CHEMISTRY
an atoms-focused approach

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Chapter 1

Properties of Gases

The Air We Breathe

Chapter Outline

Chemistry: Definitions

- Chemistry -the study of the composition, structure, and properties of matter and of the energy consumed or given off when matter undergoes a change
- Matter - anything that occupies space and has mass
- Mass - defines the quantity of matter in an object
- Energy – capacity to do work

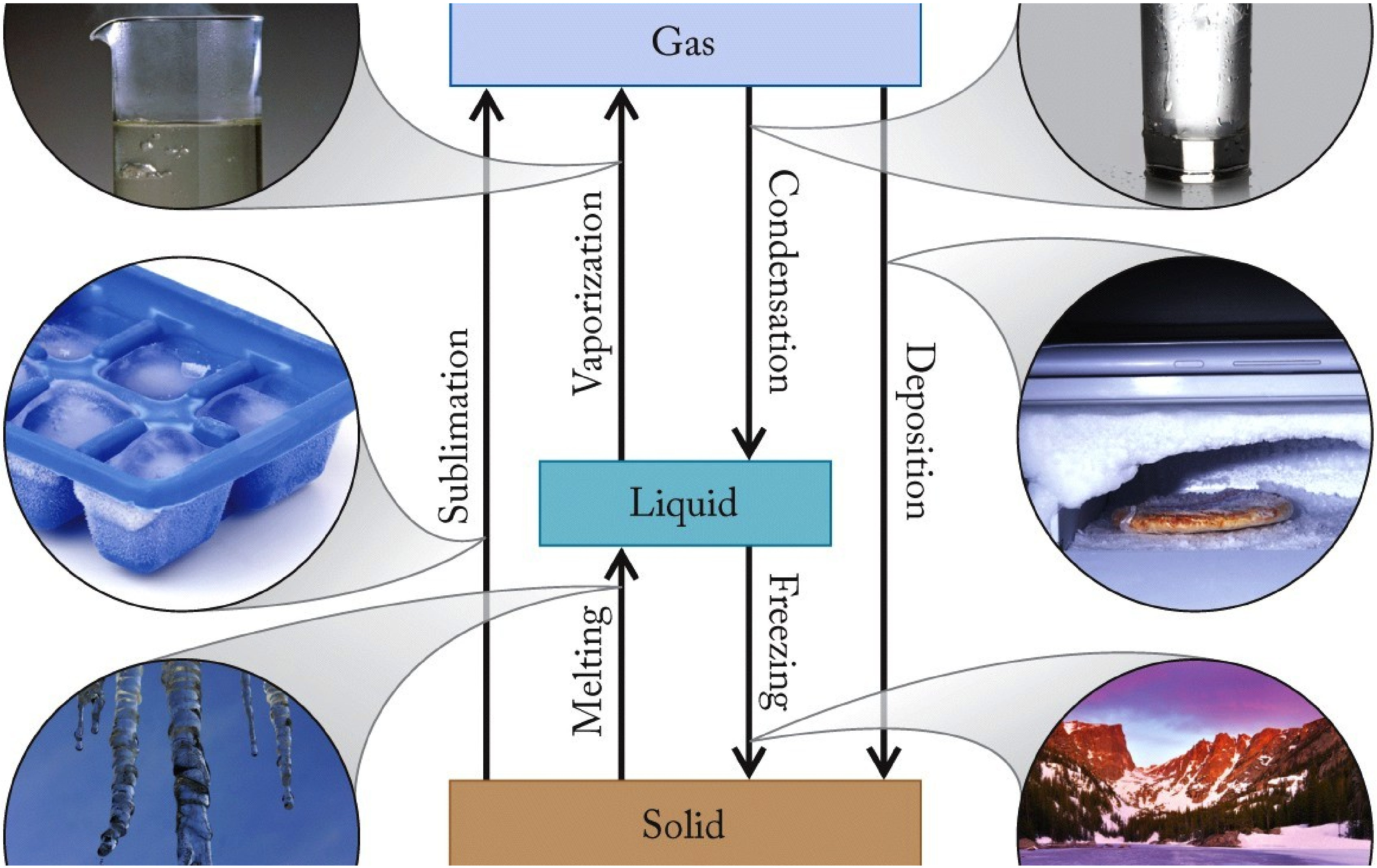
Particles Making Up Matter

- Atoms – tiny, indivisible particles that are the building blocks of matter
- Molecules – groups of atoms held together in particular pattern and proportion
- Chemical Bonds – forces that hold atoms together
  - One type forms molecules
  - Another type forms ionic compounds

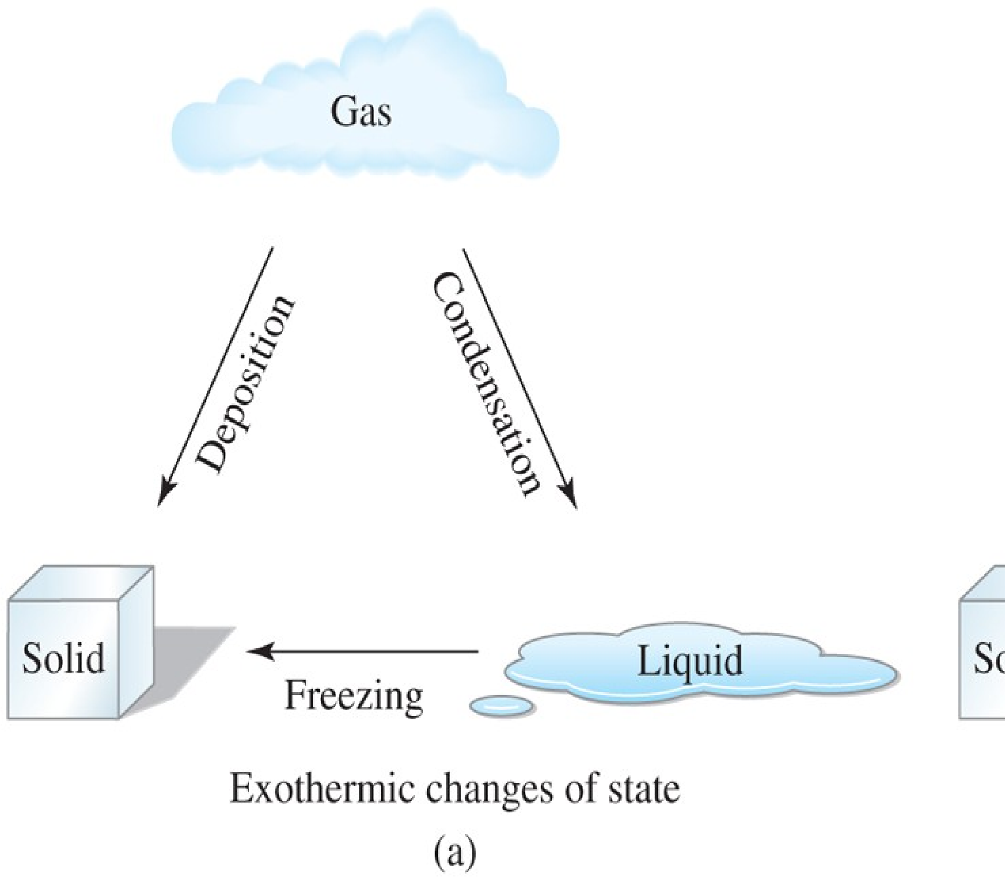
States of Matter

- Solids – definite shape and volume.
- Liquids – definite volume but flows to assume the shape of its container.
- Gases (vapors) – neither definite volume nor shape; expands to fill its container.

Changes in State



Changes in State



Chapter Outline

1.1 States of Matter
1.2 Forms of Energy
1.3 Classes of Matter
1.4 Properties of Matter
1.5 Atomic Theory: The Scientific Method in Action
1.6 A Molecular View
1.7 COAST: A Framework for Solving Problems
1.8 Making Measurements and Expressing Results
1.9 Unit Conversions and Dimensional Analysis
1.10 Temperature Scales

Forms of Energy

- Work
- Heat
- Potential Energy
- Kinetic Energy
- Law of Conservation of Energy

Chapter Outline

Classes of Matter

- Types of Matter:
  - Pure substances
    - same physical and chemical properties throughout
    - cannot be separated into simpler substances by physical processes
  - Mixtures
    - combination of two or more pure substances
    - can be separated by physical processes

Pure Substances

- Elements
  - a pure substance that cannot be separated into simpler substances by any chemical process.
- Compounds
  - pure substance composed of two or more elements bonded together in fixed proportions.
  - can be broken down into individual elements by chemical means.

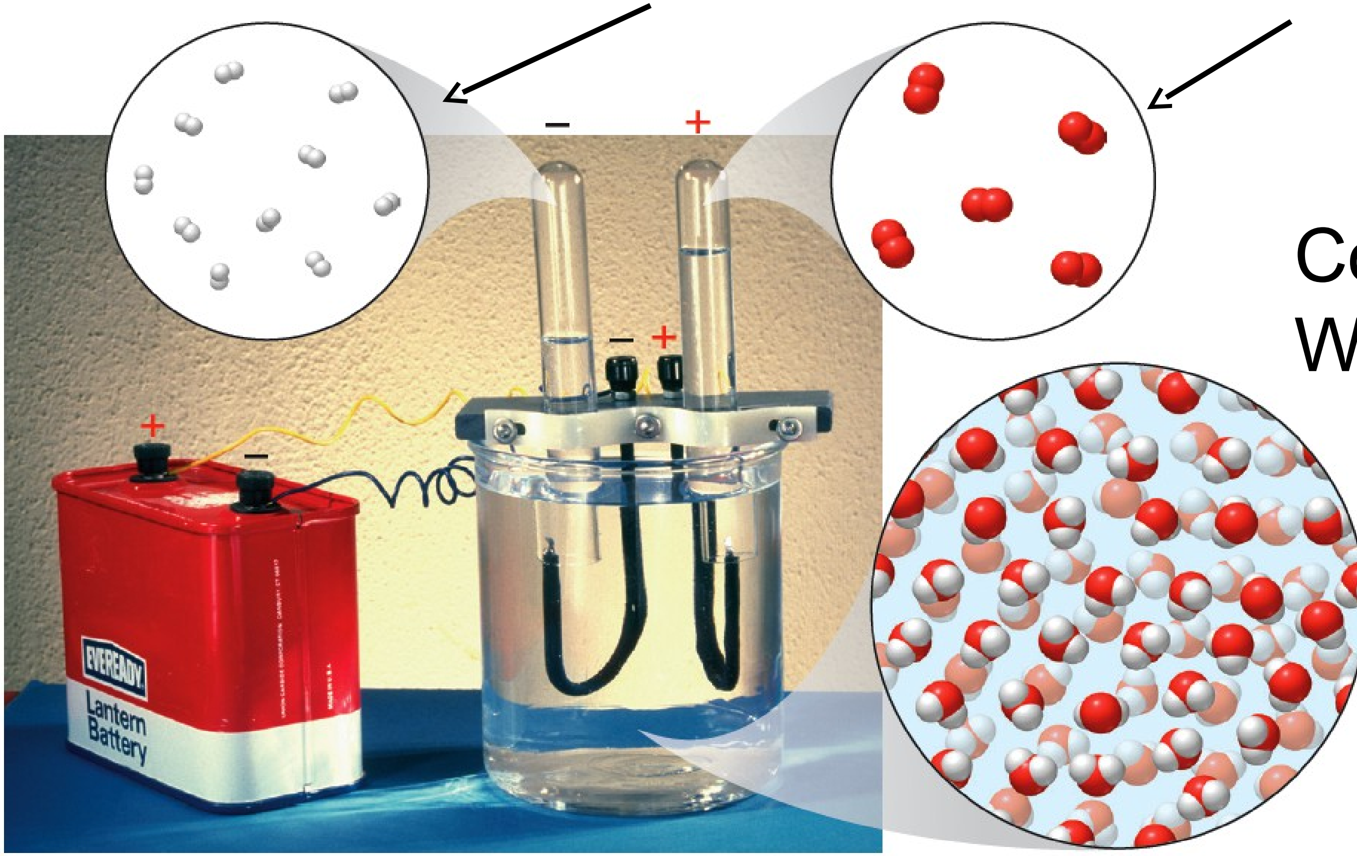
Compounds (cont.)

- Law of Constant Composition
  - All samples of a particular compound contain the same elements combined in the same proportions.
  - Example: water (H_2O)
    - Consists of two units of hydrogen (H) combined with one unit of oxygen (O)
    - Elements and proportions represented by chemical formula (Section 1.6).

Figure 1.7 Pure Substances

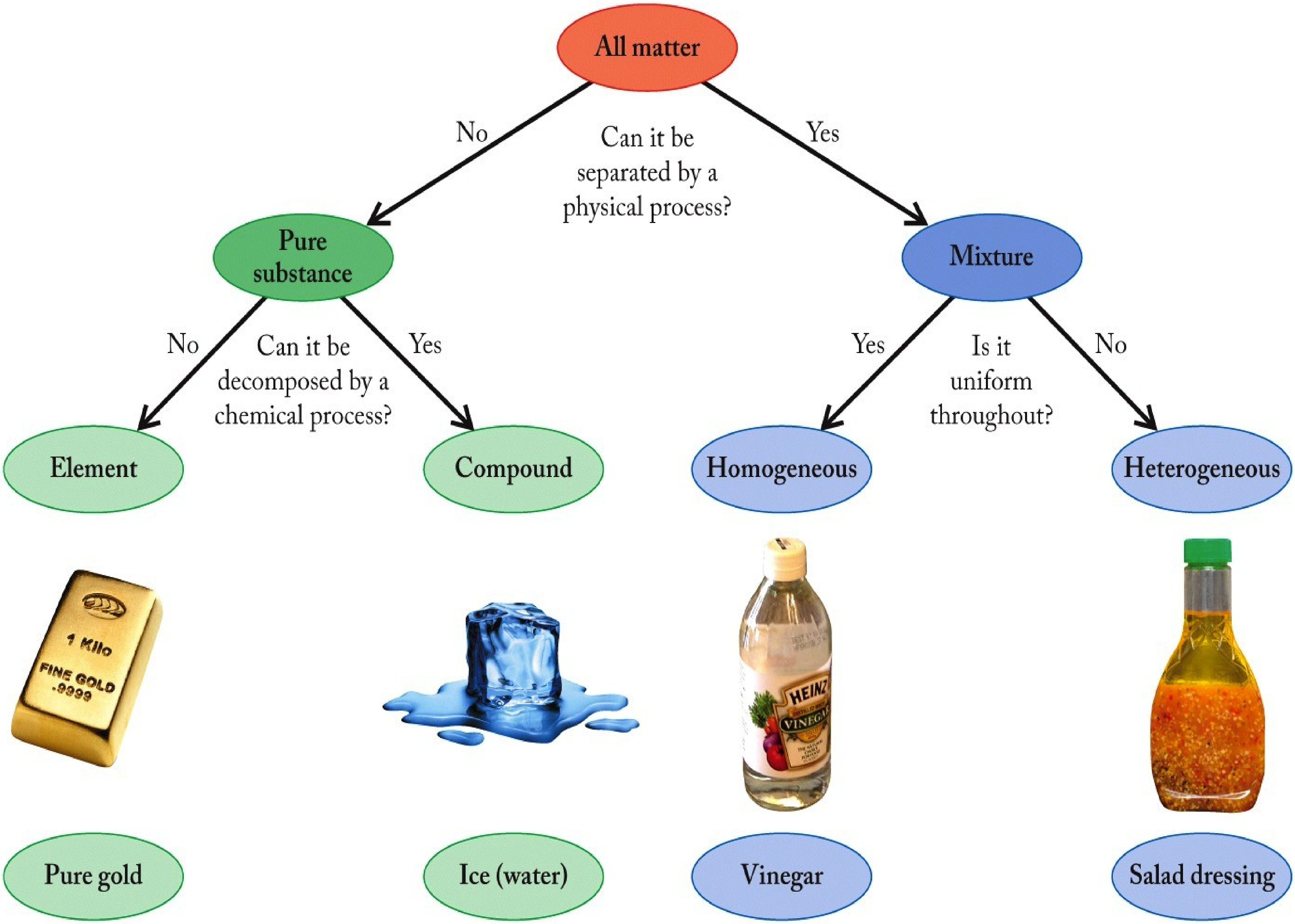
Elements: Hydrogen (H_2) and Oxygen (O_2)

Compound: Water (H_2O)



Mixtures

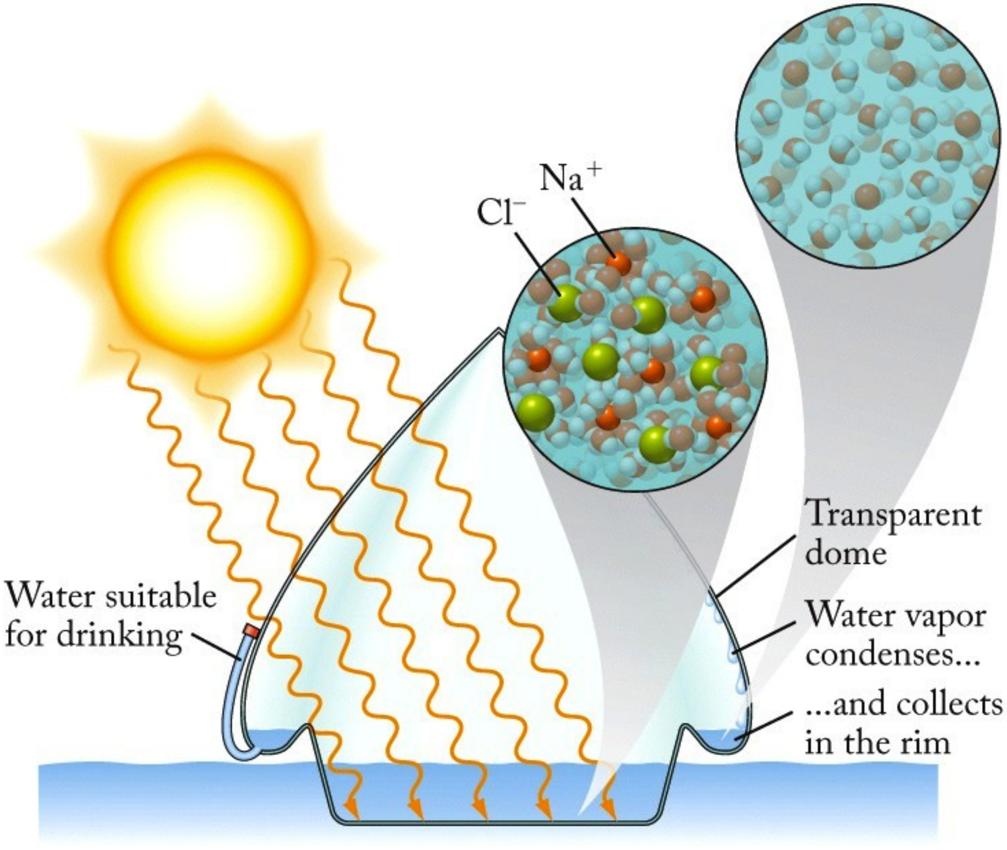
- Homogeneous
  - Also known as solutions, its components are distributed uniformly throughout the sample and have no visible boundaries or regions.
- Heterogeneous
  - Components are not distributed uniformly, contains distinct regions of different composition.


All matter
Can it be separated by a physical process?
No
Yes
Pure substance
Mixture
Can it be decomposed by a chemical process?
No
Yes
Is it uniform throughout?
Yes
No
Element
Compound
Homogeneous
Heterogeneous
1 Kilo
FINE GOLD .9999
HEINZ
VINEGAR
Pure gold
Ice (water)
Vinegar
Salad dressing

Separating Mixtures

- Distillation
  - separation using different boiling points
- Filtration
  - separation by size
- Chromatography
  - separation by solubility

Distillation



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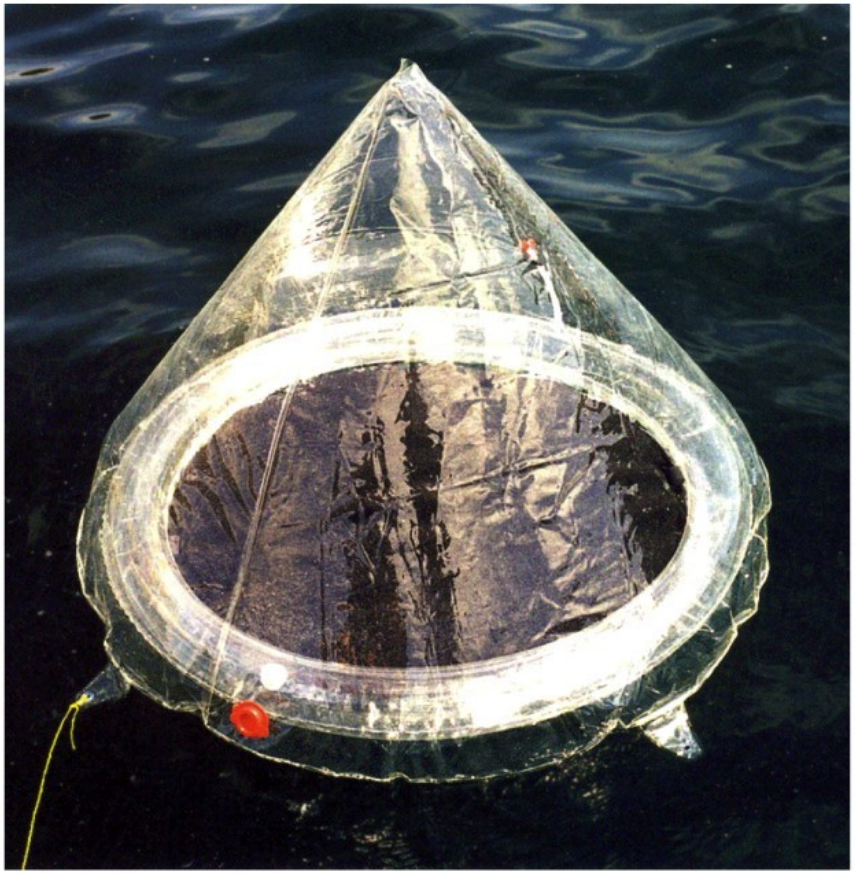

(b)

Fig. 1.8 Collection of pure water from seawater using a solar still.

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Properties of Matter

- Intensive property
  - A property that is independent of the amount of substance present
  - Examples: color, melting point
- Extensive property
  - A property that varies with the amount of the substance present
  - Examples: volume, mass

Properties of Matter

- Physical property
  - Property of a substance that can be observed without changing it into another substance
  - Examples: luster, hardness, solubility, etc.
- Chemical property
  - Property of a substance that can be observed only by reacting it to form another substance
  - Example: flammability

Density

- Density – a measure of how tightly packed the particles in a substance are
- mass of substance per unit volume of the substance

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

(Density: intensive or extensive property?)

Chapter Outline

Scientific Method

- Scientific Method – general explanation of widely observed phenomena
- Hypothesis
- Scientific theory

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A Molecular View

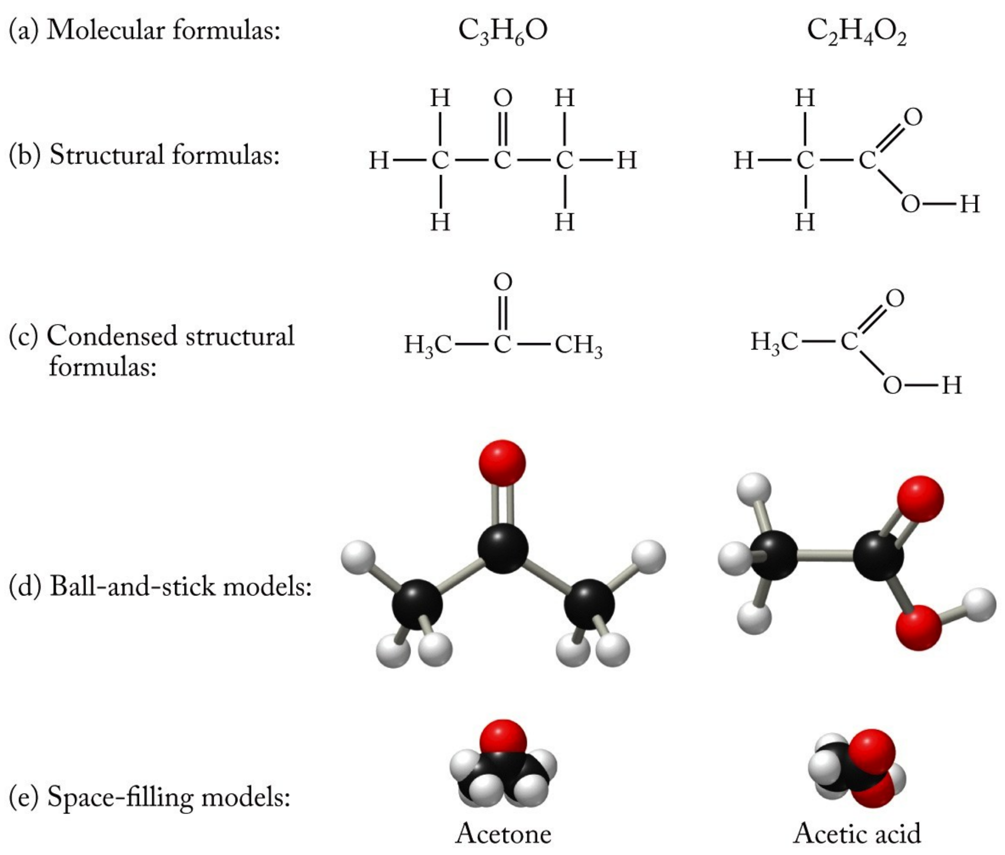
- Chemical Formula
  - Notation for representing elements and compounds
  - Consists of symbols of constituent elements, and subscripts identifying the number of atoms of each element in one molecule.

- Molecular Formula

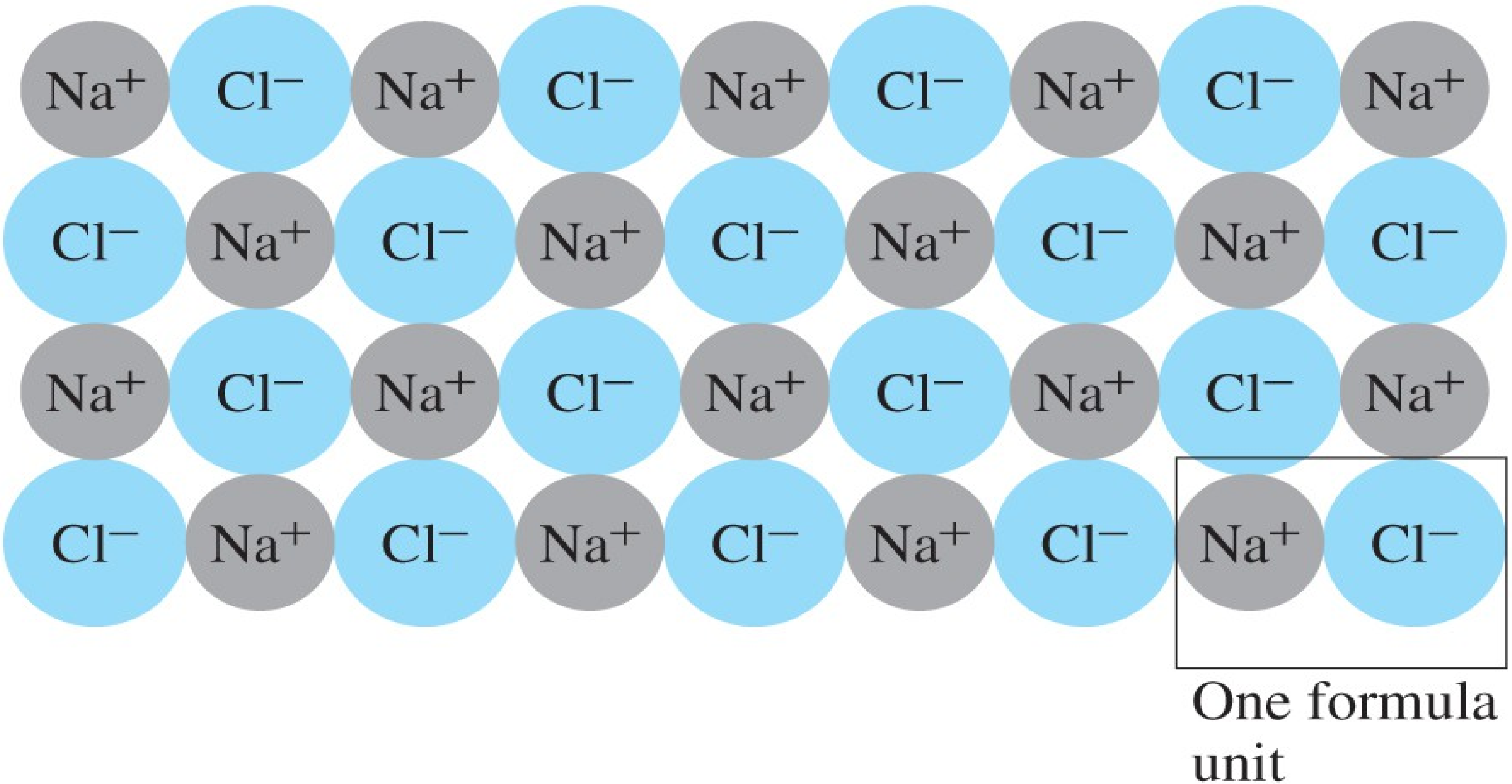
- Structural Formula

Chemical Formulas

- Some chemical bonds link atoms together to make molecules.
- Chemical formulas can be represented in four ways:
  - Molecular formulas
  - Structural formulas
  - Condensed structural formulas
  - Ball-and-stick models
  - Space-filling models



Ionic Compounds



Chapter Outline

COAST

- Collect and Organize
  - Identify key concepts, skills required to solve problem, and assemble information needed.
- Analyze
  - Evaluate information and relationships or connections; sometimes units will help identify steps needed to solve the problem.
- Solve
  - Perform calculations, **check units**, etc.
- Think about it
  - Is the answer reasonable? Are the units correct?

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Making Measurements

- Measurements
  - Essential for characterizing physical and chemical properties of matter
  - Two parts of every measurement:

1875 lbs

Number → 1875; Unit → lbs

- Standardization of the units of measurement is essential.

TABLE 1.2 SI Base Units

Quantity or Dimension	Unit Name	Unit Abbreviation
Mass	kilogram	kg
Length	meter	m
Temperature	kelvin	K
Time	second	s
Electric current	ampere	A
Quantity of a substance	mole	mol

The Mole

•A mole is a very large number

1 mole = 6.022E23

•Moles are useful when counting atoms, molecules, or sub atomic particles

•Think of a mole as we do with other abbreviations

1 dozen = 12

1 gross = 144

1 thousand = 1000

1 mole = 6.022E23

(1 gram of hydrogen contains 1 mole of atoms)

TABLE 1.1 Commonly Used Prefixes for SI Units

PREFIX		VALUE	
Name	Symbol	Numerical	Exponential
zetta	Z	1,000,000,000,000,000,000,000	10^{21}
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}
zepto	z	0.000000000000000000001	10^{-21}

SI Measurements

The prefix is placed before the base unit.

This multiplies the base unit and creates a new larger or smaller unit.

1 m = 1 meter

1 km = 1 * 1000 * m = 1000 meters

1 cm = 1 * 1E-2 * m = 0.01 meters

1 mg = 1 * 1E-3 * g = 0.001 grams

The Liter

- The liter is a derived unit since it is based on the meter:
  Volume = length x width x height
  Liter = 0.1 m * 0.1 m * 0.1 m

- One liter is a little more than a quart

- 1 mL = 0.001 L = 1 cm x 1 cm x 1 cm = 1 cm^3

- 1 cm^3 = 1 mL = 1 cc (cubic centimeter)

- The mass of 1 mL of water is 1 gram

TABLE 1.3 Conversion Factors for SI and Other Commonly Used Units

Quantity or Dimension	Equivalent Units
Mass	1 kg = 2.205 pounds (lb); 1 lb = 0.4536 kg = 453.6 g 1 g = 0.03527 ounce (oz); 1 oz = 28.35 g
Length (distance)	1 m = 1.094 yards (yd); 1 yd = 0.9144 m (exactly) 1 m = 39.37 inches (in); 1 foot (ft) = 0.3048 m (exactly) 1 in = 2.54 cm (exactly) 1 km = 0.6214 miles (mi); 1 mi = 1.609 km
Volume	1 m^3 = 35.31 ft^3; 1 ft^3 = 0.02832 m^3 1 m^3 = 1000 liters (L) (exactly) 1 L = 0.2642 gallon (gal); 1 gal = 3.785 L 1 L = 1.057 quarts (qt); 1 qt = 0.9464 L

Uncertainty in Measurements

- All measurements contain uncertainty.
  - depends on instruments used to make measurement.
- A digit that must be estimated is called uncertain (last recorded digit).

± 0.01 g

± 0.0001 g

Significant Figures

- Include all digits known with certainty plus one digit that is uncertain
- Rules for counting significant figures
  - All nonzero integers are significant.
  - Zeros depend on location.
    - leading zeros – not significant
    - “captive” zeros - significant
    - trailing zeros – significant is there is a decimal anywhere in the value
  - Exact numbers

Counting Significant Figures

- Nonzero integers - always significant.
  - 283 → 3 sig. figs.
  - 7.315 → 4 sig. figs.
- Zeros
  - leading zeros not significant.
    - 0.0392 → 3 sig. figs.

Counting Significant Figures

- Zeros (cont.)
  - trailing zeros not significant <u>unless there is a decimal anywhere in the number.</u>
    - 600 → 1 sig. fig.
    - 30.000 → 5 sig. figs.
  - Captive zeros <u>always</u> significant.
    - 50.06 → 4 sig. figs.
- Exact numbers → infinite # of sig. figs.
  - Often defined values or relationships

Practice: Significant Figures

How many significant figures are in the following numbers?

0.0280 g

2100 lb

10.50 mL

670.1 g

Significant Figures in Mathematical Operations

- Rounding off
  - Dropping “insignificant” digits
  - Only at the end of calculations!
- “Weakest link” Principle
  - The number of significant figures in the final result cannot be greater than the “weakest link” used in the calculation.
  - Actual rule depends on mathematical operation.

Significant Figures in Mathematical Operations

- Multiplication / Division:
  - # of sig figs in the result = # of sig figs in the initial value with the least number of sig figs.

  - $6.38 \times \underline{2.0} = 12.76 \rightarrow 13$ (2 sig. figs.)
  - $16.84 / \underline{2.54} = 6.6299 \rightarrow 6.63$ (3 sig. figs.)

Significant Figures in Mathematical Operations

Addition / Subtraction:

- # of sig figs in the result depends on the number of <u>decimal places</u> (precision) in the least precise measurement.

  6.8 g + 11.934 g = 18.734 g $\rightarrow$ 18.7 g

  (first decimal)

Practice: Rounding

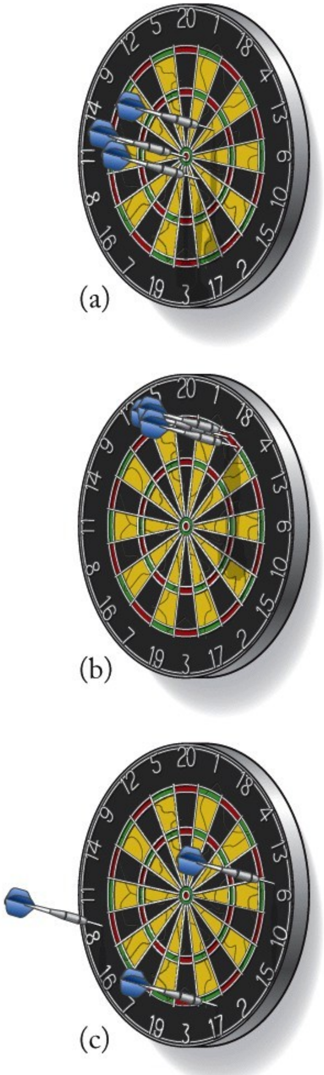
Round the answer for the mathematical operation below to the appropriate number of significant figures.

$$\frac{1.23\text{ g} - 0.567\text{ g}}{0.34442\text{ m}} = 1.924975\text{ g/m?}$$

(Answer: 1.9)

Precision and Accuracy

- Accuracy – agreement between measured value and accepted or true value.
- Precision – agreement among repeated measurements.
- Which is accurate? Which is precise?



Chapter Outline

Changing Units: Conversion Factors

We often need to change values from one unit to another.

To do this, we multiply the initial value by a conversion factor.

Conversion factors are ratios of the desired and initial units.

Conversion factors equal 1, and so the magnitude of the initial value is not changed.

Changing Units: Conversion Factors

- Conversion factor
  - a fraction in which the numerator and denominator represent equivalent quantities, but expressed in different units.
  - 1 km = 0.6214 mi → $\frac{1\ \text{km}}{0.6214\ \text{mi}}$ or $\frac{0.6214\ \text{mi}}{1\ \text{km}}$
- Converting a value from one unit to another:

$$\cancel{\text{initial units}} \times \frac{\text{desired units}}{\cancel{\text{initial units}}} = \text{desired units}$$

Changing Units: Conversion Factors

Most conversion factors come from equalities:

e.g. to convert 2.55 m to mm, we use the definition of mm:
 1 mm = 1E-3 m

so to perform this conversion:

$$2.55\,m\left(\frac{1\,mm}{1\text{E-}3\,m}\right)=2.55\text{E}3\,mm$$

To convert 9.33E2 g to cg:

Start with the initial value.

$$9.33\text{E}2 \text{ g} \left(\frac{1 \text{ cg}}{1\text{E-}2 \text{ g}}\right) = 9.33\text{E}4 \text{ cg}$$

Place units in the c. f. to cancel original units.

Place units in the c. f. to obtain desired units.

Place numbers in the c. f. to make the top and bottom of the factor equal to each other.

Complete the calculation, canceling units.

Using Density as a Conversion Factor

As a conversion factor:

- to convert masses into volumes.

$$18.550\ \cancel{g}\ \text{of copper}\left[\frac{1\ \text{mL}}{8.96\ \cancel{g}}\right] = 2.07\ \text{mL of Cu}$$

- to convert volumes into masses.

$$2.07\ \cancel{\text{mL}}\ \text{of Cu}\left[\frac{8.96\ \text{g}}{1\ \cancel{\text{mL}}}\right] = 18.6\ \text{g of Cu}$$

Practice: Using Density

The density of Ti is 4.50 g/cm^3. What is the volume of 7.20 g of Titanium?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Sig. Figs. and Conv. Factors

Conversion factors based on defined relationships are considered exact and have infinite sig. figs.:
12 in. = 1 ft.
1 km = 1E3 m
60 s = 1 min

For measured relationships, the sig. figs. depend on the precision of the measurements:
1 L = 0.264 gallons   (3 sig. figs.)

Howerver, “1 in. = 2.54 cm” is an exact definition and so has an infinite number of sig. figs.

Chapter Outline

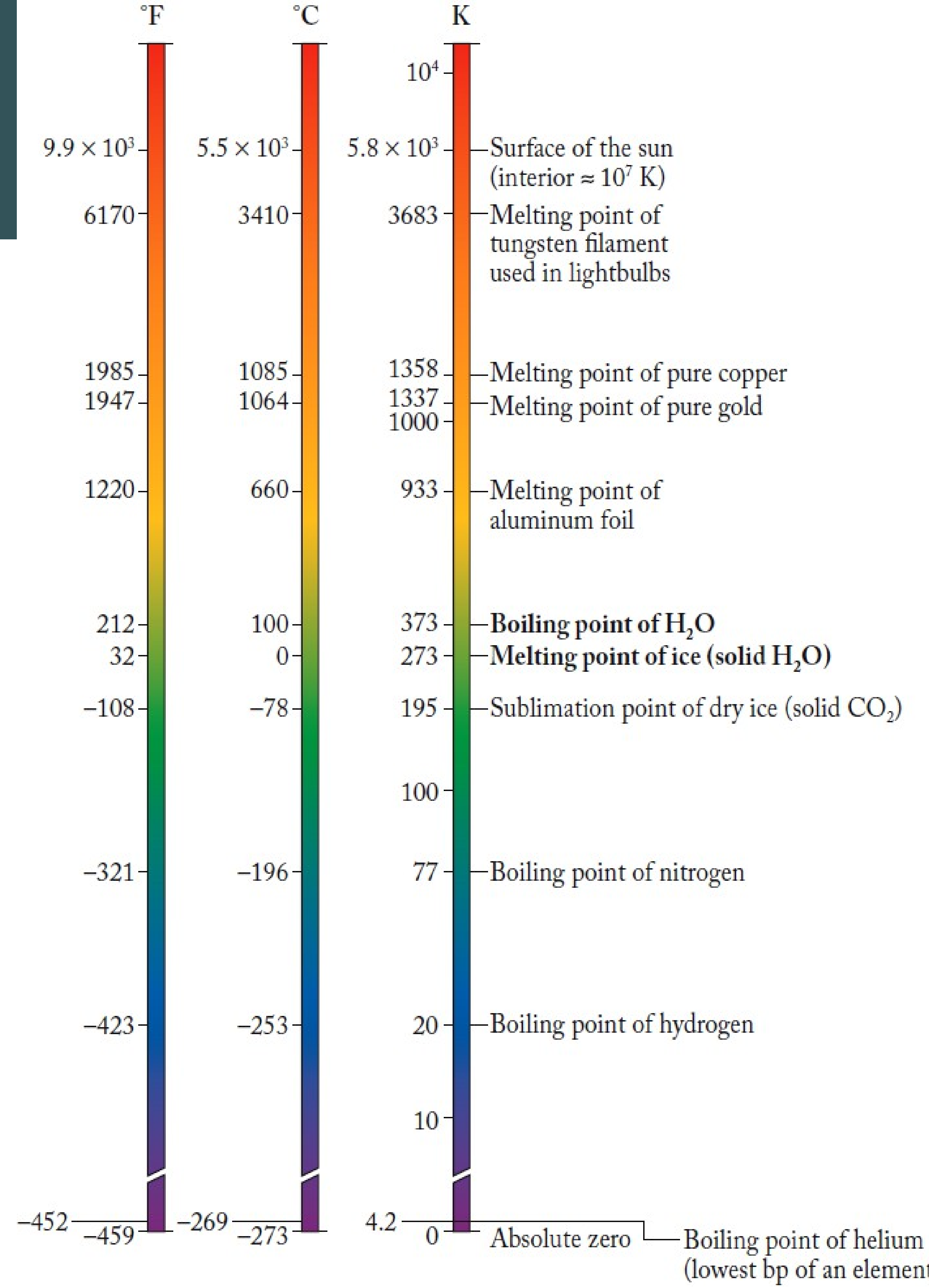
Temperature Scales

Fahrenheit (°F)
Celsius (°C)
Kelvin (K)

Temperature Conversions:

$$K = °C + 273.15$$

$$°C = 5/9\,(°F - 32)$$



Practice: Temperature Conversions

The lowest temperature measured on the Earth is -128.6 °F, recorded at Vostok, Antarctica, in July 1983. What is this temperature in °C and in Kelvin?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It: