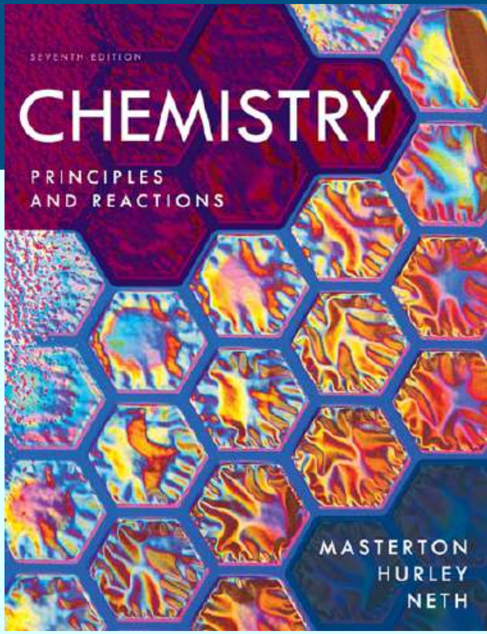

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Chapter 1
Matter and Measurements

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Chemistry

- Chemistry is concerned with matter and energy and how the two interact with each other
- Chemistry is a foundation for other disciplines
  - Engineering
  - Health sciences
  - Pharmacy and pharmacology
  - Scientific literacy

Current Issues with Chemical Relevance

- Chemistry-related issues
  - Depletion of the ozone layer
  - Alternative sources of fuel
  - Nuclear energy

Outline

- Matter
- Measurements
- Properties of substances

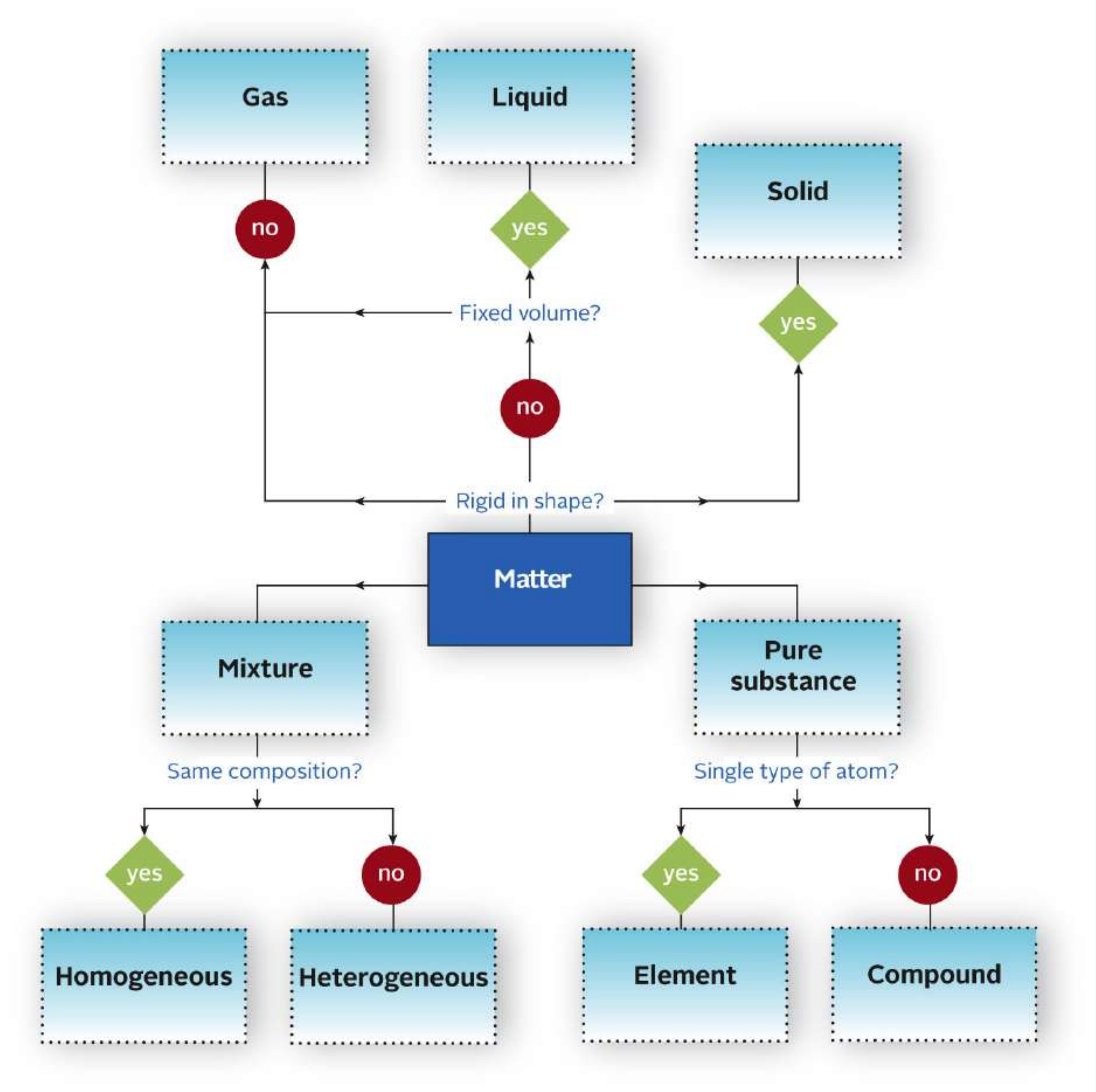
Matter

- Matter has mass
  - Weight is what we normally consider
- Matter occupies space
- Phases of matter
  - Solids
    - Fixed volume and shape
  - Liquids
    - Fixed volume, indefinite shape
  - Gases
    - Indefinite shape and volume

Matter

- Pure substances
  - Fixed composition
  - Unique set of properties
- Mixtures
  - Two or more substances in some combination

Figure 1.1 - Classification of Matter



Elements

- Elements cannot be broken down into two or more pure substances
  - 115 elements; 91 occur naturally
- Common elements
  - Carbon (found in charcoal)
  - Copper (found in pipes, jewelry, etc.)
- Rare elements
  - Gold
  - Uranium

Atomic Symbols

- Elements are given symbols
  - Chemical identifier
  - Elements known to ancient times often have symbols based on Latin names
    - Copper, Cu (cuprum)
    - Mercury, Hg (hydrargyrum)
    - Potassium, K (kalium)
  - One element has a symbol based on a German name
    - Tungsten, W (wolfram)

Table 1.1 - Elements and Abundances

- Some elements are common, some are rare

TABLE 1.1 Some Familiar Elements with Their Percentage Abundances

Element	Symbol	Percentage Abundance	Element	Symbol	Percentage Abundance
Aluminum	Al	7.5	Manganese	Mn	0.09
Bromine	Br	0.00025	Mercury	Hg	0.00005
Calcium	Ca	3.4	Nickel	Ni	0.010
Carbon	C	0.08	Nitrogen	N	0.03
Chlorine	Cl	0.2	Oxygen	O	49.4
Chromium	Cr	0.018	Phosphorus	P	0.12
Copper	Cu	0.007	Potassium	K	2.4
Gold	Au	0.0000005	Silicon	Si	25.8
Hydrogen	H	0.9	Silver	Ag	0.00001
Iodine	I	0.00003	Sodium	Na	2.6
Iron	Fe	4.7	Sulfur	S	0.06
Lead	Pb	0.0016	Titanium	Ti	0.56
Magnesium	Mg	1.9	Zinc	Zn	0.008

Compounds

- Compounds are combinations of two or more elements
  - Carbon and hydrogen
    - Hydrocarbons
    - Methane, acetylene, naphthalene
    - Different proportions of each element

Composition of Compounds

- Compounds always contain the same elements in the same composition by mass
  - Water by mass:
    - 11.19% hydrogen
    - 88.81% oxygen
- Properties of compounds are often very different from the properties of elements from which the compounds form

Resolving compounds into elements

- Many methods
  - Heating mercury(II) oxide releases mercury and oxygen
    - Priestley, 200 years ago
  - Aluminum
    - Not known until about 100 years ago
    - Difficult to resolve aluminum from rocks and minerals where it is commonly found
  - Electrolysis is required to prepare aluminum from its compounds

Mixtures

- Two or more substances in such a combination that each substance retains a separate chemical identity
  - Copper sulfate and sand
    - Identity of each is retained
  - Contrast with the formation of a compound
    - Sodium and chlorine form sodium chloride

Mixtures

- Homogeneous mixtures
  - Uniform
  - Composition is the same throughout
  - Example: seawater
- Heterogeneous mixtures
  - Not uniform
  - Composition varies throughout
  - Example: rocks

Figure 1.3 – Sodium, Chlorine and Sodium Chloride

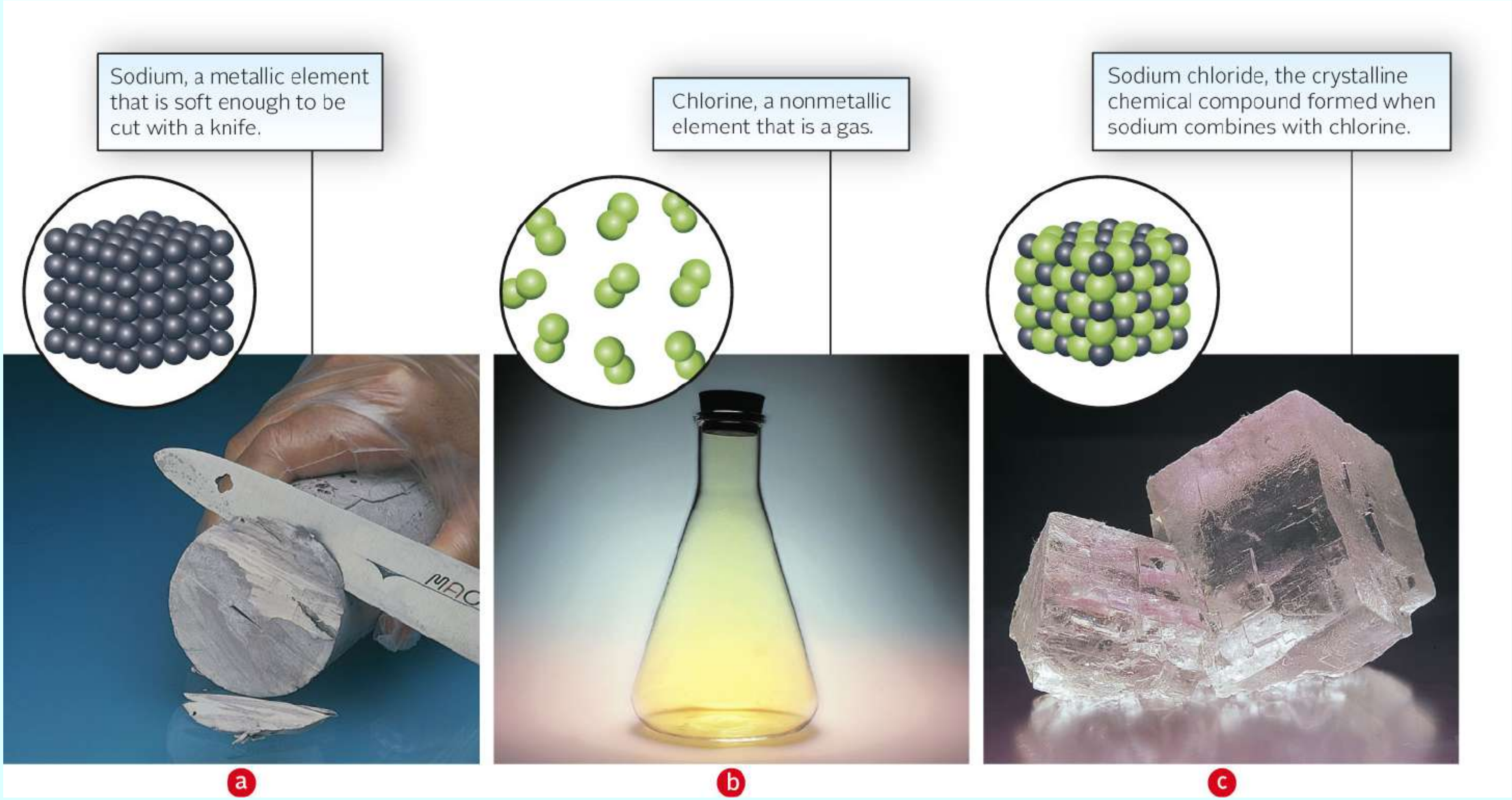


Figure 1.2 – Cinnabar and Mercury

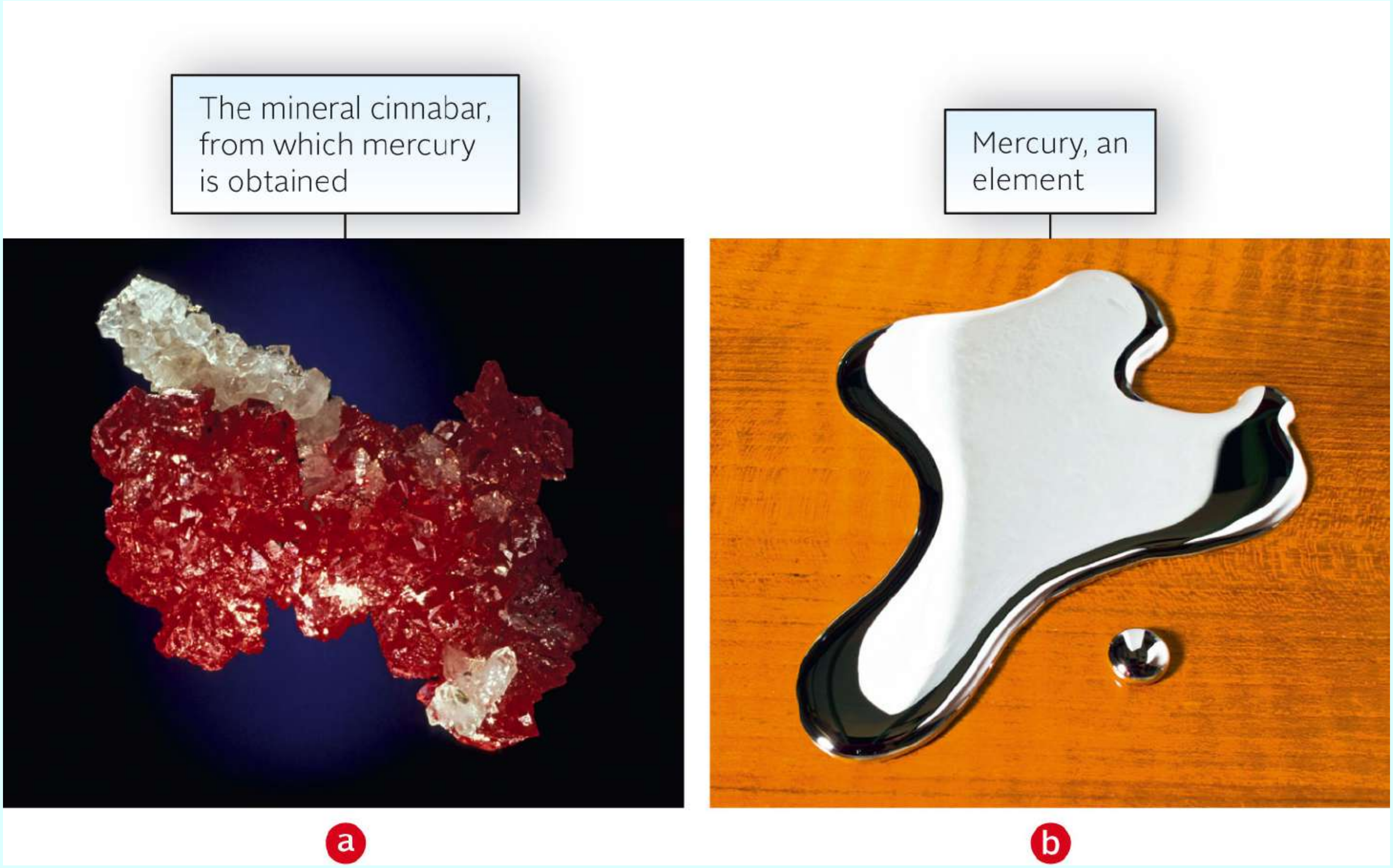
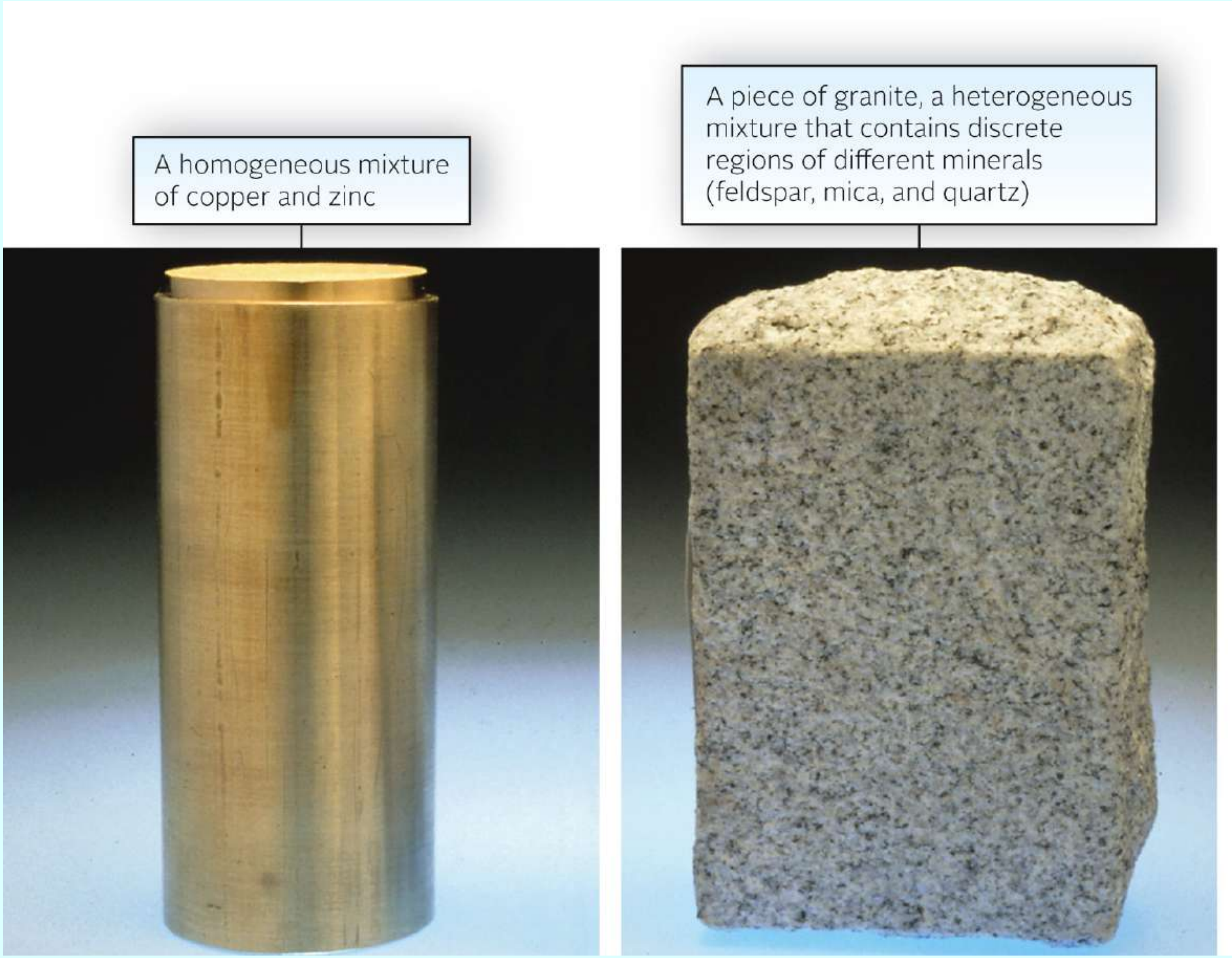


Figure 1.4 – Copper Sulfate and Sand

Figure 1.5 – Two Mixtures



Solutions

- Common homogeneous mixture
  - Components
    - Solvent
      - Most commonly a liquid
    - Solute
      - May be solid, liquid or gas
  - Seawater
    - Water is the solvent
    - Solutes may be one of a variety of salts

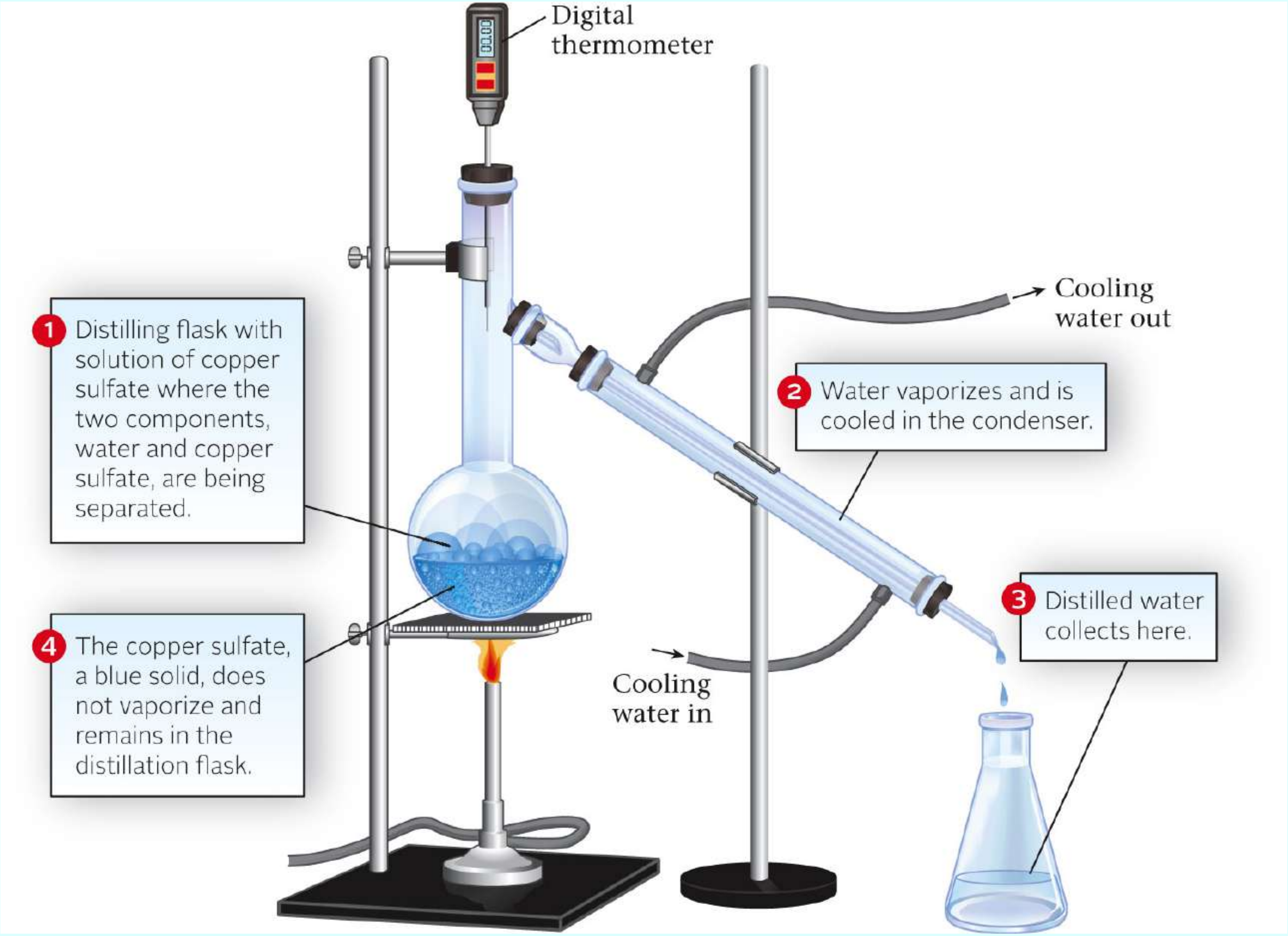
Separating Mixtures

- Filtration
  - Separate a heterogeneous solid-liquid mixture
  - Barrier holds back one part of the mixture and lets the other pass
  - Filter paper will hold back sand but allow water to pass through
- Distillation
  - Resolves homogeneous mixtures
  - Salt water can be distilled, allowing water to be separated from the solid salt

Chromatography

- Separation of mixtures in industry and research
  - Many mixtures can be separated by chromatography
    - Gas mixtures
    - Liquid mixtures

Figure 1.6 – Distillation Apparatus



Forensic Chemistry

- Forensic chemistry is the study of materials or problems where evidence is sought for criminal or civil cases tried in court
  - Chromatography is a fundamental tool of forensic chemistry
  - Biochemistry relies heavily on chromatography and on mass spectrometry, which we will briefly examine in Chapter 2

Figure 1.7 – Airport Security Scanner

Measurements

- Quantitative
  - Identify the amount of substance present
  - Chemistry is a quantitative science
- Measurement
  - Needed to quantify the amount of substance present
  - SI, the international system of measurements
    - Common name: the metric system

Metric System

- Based on the decimal
  - Powers of ten
  - Four units
    - Length
    - Volume
    - Mass
    - Temperature

Table 1.2 - Powers of Ten

TABLE 1.2 Metric Prefixes

Factor	Prefix	Abbreviation	Factor	Prefix	Abbreviation
10^{6}	mega	M	10^{-3}	**milli**	m
10^{3}	**kilo**	k	10^{-6}	micro	μ
10^{-1}	deci	d	10^{-9}	**nano**	n
10^{-2}	**centi**	c	10^{-12}	pico	p

Instruments and Units

- Length
  - In the SI system, the unit of length is the meter
    - A meter is slightly longer than a yard
    - Precise definition is the distance light travels in 1/299,792,458 of one second
- Volume
  - Volume is related to length
  - Units of volume
    - Cubic centimeters
    - Liters
    - Milliliters
      - 1 mL = 1 cm^3

Table 1.3 – Units and Unit Relations

TABLE 1.3 Relations Between Length, Volume, and Mass Units

Metric		English		Metric-English	
Length					
1 km	$= 10^3$ m	1 ft	= 12 in	1 in	= 2.54 cm*
1 cm	$= 10^{-2}$ m	1 yd	= 3 ft	1 m	= 39.37 in
1 mm	$= 10^{-3}$ m	1 mi	= 5280 ft	1 mi	= 1.609 km
1 nm	$= 10^{-9}$ m = 10 Å				
Volume					
1 m^3	$= 10^6$ cm^3 $= 10^3$ L	1 gal	= 4 qt = 8 pt	1 ft^3	= 28.32 L
1 cm^3	= 1 mL $= 10^{-3}$ L	1 qt (U.S. liq)	= 57.75 in^3	1 L	= 1.057 qt (U.S. liq)
Mass					
1 kg	$= 10^3$ g	1 lb	= 16 oz	1 lb	= 453.6 g
1 mg	$= 10^{-3}$ g	1 short ton	= 2000 lb	1 g	= 0.03527 oz
1 metric ton	$= 10^3$ kg			1 metric ton	= 1.102 short ton

*This conversion factor is exact; the inch is defined to be exactly 2.54 cm. The other factors listed in this column are approximate, quoted to four significant figures. Additional digits are available if needed for very accurate calculations. For example, the pound is defined to be 453.59237 g.

Measuring volume

- Graduated cylinder
- Pipet or buret
  - Used when greater accuracy is required

Figure 1.8 – Measuring Volume

Mass

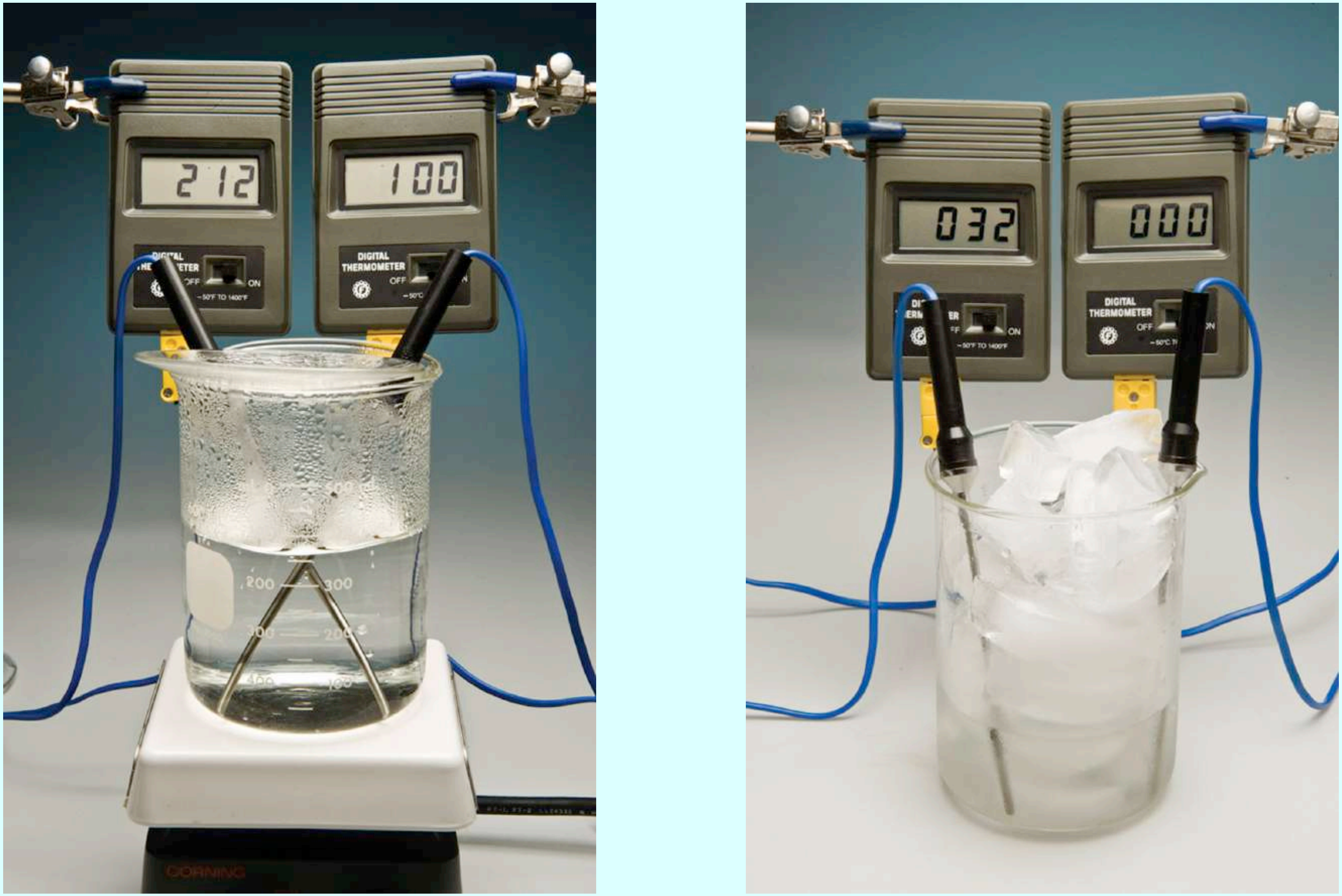
- In the metric system, mass is expressed in grams
- Powers of ten modify the unit
  - Milligram, 0.001 g
  - Kilogram, 1000 g

Figure 1.9 – Weighing a Solid

Temperature

- Factor that determines the direction of heat flow
- Temperature is measured indirectly
  - Observing its effect on the properties of a substance
  - Mercury thermometer
    - Mercury expands and contracts in response to temperature
  - Digital thermometer
    - Uses a device called a thermistor

Figure 1.10 – Fahrenheit and Celsius Scales



Temperature Units

- Degrees Celsius
  - Until 1948, degrees centigrade
- On the Celsius scale
  - Water freezes at 0 °C
  - Water boils at 100 °C

The Fahrenheit Scale

- On the Fahrenheit scale
  - Water freezes at 32 °F
  - Water boils at 212 °F
- Comparing scales
  - 0 C is 32 °F
  - 100 C is 212 °F
  - There are 180 F for 100 °C, so each °C is 1.8 times larger than each °F

The Kelvin Scale

- The Kelvin is defined as
  - 1/273.16 of the difference between the lowest attainable temperature (0 K) and the triple point of water (0.01 °C)
  - Unlike the other two scales, no degree sign is used to express temperature in K

Relationships Between Temperature Scales

- Fahrenheit and Celsius

$$t_{^\circ F} = 1.8 t_{^\circ C} + 32^\circ$$

- Celsius and Kelvin

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

Example 1.1

EXAMPLE 1.1

Mercury thermometers have been phased out because of the toxicity of mercury vapor. A common replacement for mercury in glass thermometers is the organic liquid isoamyl benzoate, which boils at 262°C. What is its boiling point in (a) °F? (b) K?

ANALYSIS

Information given:	Boiling point (262°C)
Asked for:	boiling point in °F and K

STRATEGY

1. Substitute into Equation 1.1.
2. Substitute into Equation 1.2.

SOLUTION

(a) °F	°F = 1.8(°C) + 32 = 1.8(262°C) + 32 = 504°F
(b) K	K = 273.15 + 262°C = 535 K

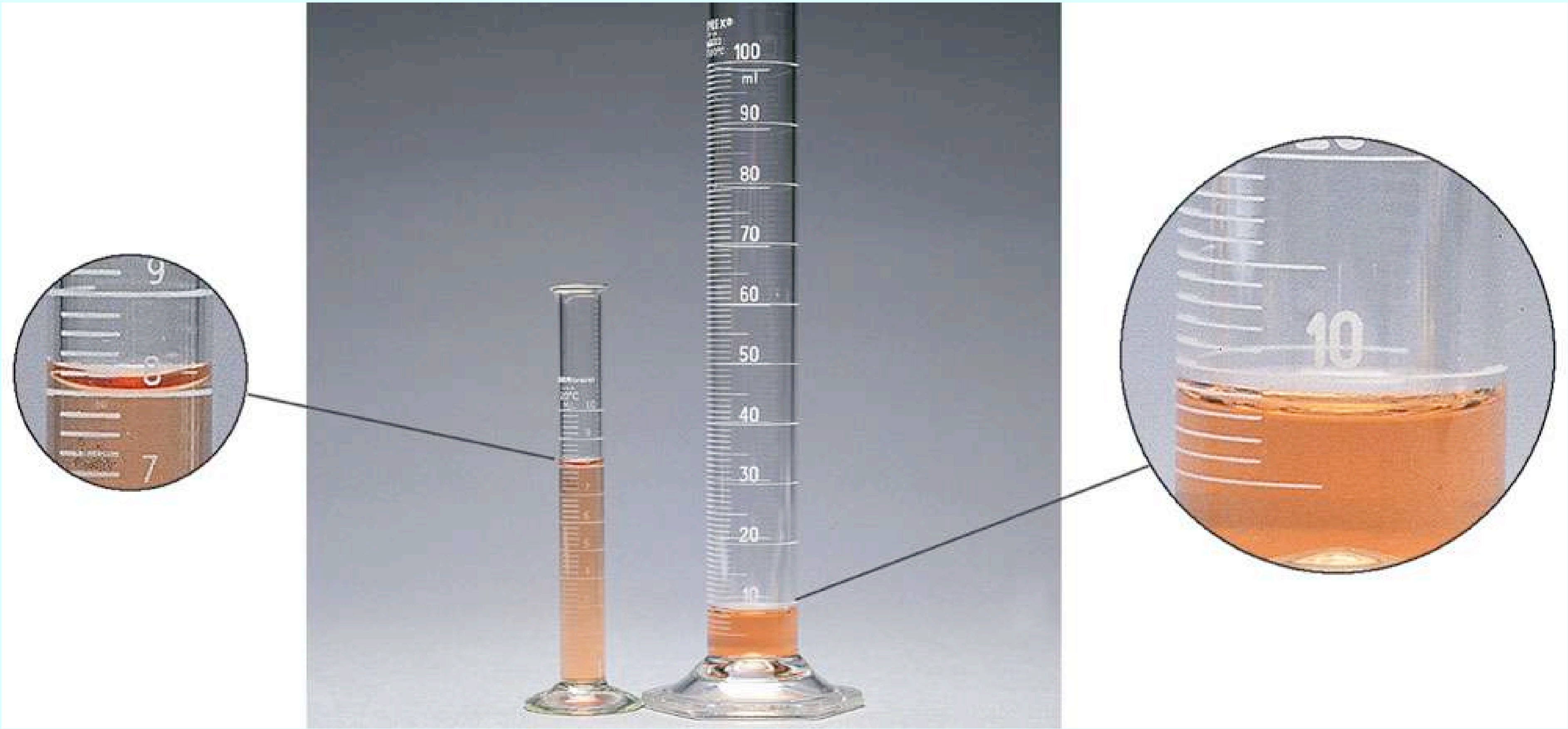
Uncertainties in Measurements

- Significant Figures
  - Every measurement carries uncertainty
  - All measurements must include estimates of uncertainty with them
  - There is an uncertainty of at least one unit in the last digit

Uncertainty in Measuring Volume

- Three volume measurements with their uncertainties
  - Large graduated cylinder, 8 ± 1 mL
  - Small graduate cylinder, 8.0 ± 0.1 mL
  - Pipet or buret, 8.00 ± 0.01 mL
- Text convention
  - Uncertainty of ± in the last digit is assumed but not stated

Figure 1.11 – Uncertainty in Measuring Volume



Example 1.2

EXAMPLE 1.2

Using different balances, three different students weigh the same object. They report the following masses:

(a) 1.611 g (b) 1.60 g (c) 0.001611 kg

How many significant figures does each value have?

STRATEGY

Assume each student reported the mass in such a way that the last number indicates the uncertainty associated with the measurement.

SOLUTION

(a) 1.611 g	4	
(b) 1.60 g	3	The zero after the decimal point is significant. It indicates that the object was weighed to the nearest 0.01 g.
(c) 0.001611 kg	4	The zeros at the left are not significant. They are only there because the mass was expressed in kilograms rather than grams. Note that 1.611 g and 0.001611 kg represent the same mass.

END POINT

If you express these masses in exponential notation as 1.611×10^0 g, 1.60×10^0 g, and 1.611×10^{-3} kg, the number of significant figures becomes obvious.

Significant Figures

- Significant figures are meaningful digits in measurements
  - In 8.00 mL, there are three significant figures
  - In 8.0 mL, there are two significant figures
  - In 8 mL, there is one significant figure

Ambiguity in Significant Figures

- Consider the measurement, 500 g
  - If the measurement was made to the nearest 1 g, all three digits are significant
  - If the measurement was made to the nearest 10 g, only two digits are significant
  - Resolve by using scientific notation
    - 5.00×10^2 g
    - 5.0×10^2 g

Example 1.3

EXAMPLE 1.3

A US Airways flight leaves Philadelphia in the early evening and arrives in Frankfurt 8.05 hours later. The airline distance from Philadelphia to Frankfurt is about 6.6×10^3 km, depending to some extent on the flight path followed. What is the average speed of the plane, in kilometers per hour?

ANALYSIS

Information given:	distance traveled (6.6×10^3 km) time elapsed (8.05 h)
Asked for:	average speed in km/h

STRATEGY

1. Substitute into a formula that relates time and distance.

$$\text{speed} = \frac{\text{distance}}{\text{time}}$$

2. Recall the rules for significant figures.

SOLUTION

average speed	$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{6.6 \times 10^3 \text{ km}}{8.05 \text{ h}} = 819.8757764 \text{ km/h}$
significant figures	numerator: 2; denominator: 3 The answer should have 2 significant figures.
average speed	8.2×10^2 km/h

Rounding

- Rounding off numbers
  - If the digits to be discarded ***less than --500***, ***leave the last digit unchanged***
    - 23.317 and 23.485 round to 23
  - If the digits to be discarded is ***greater than –500, add 1 to the last digit***
    - 23.692 and 23.514 round to 24
  - If the digits to be discarded are equal to ***-- 500, round off so that the last digit is an even number***
    - 23.5 and 24.5 both round to 24

Significant Figures in Addition and Subtraction

- When two numbers are added or subtracted
  - Perform the addition(s) and/or subtraction(s)
  - Count the number of decimal places in ***each number***
  - Round off so that the resulting number has the ***same number of decimal places as the measurement with the greatest uncertainty*** (i.e., the fewer number of decimal places).

Significant Figures in Multiplication and Division

- When measured quantites are multiplied or divided, the number of significant figures in the result is equal to that in the quantity with the fewest significant figures.
  - 2.40 X 2 = 5
  - 1.775 /2.0 = 0.89

Exact Numbers

- Some numbers carry an infinite number of significant figures
- These are exact numbers
- Exact numbers do not change the number of significant figures in a calculation
- The numbers 1.8 and 32 in the conversion between Fahrenheit and Celsius temperature are exact:

$$t_{^\circ F} = 1.8 t_{^\circ C} + 32^\circ$$

More on Exact Numbers

- In some problems in the text, numbers will be spelled out in words
- “Calculate the heat evolved when one kilogram of coal burns”
- Consider these numbers to be exact

Dimensional Analysis

- In many cases throughout your study of chemistry, the units (dimensions) will guide you to the solution of a problem
- Always be sure your answer is reported with both a number and a set of units!

Converting Units

- Conversion factors are used to convert one set of units to another
  - Only the units change
  - Conversion factors are numerically equal to 1
    - $1L = 1000\ cm^3$

$$\frac{1\,L}{1000\ cm^3} = \frac{1000\ cm^3}{1000\ cm^3} = 1$$

Choosing a conversion factor

- Choose a conversion factor that puts the initial units in the denominator
  - The initial units will cancel
  - The final units will appear in the numerator

Example 1.4

EXAMPLE 1.4

A red blood cell has a diameter of 7.5 μm (micrometers). What is the diameter of the cell in inches? (1 inch = 2.54 cm)

ANALYSIS	
Information given:	cell diameter (7.5 μm) bridge conversion (1 in = 2.54 cm)
Information implied:	relation between micrometers and centimeters
Asked for:	7.5 μm in inches
STRATEGY	
Follow the plan:	$\mu m \rightarrow cm \rightarrow inch$
SOLUTION	
7.5 μm in inches	$7.5\ \mu m \times \frac{1 \times 10^{-6}\ m}{1\ \mu m} \times \frac{100\ cm}{1\ m} \times \frac{1\ in}{2.54\ cm} = 3.0 \times 10^{-4}\ in$

Properties of Substances

- There are two fundamental properties of matter
- Chemical properties
    - Accompany chemical change
- Physical properties
    - Occur without chemical change

Gold Metal

Chemical Properties

- Examples
  - Mercury(II) oxide decomposes to mercury and oxygen gas when heated
  - Silver tarnishes on exposure to sulfides in air

Physical Properties

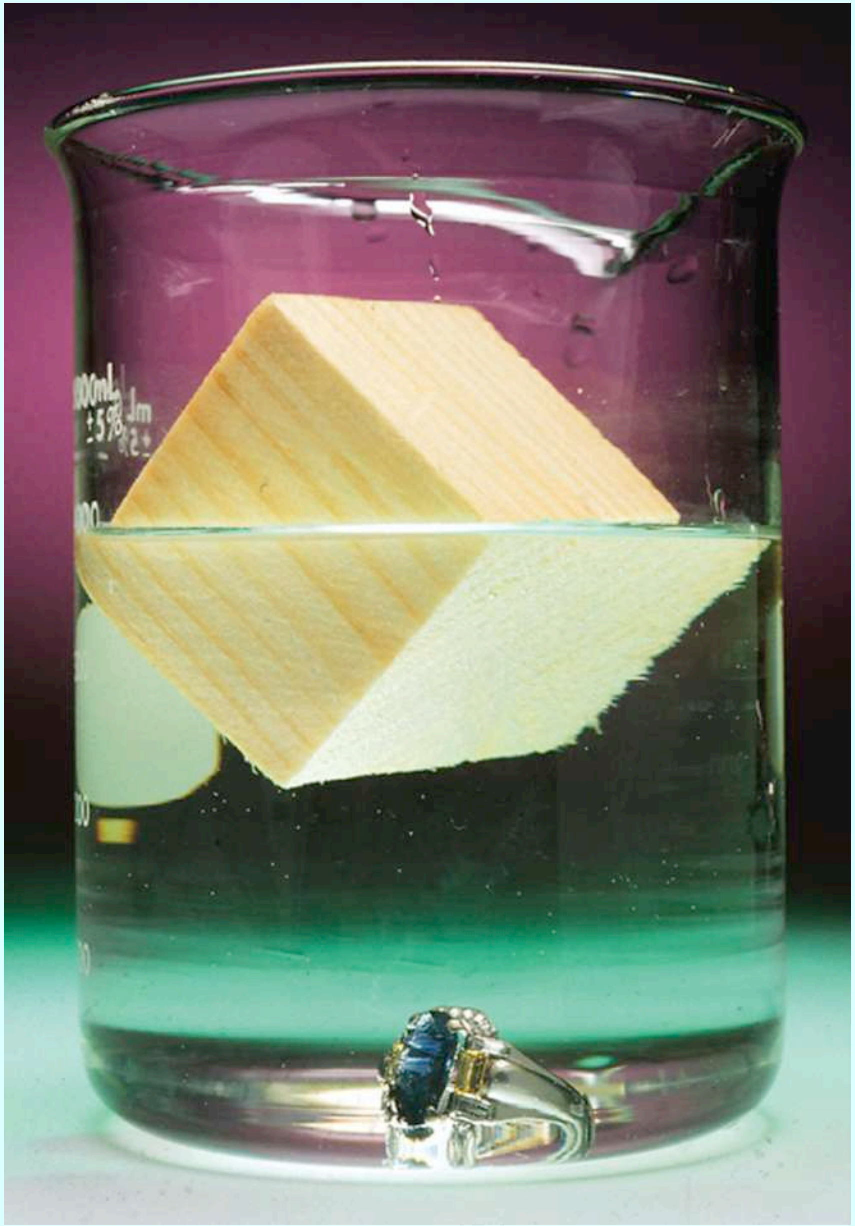
- Melting point
  - Temperature at which a solid changes to a liquid
- Boiling point
  - Temperature at which a liquid changes to a gas
- Both boiling and melting are reversible simply by changing the temperature

Density

- The density of a substance is its mass divided by its volume

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

Figure – Density of Wood and Water



Example 1.5

EXAMPLE 1.5

The beds in your dorm room have extra-long matresses. These mattresses are 80 inches (2 significant figures) long and 39 inches wide. (Regular twin beds are 72 inches long.)

What is the area of the mattress top in m^2? (1 inch = 2.54 cm)

ANALYSIS

Information given:	mattress length (80 in) and width (39 in) bridge conversion (1 inch = 2.54 cm)
Information implied:	centimeter to meter conversion
Asked for:	area in m^2

STRATEGY

1. Recall equation for finding the area of a rectangle: area = length × width
2. Follow the plan: $in^2 \rightarrow cm^2 \rightarrow m^2$

SOLUTION

area in in^2	$80\text{ in} \times 39\text{ in} = 3.12 \times 10^3\text{ in}^2$ (We will round off to correct significant figures at the end.)
area in m^2	$3.12 \times 10^3\text{ in}^2 \times \frac{(2.54)^2\text{ cm}^2}{(1)^2\text{ in}^2} \times \frac{(1)^2\text{ m}^2}{(100)^2\text{ cm}^2} = 2.0\text{ m}^2$

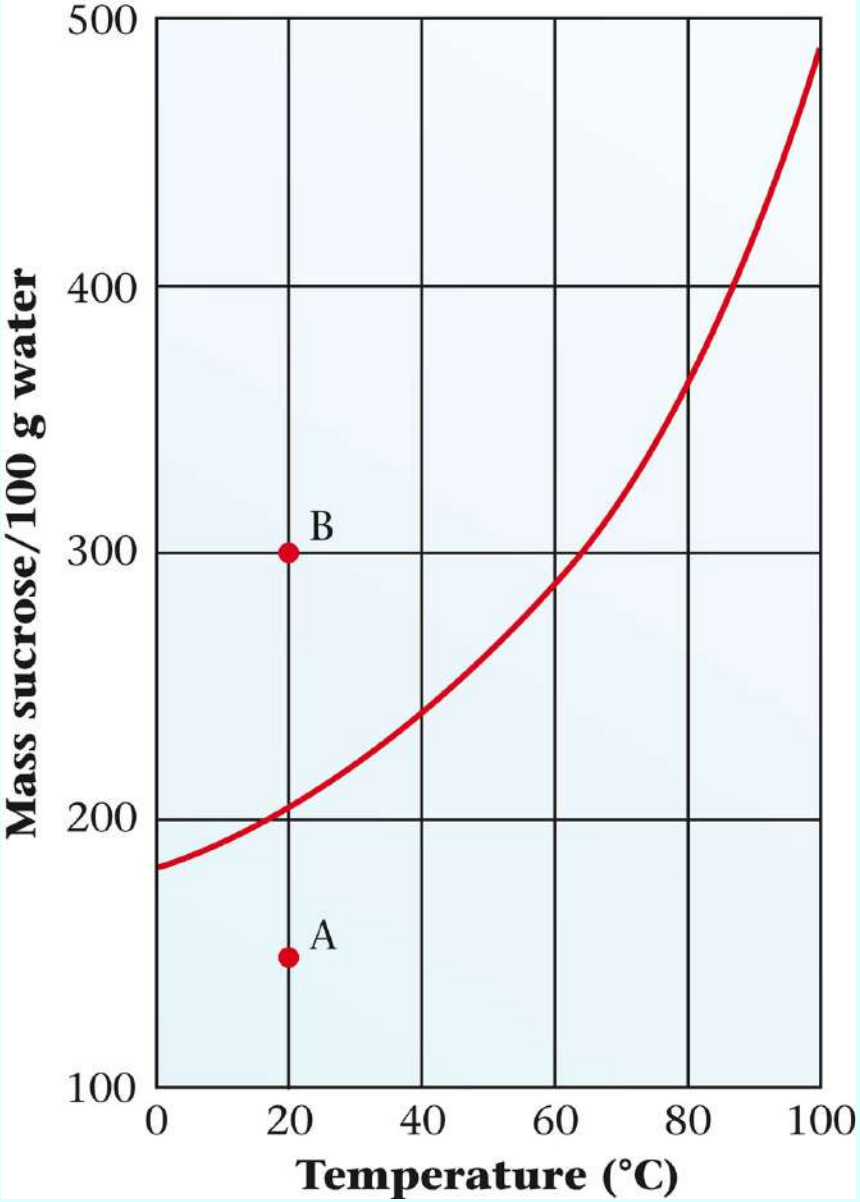
END POINT

There are 36 inches in one yard, so the dimensions of the mattresses are approximately 1 yd wide and 2 yd long or 2 yd^2. A meter is almost equivalent to a yard (see Table 1.3) so the calculated answer is in the same ball park.

Solubility

- The process by which one substance dissolves in another is ordinarily a physical change
- The resulting mixture is a solution
- Solutions may be classified by the relative amount of solute and solvent
  - Saturated: maximum amount of solute
  - Unsaturated: less than maximum amount of solute
  - Supersaturated: more than maximum amount of solute

Figure 1.12 – Solubility and Temperature



Example 1.6

EXAMPLE 1.6

To determine the density of ethyl alcohol, a student pipets a 5.00-mL sample into an empty flask weighing 15.246 g. He weighs the flask with the sample and finds the mass to be 19.171 g. What is the density of the ethyl alcohol?

ANALYSIS

Information given:	mass of empty flask (15.246 g) mass of flask + sample (19.171 g) volume of sample (5.00 mL)
Asked for:	density of the sample

STRATEGY

1. Find the mass of the sample by difference.

$$\text{mass of sample} = (\text{mass of flask} + \text{sample}) - (\text{mass of sample})$$

2. Recall the formula for density.

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

continued

Example 1.6 cont.

SOLUTION	
1. mass of sample	mass of sample = (mass of flask + sample) − (mass of flask) = 19.171 g − 15.246 g = 3.925 g
2. density	$d = \frac{\text{mass}}{V} = \frac{3.925\ \text{g}}{5.00\ \text{mL}} = 0.785\ \text{g/mL}$

END POINT

The density is expressed in 3 significant figures because the volume is measured only to 3 significant figures.

Key Concepts

1. Convert between Fahrenheit, Celsius and Kelvin.
2. Determine the number of significant figures in a measured quantity.
3. Determine the number of significant figures in a calculated quantity.
4. Use conversion factors to change from one quantity to another.
5. Use density to relate mass and volume.
6. Given the solubility, relate mass to volume for a substance.