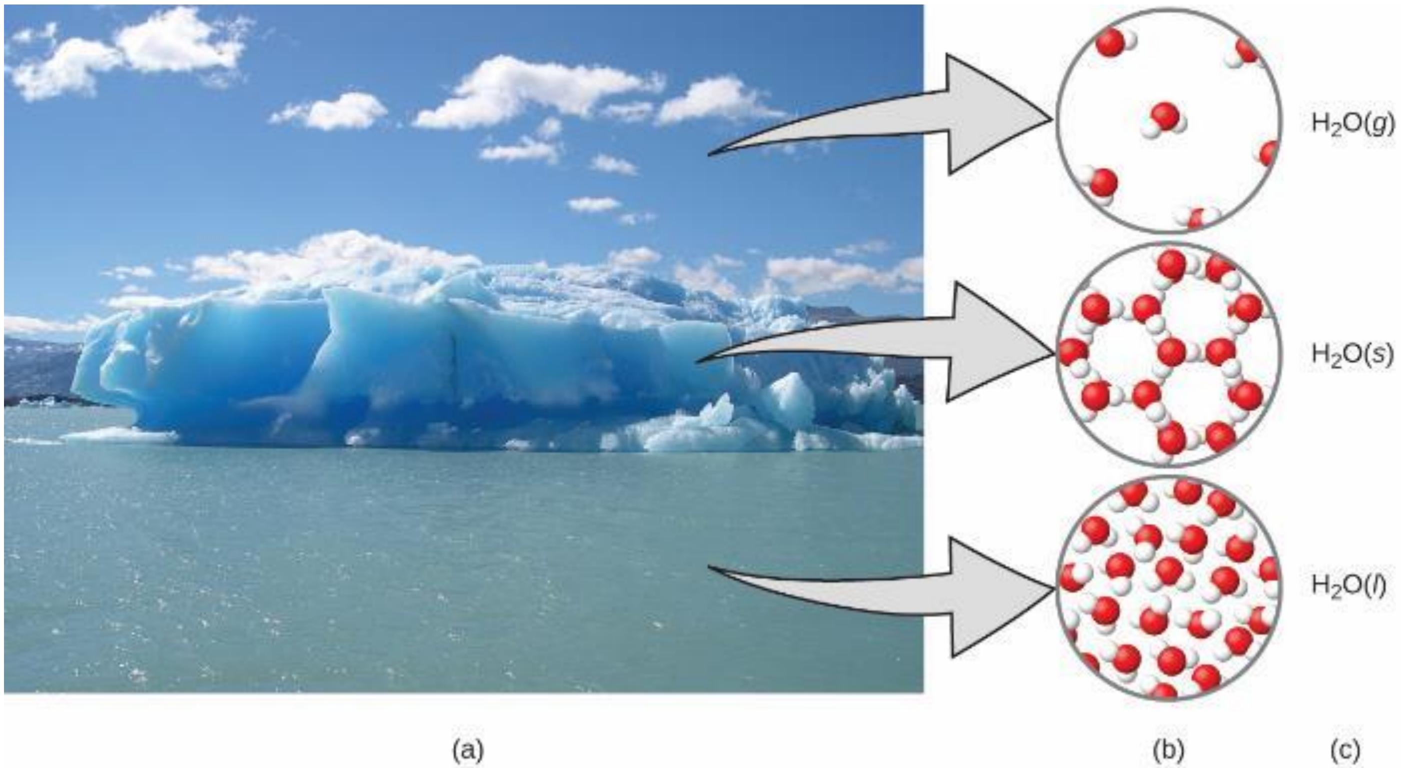
1) The **macroscopic domain** is familiar to us: It is the realm of everyday things that are large enough to be sensed directly by human sight or touch.

2) The **microscopic domain** of chemistry is almost always visited in the imagination. *Micro* also comes from Greek and means “small.” Some aspects of the microscopic domains are visible through a microscope.

3) The **symbolic domain** contains the specialized language used to represent components of the macroscopic and microscopic domains, such as chemical symbols.

(minor point, but the fact that the study of chemistry uses macroscopic representations of microscopic topics can be challenging for beginners)

THE DOMAINS OF CHEMISTRY



(a) Moisture in the air, icebergs, and the ocean represent water in the macroscopic domain.

(b) At the molecular level (microscopic domain), gas molecules are far apart and disorganized, solid water molecules are close together and organized, and liquid molecules are close together and disorganized.

(c) The formula H_2O symbolizes water, and (*g*), (*s*), and (*l*) symbolize its phases. Note that clouds are actually comprised of either very small liquid water droplets or solid water crystals; gaseous water in our atmosphere is not visible to the naked eye, although it may be sensed as humidity. (credit a: modification of work by “Gorkaazk”/Wikimedia Commons)

1.2 PHASES AND CLASSIFICATION OF MATTER

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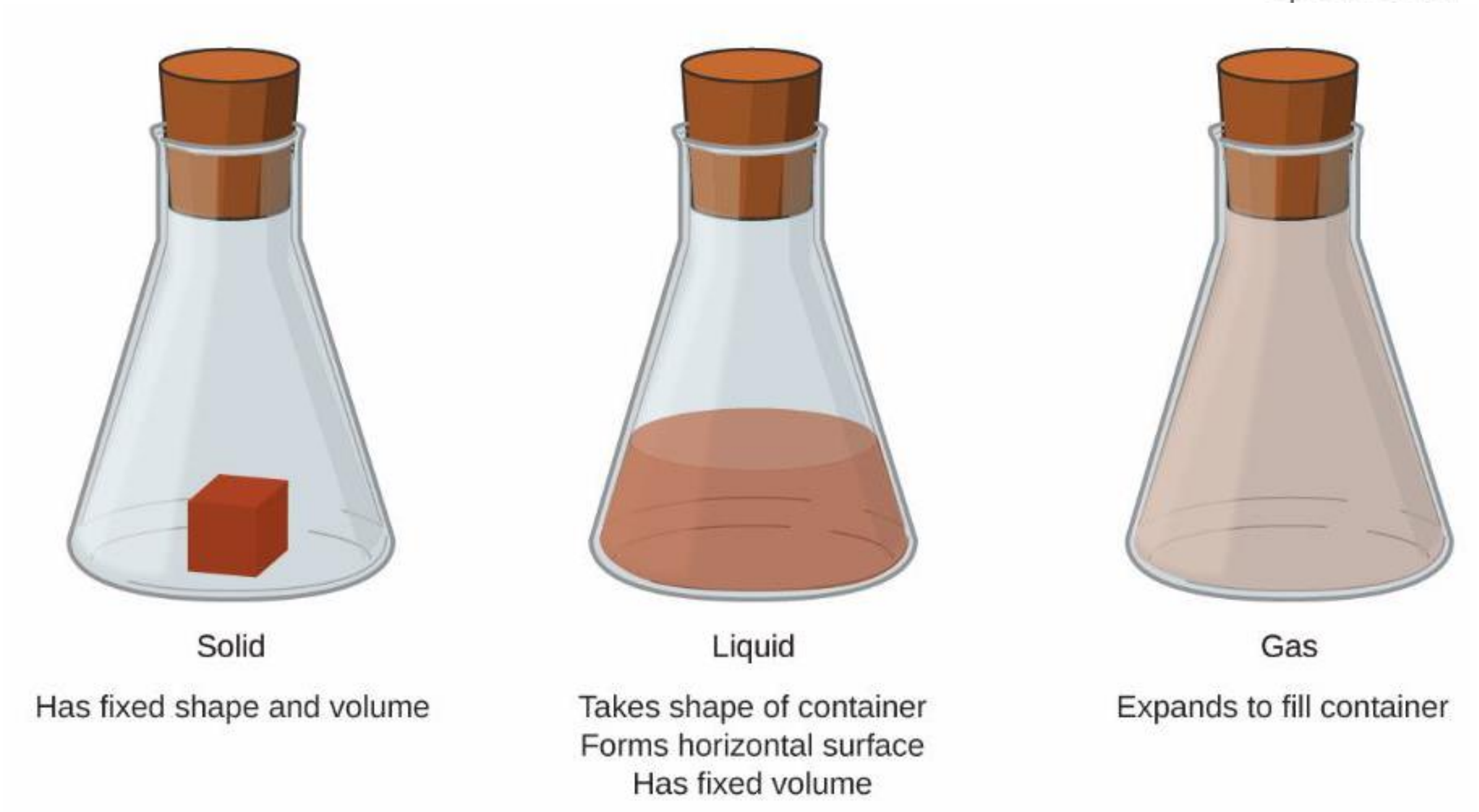
Matter - **Anything that occupies space and has mass.**

The three most common states or phases of matter:

1) A solid is rigid and posseses a definite shape.

2) A liquid flows and takes the shape of its container.

3) A gas takes both the shape and volume of its container.



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

PLASMA: A FOURTH STATE OF MATTER

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

- **Plasma has unique properties distinct from ordinary gases.**

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

- (There are other states of matter- liquid crystal, Bose-Einstein condensate…)

PLASMA: A FOURTH STATE OF MATTER

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

MASS VS. WEIGHT

- **Mass is a measure of the amount of matter in an object.**

- **Weight refers to the force that gravity exerts on an object.**

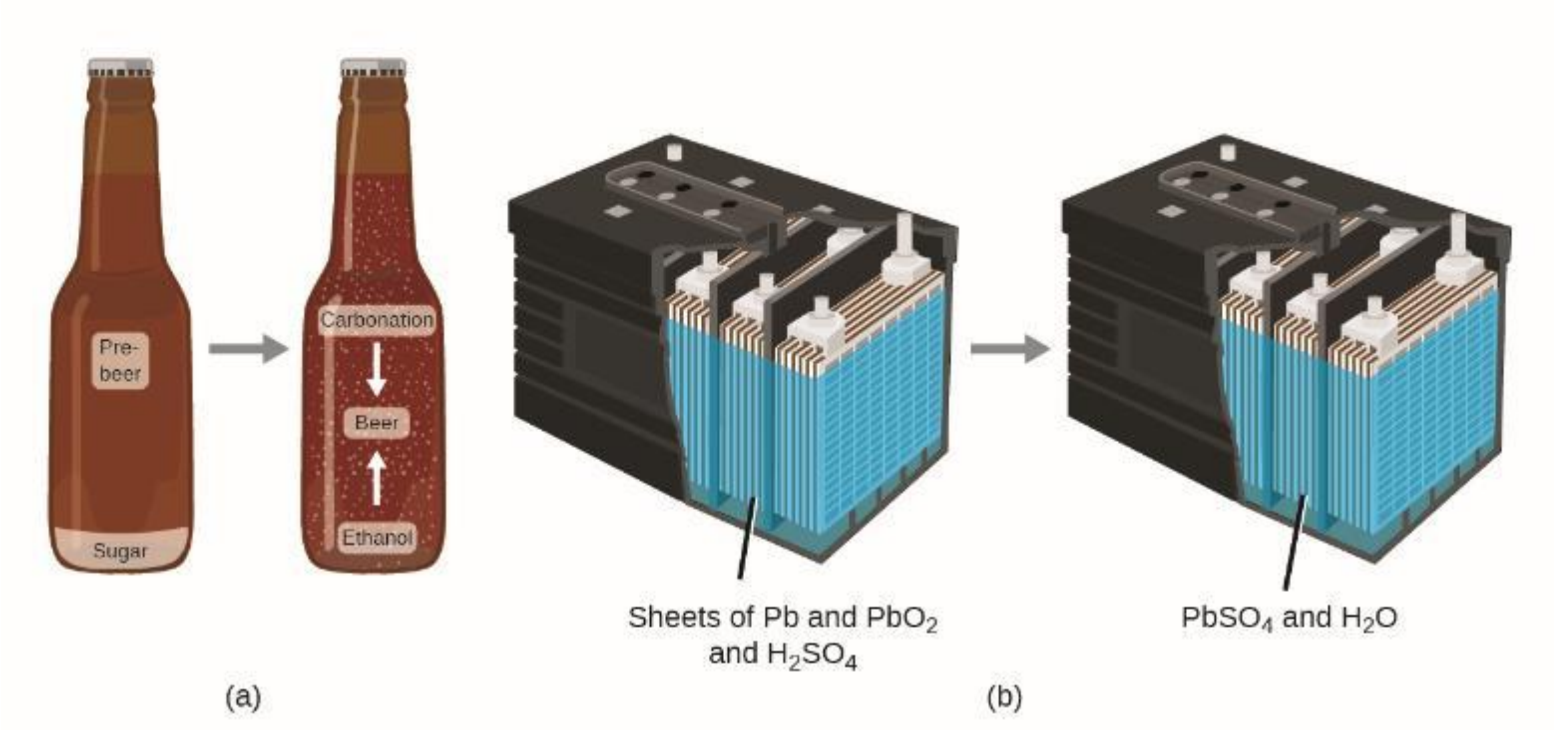
- **An object’s mass is the same on the earth and the moon but its weight is different.**

LAW OF CONSERVATION OF MATTER

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

- **This is true for both chemical and physical changes.**

LAW OF CONSERVATION OF MATTER



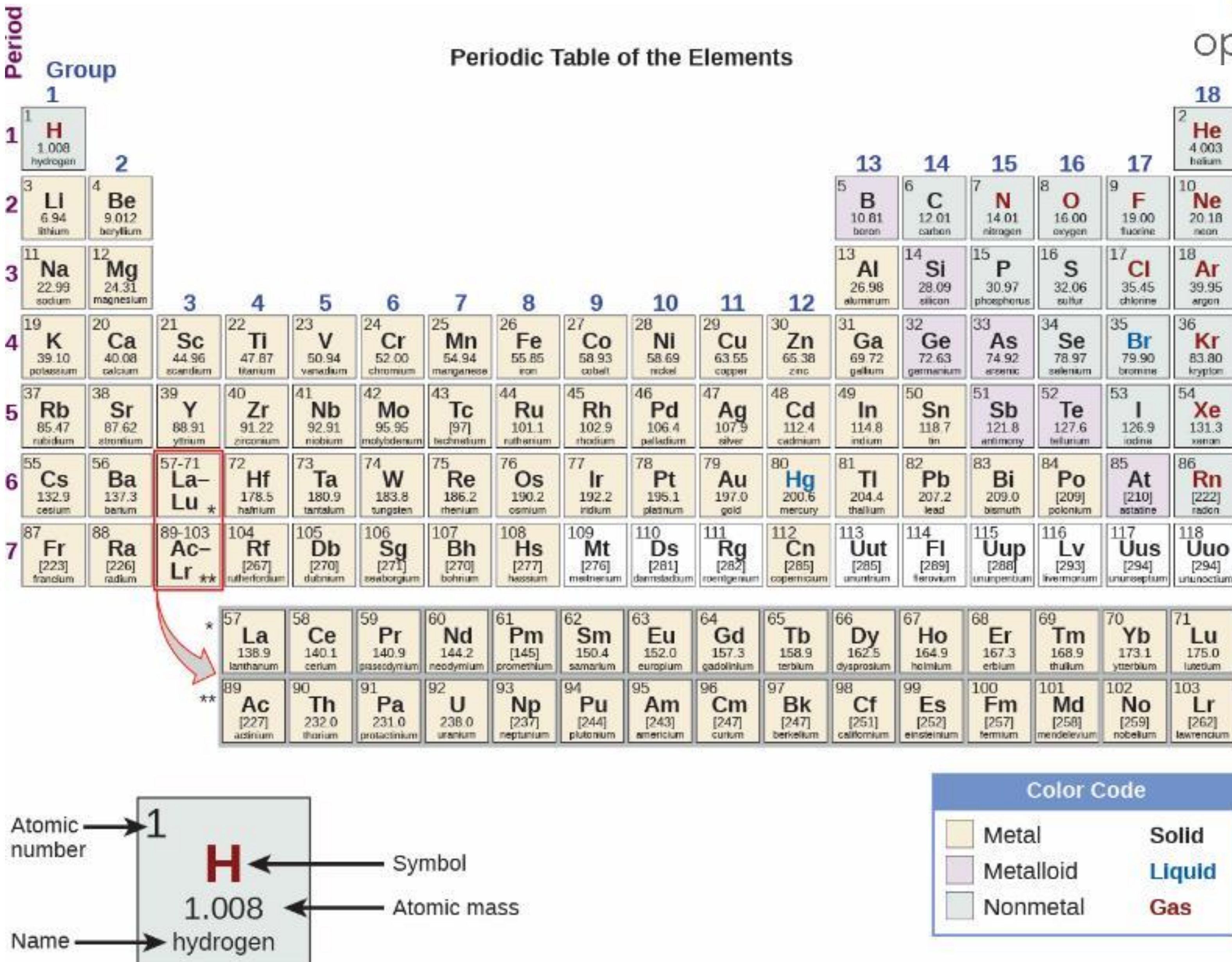
(a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbonation.

(b) The mass of the lead, lead oxide plates, and sulfuric acid that goes into the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

ELEMENTS

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

- **The known elements are displayed in the periodic table.**
  - There are more than 100 known elements
  - 90 of these occur naturally
  - Two dozen or so have been created in laboratories



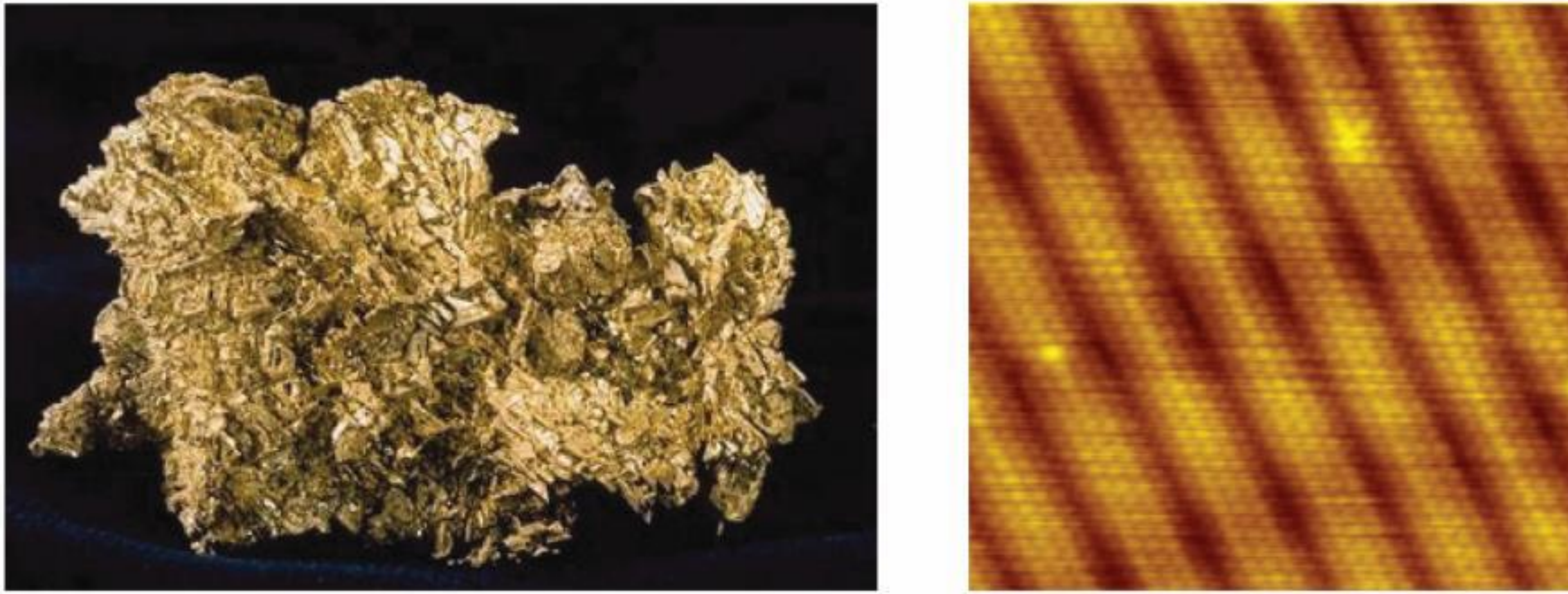
Note: all elements up to 118 have been named.

ATOMS AND MOLECULES

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

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

GOLD ATOMS



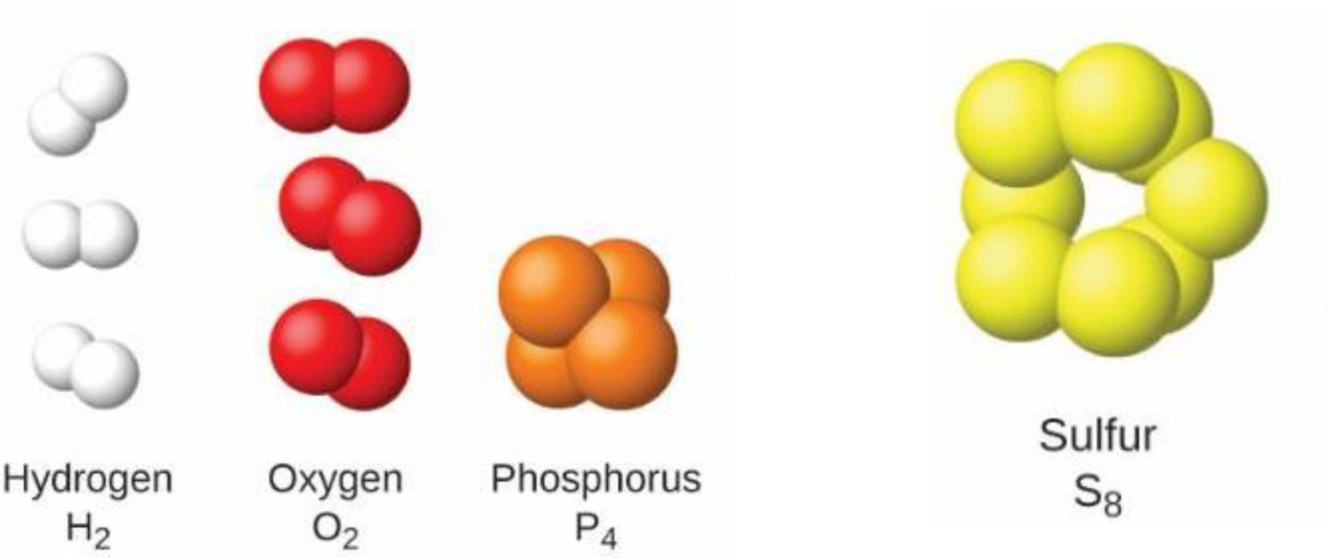
(a) (b)

(a) This photograph shows a gold nugget.

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

MOLECULES

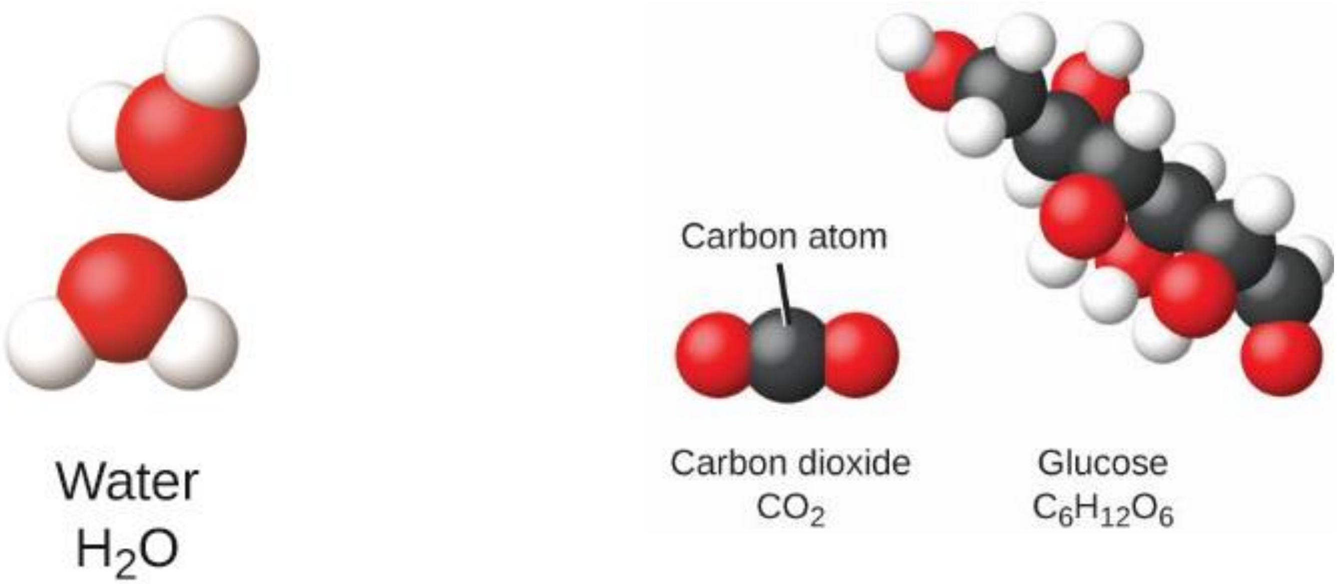
- Only a few elements exist as individual atoms.
- Most elements exist as molecules where two or more atoms of the same element are bonded together.



The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element.

MOLECULES

- Many molecules consist of two or more atoms of different elements.
- Atoms in all types of molecules move as a unit.



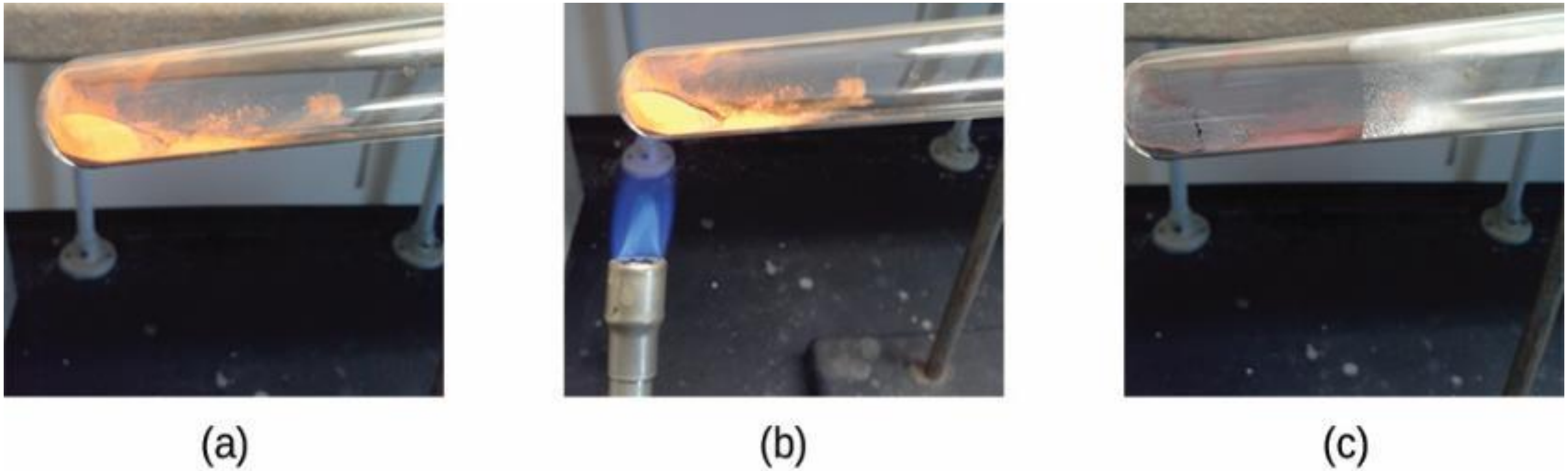
The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

PURE SUBSTANCES AND MIXTURES

- **Pure substances have constant composition.**
  - **Elements – ** Pure substance that **CANNOT** be broken down into simpler substances by chemical changes.
    - Consist of one type of element.
    - Examples: Gold (Au), Phosphorus (P_4), Oxygen (O_2)
  - **Compounds –** Pure substances that **CAN** be broken down into simpler substances by chemical changes.
    - Consist of two or more types of elements chemically bonded.
    - Examples: H_2O, $C_6H_{12}O_6$, AgCl
    - The properties of compounds are different from the uncombined elements making up the compound.

BREAKING DOWN A COMPOUND

- Upon heating, the compound, mercury (II) oxide, is broken down into its elements, mercury and oxygen.



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

PURE SUBSTANCES AND MIXTURES

- **A mixture is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes.**

- Evaporation is an example of a physical change.

- There are two types of mixtures: homogenous mixtures and heterogeneous mixtures.

TWO TYPE OF MIXTURES

1) A **homogenous mixture** exhibits a uniform composition and appears visually the **same throughout.**

Another name for a homogenous mixture is a ***solution.***

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

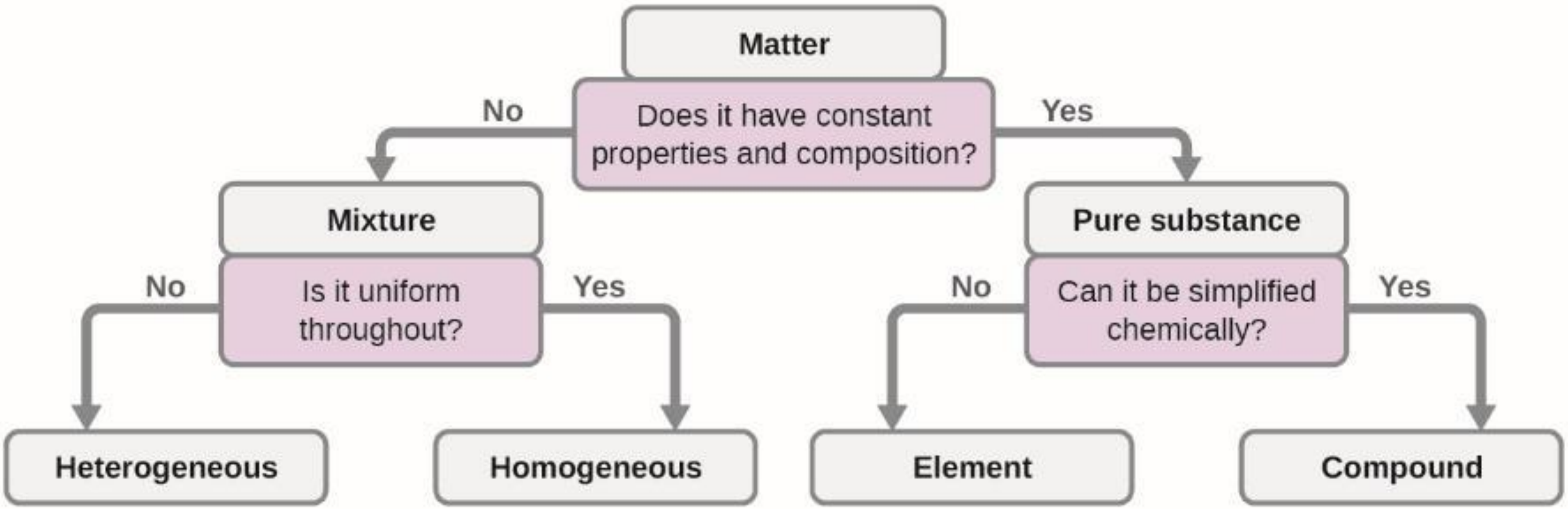
TWO TYPE OF MIXTURES

(a) (b)

(a) Oil and vinegar salad dressing is a heterogeneous mixture because its composition is not uniform throughout.

(b) A commercial sports drink is a homogeneous mixture because its composition is uniform throughout. (credit a “left”: modification of work by John Mayer; credit a “right”: modification of work by Umberto Salvagnin; creditb “left: modification of work by Jeff Bedford)

CLASSIFYING MATTER



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

1.3 PHYSICAL AND CHEMICAL PROPERTIES

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

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

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

PHYSICAL CHANGES

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

(a) Wax undergoes a physical change when solid wax is heated and forms liquid wax.

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

PHYSICAL AND CHEMICAL PROPERTIES

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

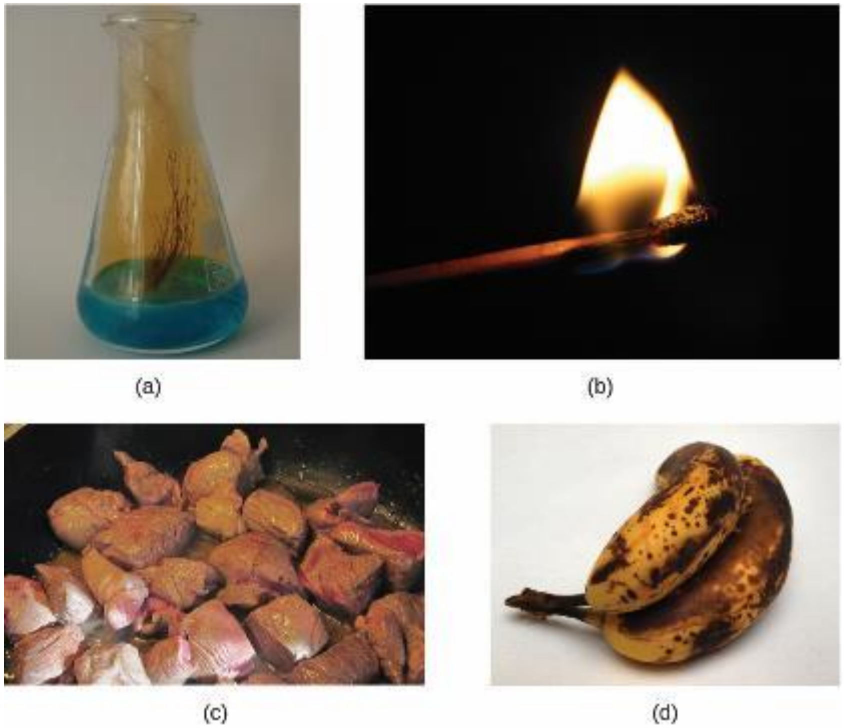
  - Examples: flammability, toxicity, acidity, reactivity, and heat of combustion.

CHEMICAL PROPERTIES

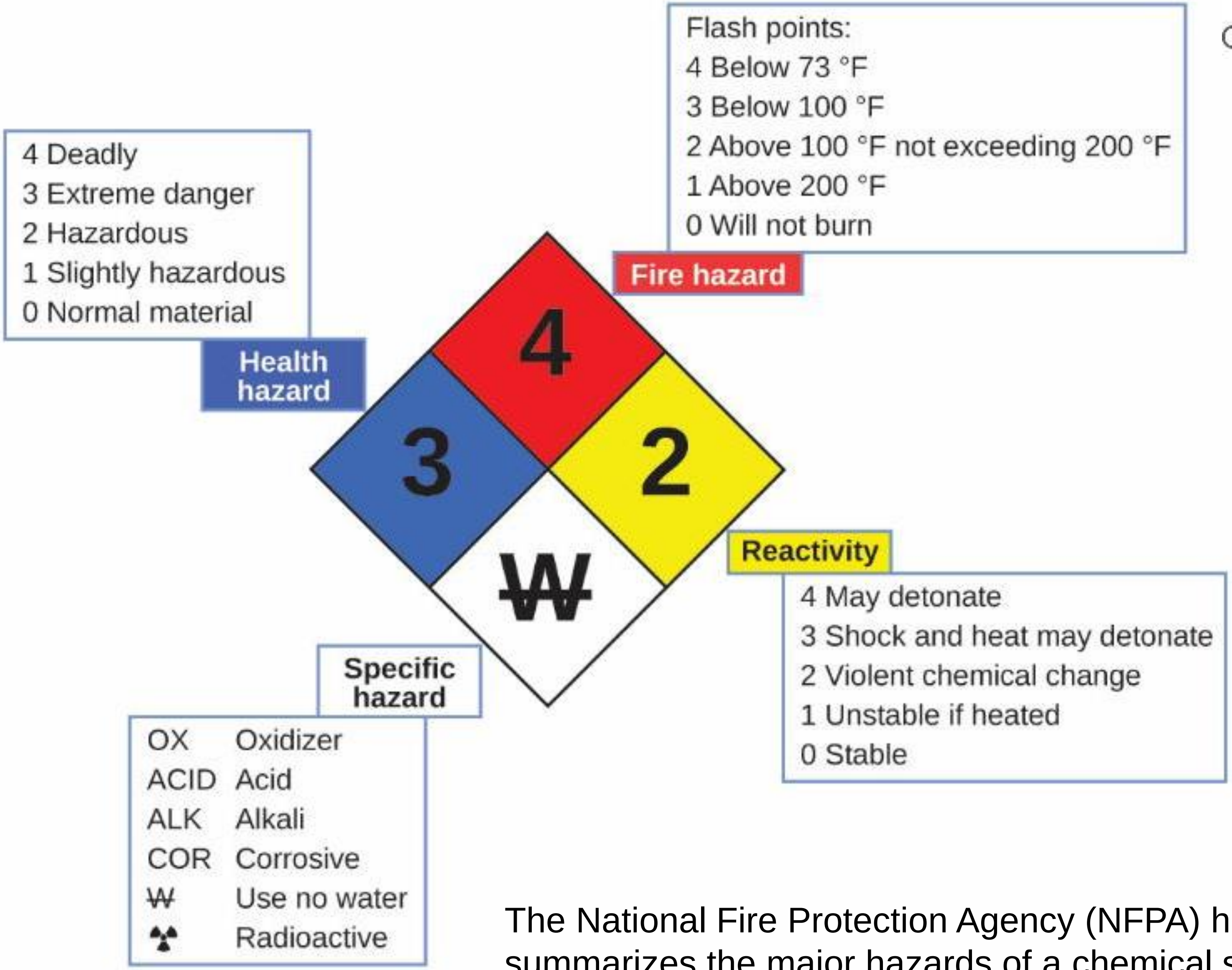
(a) (b)

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

CHEMICAL CHANGES



(a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo)



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

EXTENSIVE PROPERTIES AND INTENSIVE PROPERTIES

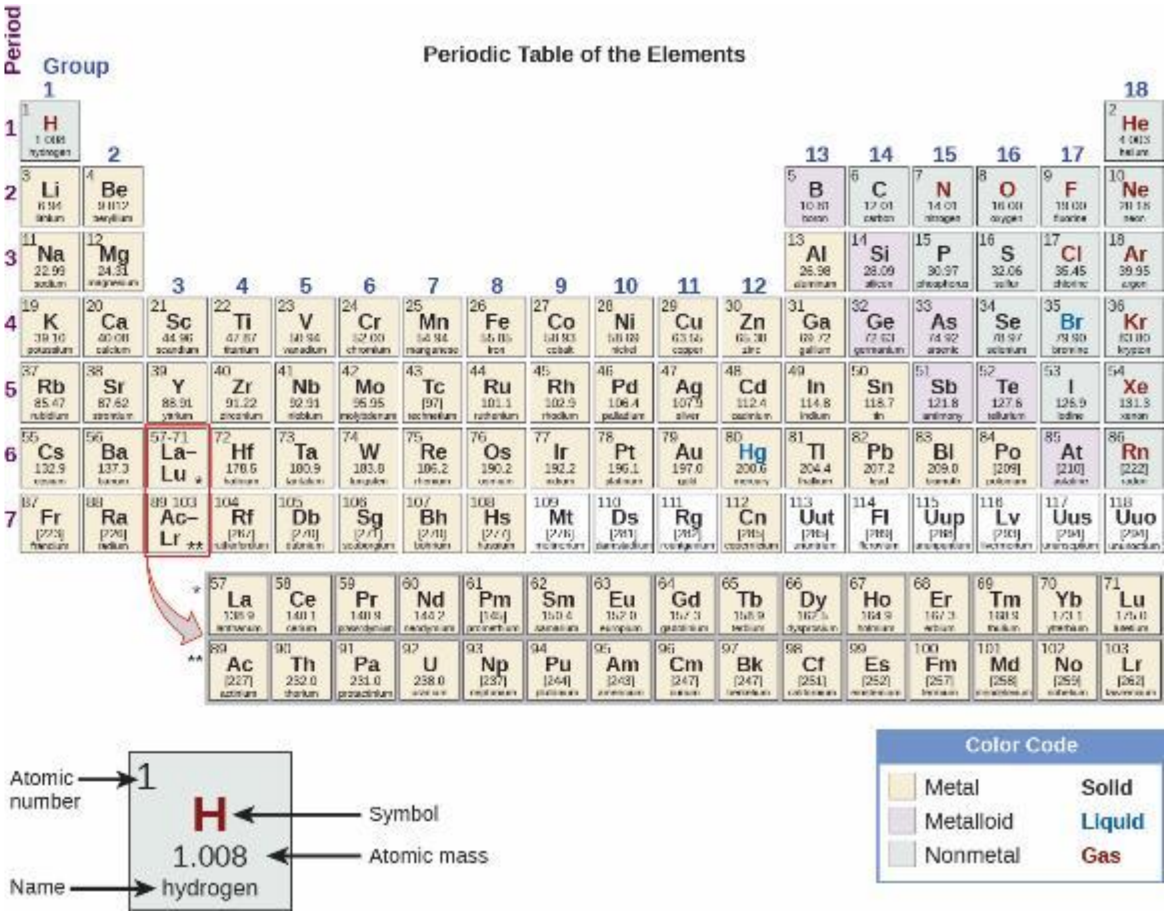
Extensive property

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

Intensive property

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

PERIODIC TABLE OF THE ELEMENTS



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

1.4 MEASUREMENTS

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• Measurements provide the information that is the basis of most of the hypotheses, theories, and laws in chemistry.

• Every Measurement provides three kinds of information:

1) The size or magnitude of the measurement - **A Number**

2) A standard of comparison for the measurement - **A Unit**

3) An indication of the uncertainty of the measurement.

UNITS

- **Without units a number can be meaningless or confusing!**

- **In chemistry we use an updated version of the metric system known as the International System of Units or SI Units.**
  - **Used since 1964**

SI BASE UNITS

Property	Name of Unit	Symbol of Unit
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous Intensity	candela	cd

SI UNIT PREFIXES

- Fractional or multiple SI units are named using a prefix and the name of the base unit.

Prefix	Symbol	Factor
femto	f	10^{-15}
pico	p	10^{-12}
nano	n	10^{-9}
micro	µ	10^{-6}
milli	m	10^{-3}
centi	c	10^{-2}
deci	d	10^{-1}

SI UNIT PREFIXES

- Fractional or multiple SI units are named using a prefix and the name of the base unit.

Prefix	Symbol	Factor
kilo	k	10^3
mega	M	10^6
giga	G	10^9
tera	T	10^{12}

Note: I will always give you a prefix table. No need to memorize. You'll end up knowing the ones you use often.

COMMON SI BASE UNITS: LENGTH

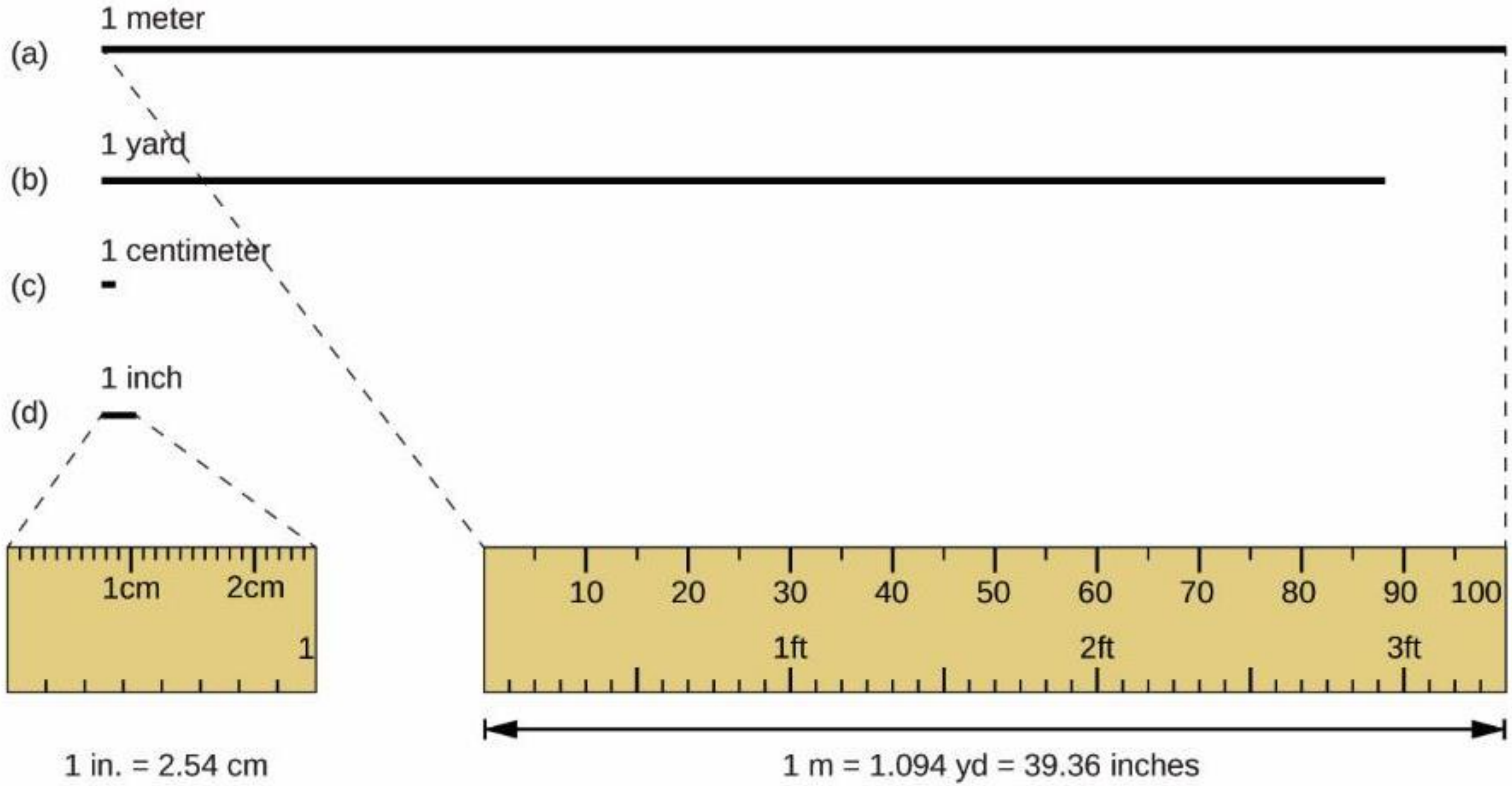
- **The SI unit of length is the meter (m).**

- The meter was originally intended to be 1/10,000,000 of the distance from the North Pole to the equator.

- A meter is now defined as the distance light travels in a vacuum in 1/299,792,458 of one second.

- A meter is about three inches longer than a yard.

RELATIVE LENGTHS



The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

COMMON SI BASE UNITS: MASS

- **The SI unit of mass is the kilogram (kg).**

- A kilogram was originally defined as the mass of a liter of water.

- It is now defined by a certain cylinder of platinum-iridium alloy, which is kept in France.

- One kilogram is about 2.2 pounds (lbs.)

KILOGRAM

This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology)

COMMON SI BASE UNITS: TEMPERATURE

- **The SI unit of temperature is the kelvin (K).**
- No degree word nor symbol (°) is used with kelvin.
- The degree Celsius (° C) is also allowed in the SI system.
- Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places.
- Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C).

COMMON SI BASE UNITS: TIME

- **The SI unit of time is the second (s).**

- Smaller and larger time intervals can be expressed with the appropriate prefixes.

- Alternatively, hours, days, and years can be used.

DERIVED SI UNITS: VOLUME AND DENSITY

- We can derive many units from the seven SI base units.

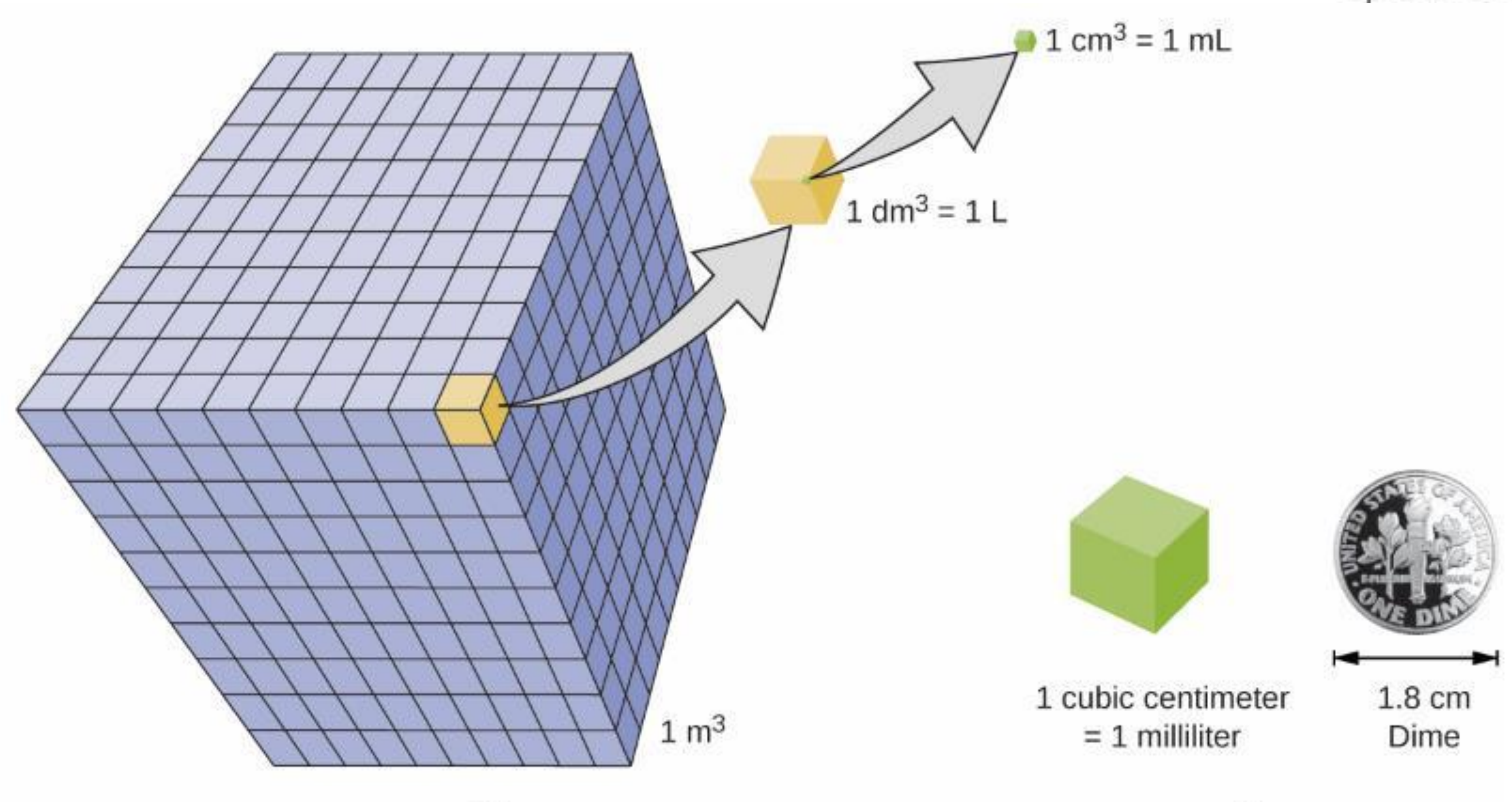
- **Volume –** The measure of the amount of space occupied by an object.

- **The standard SI unit for volume is the cubic meter (m^3),** which is derived from the SI base unit of length.

- Other units for volume are the *liter* (L) and *milliliter* (mL).

- 1 dm^3 = 1 L
- 1 cm^3 = 1 mL

VOLUME UNITS



(a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale).

(b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

DENSITY

- The ***density*** of a substance is the ratio of the mass of a sample of the substance to its volume.

$$density = \frac{mass}{volume}$$

- The standard SI unit for density is the kilogram per cubic meter (kg/m^3).

- Commonly used density units based on state of matter:

  g/cm^3 (solids, liquids)

  g/L (gases)

1.5 MEASUREMENT UNCERTAINTY, ACCURACY, AND PRECISION

- Counting is the only type of measurement that is free from uncertainty.

- The result of a counting measurement is an example of an ***exact number.***

- The numbers for defined quantities are also exact.

  - 1 foot is exactly 12 inches
  - 1 inch is exactly 2.54 cm
  - 1 gram is exactly 0.001 kg

1.5 MEASUREMENT UNCERTAINTY, ACCURACY, AND PRECISION

- Quantities derived from measurements other than counting are uncertain to varying extents.

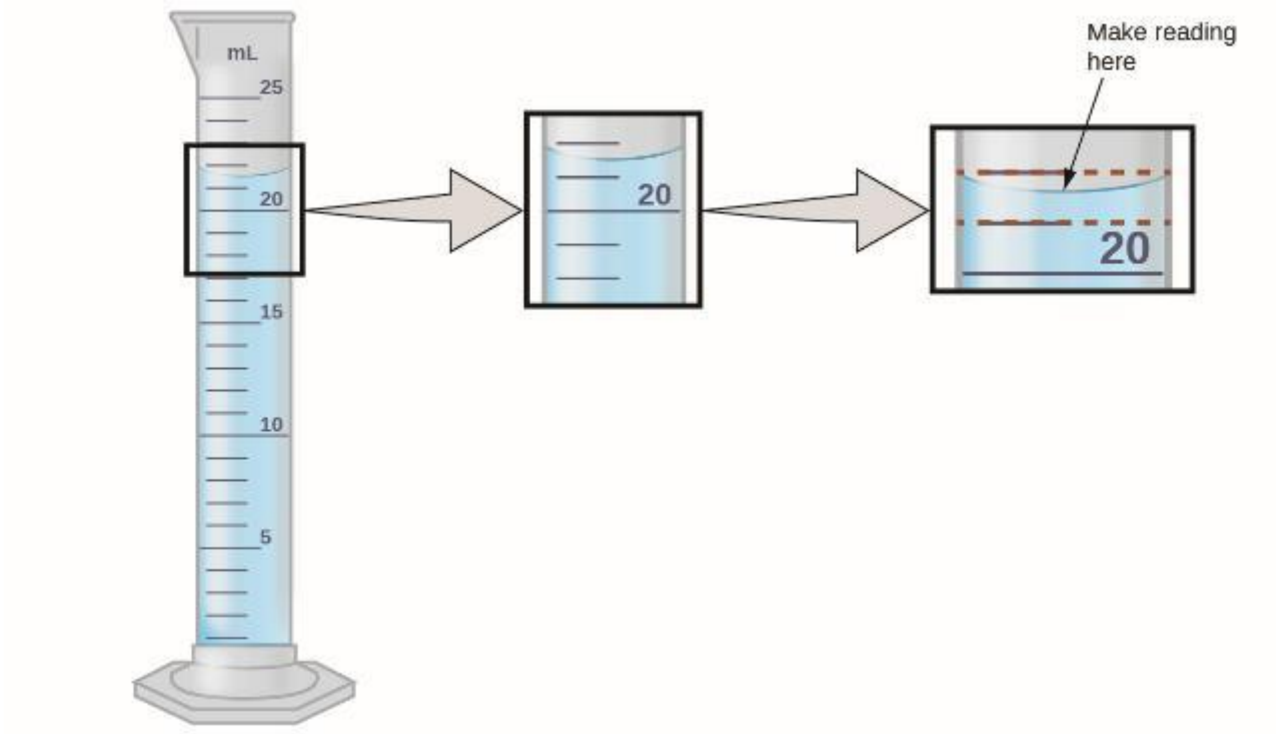
- These numbers are ***not exact.***

- There are always practical limitations of the measurement process used.

- A measured number must be reported in a way to indicate its uncertainty.

- In general, when recording a measurement you are allowed to estimate one uncertain digit.

MEASUREMENT UNCERTAINTY



To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

SIGNIFICANT FIGURES

- On the previous slide, if one recorded the volume in the graduated cylinder to be 21.6 mL.
  - 2 and 1 are certain digits
  - 6 is an estimate.
  - Someone else might perceive the volume to be 21.5 mL or 21.7mL.

- All of the digits in a measurement, including the uncertain last digit, are called ***significant figures*** or ***significant digits.***

- Frequently we need to know the number of significant figures in a measurement reported by someone else.

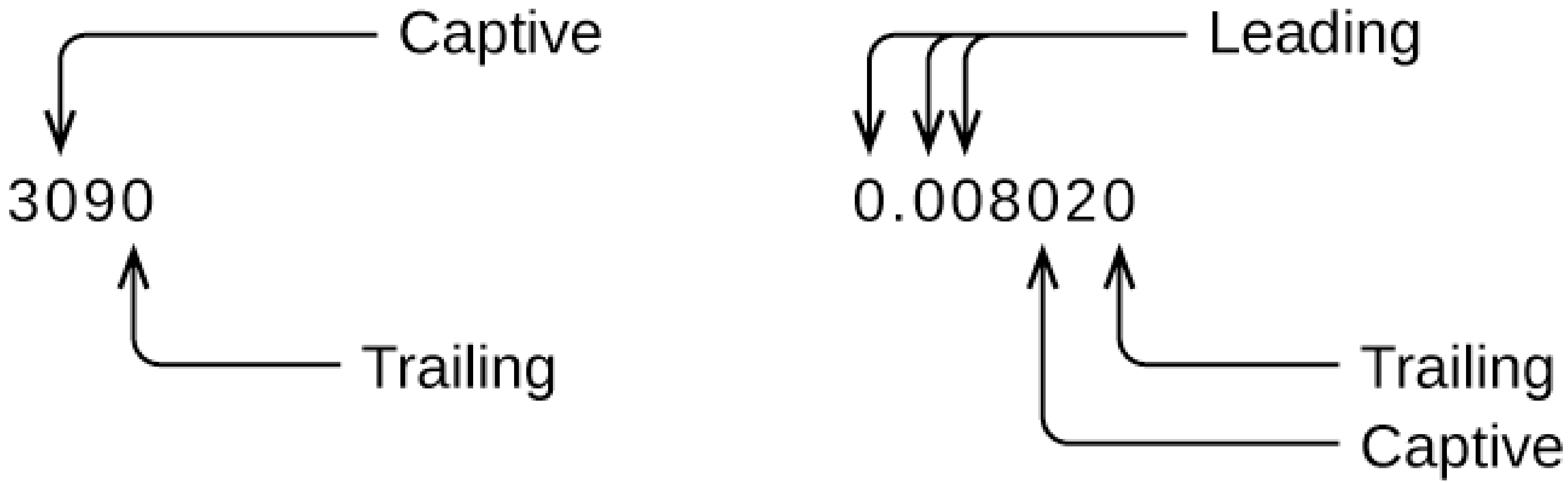
SIGNIFICANT FIGURES

These numbers are always significant.

- Nonzero digits
- Captive zeros
- Trailing zeroes
  - When they are to the right of the decimal place.
  - When in scientific notation.
    - We deal with BIG numbers of SMALL things in chemistry, so exponents are used often (positive and negative, respectively).
  - (TIP: write numbers in scientific notation… if the zeroes are there, they were significant! No need to memorize rules!)

These numbers are always not significant.

- Leading zeros
- Trailing zeros
  - When they are to the left of the decimal place.


Captive
3090
Trailing
Leading
0.008020
Trailing
Captive

SIGNIFICANT FIGURES IN CALCULATIONS

Results calculated from measured numbers are at least as uncertain as the measurement itself.

1) When we add or subtract numbers, we should round the result to the same number of decimal places as the number with the least number of decimal places (the least precise value in terms of addition and subtraction).
2) When we multiply or divide numbers, we should round the result to the same number of digits as the number with the least number of significant figures (the least precise value in terms of multiplication and division).
3) If the digit to be dropped (the one immediately to the right of the digit to be retained) is less than 5, we “round down” and leave the retained digit unchanged; if it is more than 5, we “round up” and increase the retained digit by 1; if the dropped digit *is* 5, we round up or down, whichever yields an even value for the retained digit.

SIGNIFICANT FIGURES IN CALCULATIONS

The following examples illustrate the application of this rule in rounding a few different numbers to three significant figures:

0.028675 rounds “up” to 0.0287 (the dropped digit, 7, is greater than 5)

18.3384 rounds “down” to 18.3 (the dropped digit, 3, is less than 5)

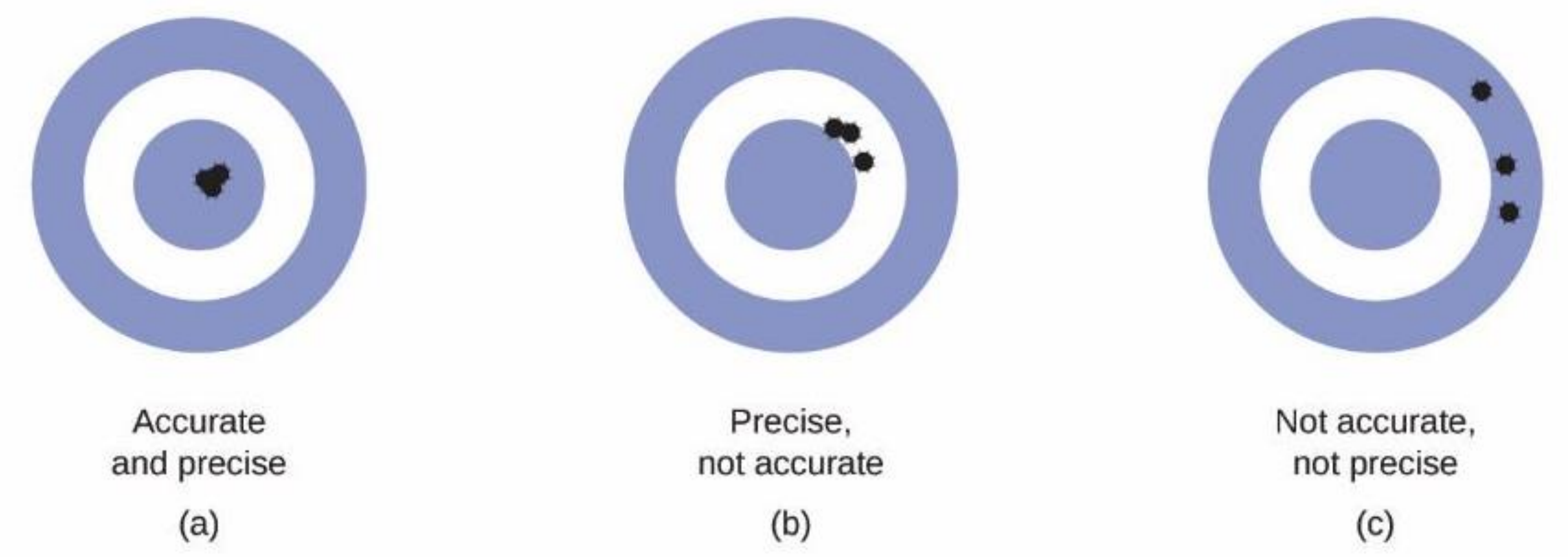
6.8752 rounds “up” to 6.88 (the dropped digit is 5, and the retained digit is even)

92.85 rounds “down” to 92.8 (the dropped digit is 5, and the retained digit is even)

ACCURACY AND PRECISION

- Measurements are said to be ***precise*** if they yield very similar results when repeated in the same manner.

- A measurement is considered ***accurate*** if it yields a result that is very close to the true or accepted value.



- (What's missing?)

ACCURACY AND PRECISION

Volume (mL) of cough medicine delivered by 10-oz (296 mL) dispensers

Dispenser #1	Dispenser #2	Dispenser #3
283.3	298.3	296.1
284.1	294.2	295.9
283.9	296.0	296.1
284.0	297.8	296.0
284.1	293.9	296.1

Dispenser #1 is precise, but not accurate.
Dispenser #2 is more accurate but less precise.
Dispenser #3 is both accurate and precise.

1.6 MATHEMATICAL TREATMENT OF MEASUREMENT RESULTS

- A quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships.

- The mathematical approach we will be using is known as ***dimensional analysis (factor labelling, multiple ratios…).***

- ***Dimensional analysis*** is based on the premise: ***the units of quantities must be subjected to the same mathematical operations as their associated numbers.***

-You can solve almost ANY problem using this method!!!

CONVERSION FACTORS AND DIMENSIONAL ANALYSIS

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- A ***conversion factor*** or ***unit conversion factor*** is a ratio of two equivalent quantities expressed with different measurement units.

- Example: The lengths 2.54 cm and 1 in. are equivalent.

$$2.54\ cm = 1\ in.$$

$$\frac{2.54\ cm}{1\ in.} \quad or \quad \frac{1\ in.}{2.54\ cm}$$

COMMON CONVERSION FACTORS

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Length	Volume	Mass
1 m = 1.0936 yd	1 L = 1.0567 qt	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft3 = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1 tbsp = 14.787 mL	1 (troy) oz = 31.103 g

Note: I will always give you conversion factors. No need to memorize.

CONVERSION FACTORS AND DIMENSIONAL ANALYSIS

- We must use the form of the conversion factors that results in the original unit canceling out, leaving only the sought unit.

- Example: Convert 34 in. to cm.

$$34\ in. \times \frac{2.54\ cm}{1\ \cancel{in.}} = 86\ cm$$

CONVERSION OF TEMPERATURE UNITS

- Temperature refers to the hotness or coldness of a substance.

- **Celsius scale**

    Water freezes at 0 ° C
    Water boils at 100 ° C

- **Fahrenheit scale**

    Water freezes at 32 ° F
    Water boils at 212 ° F

100 °C covers the same temperature interval as 180 °F.

CONVERSION OF TEMPERATURE UNITS

- The SI unit of temperature is the kelvin (K).

- Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale.

- Zero kelvin corresponds to the lowest temperature that can theoretically be achieved.

- **Kelvin scale**

    Water freezes at 273.15 K
    Water boils at 373.15 K

100 °C covers the same temperature interval as 100 K.

MATHEMATICAL RELATIONSHIPS BETWEEN TEMPERATURE SCALES

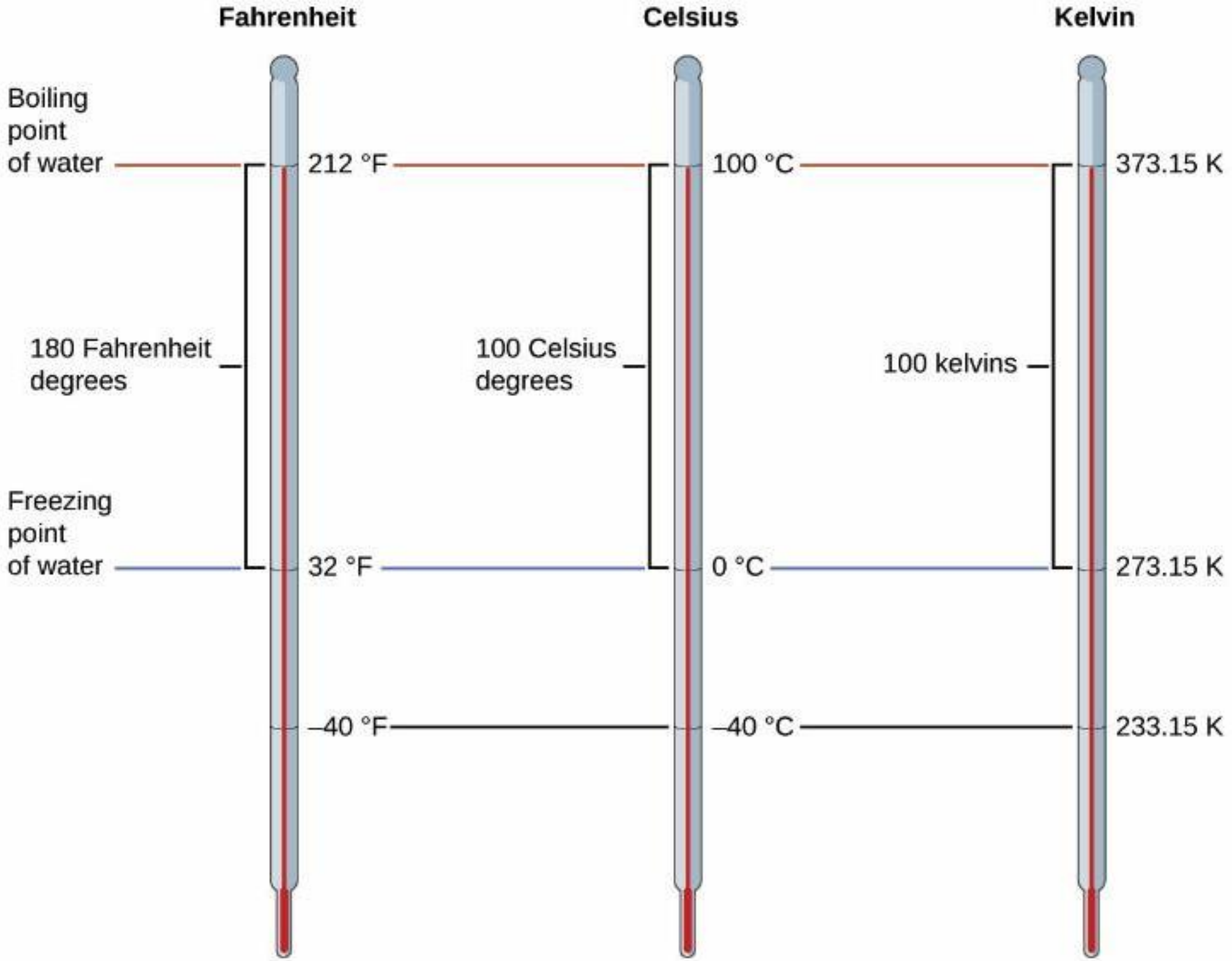
Fahrenheit and Celsius

$$T_{^\circ F} = \frac{9}{5} \times T_{^\circ C} + 32$$

Note: 9/5 = 1.8
Many textbook authors did not have calculators in high school when they learned this topic!

Kelvin and Celsius

$$T_K = T_{^\circ C} + 273.15$$



The Fahrenheit, Celsius, and kelvin temperature scales are compared.