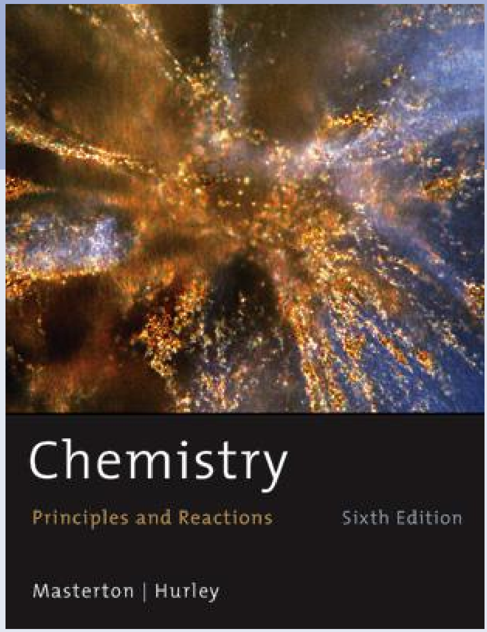

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Chapter 1
Matter and Measurements

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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

1.1 Types of Matter

1.Pure substances

- 2.Mixtures

1.2 Measurements

- 1.Instruments and Units
- 2.Uncertainties in Measurements：Significant Figures

1.3 Properties of Substances

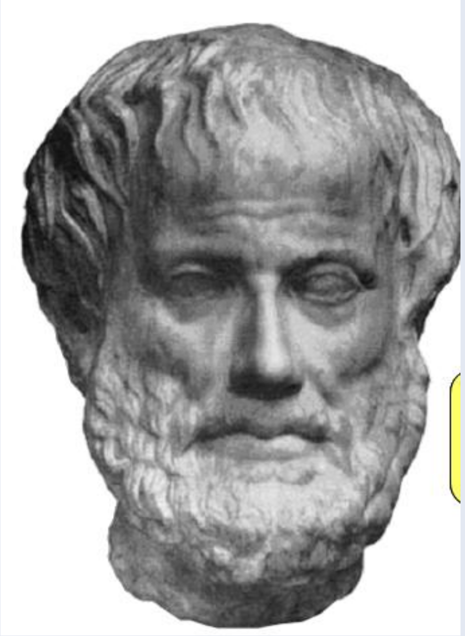
- 1.Chemical properties
- 2.Physical properties
- 3.Density
- 4.Solubility
- 5.Color；Absorption Spectrum

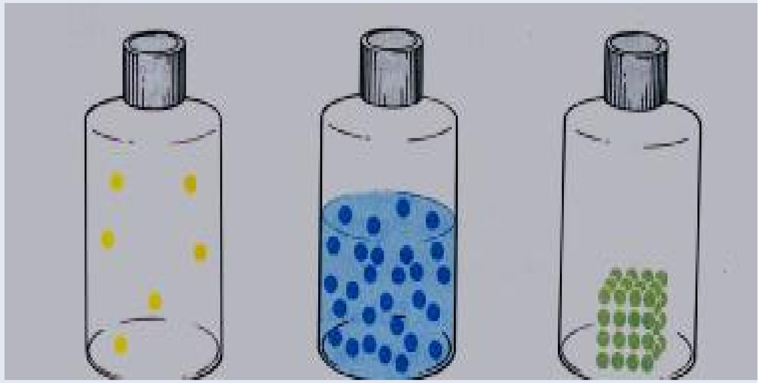
Matter

亞里斯多德

(公元前384-322)

- 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





1.1 Matter

- Pure substances
  - Fixed composition
  - Unique set of properties (mp. bp, density…)
- Mixtures
  - Two or more substances in some combination

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

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, Hg, and oxygen, O
    - 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 1.1 - Classification of Matter

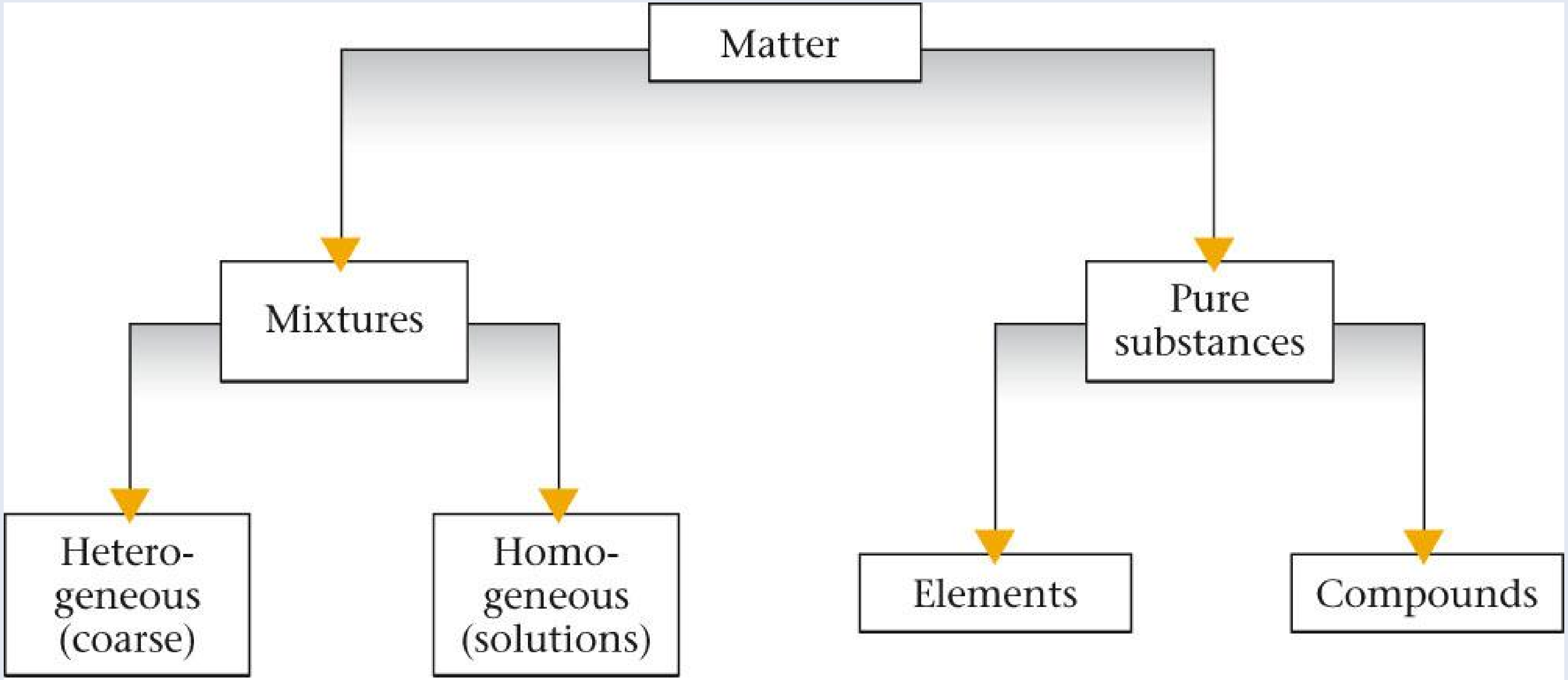
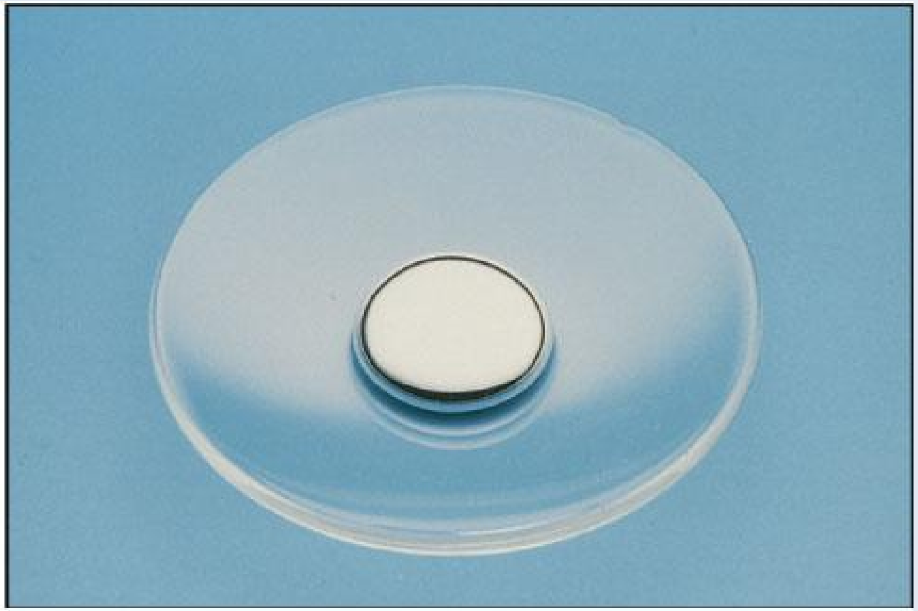


Figure 1.3 – Sodium, Chlorine and Sodium Chloride

(a) (b) (c)

Figure 1.2 – Cinnabar and Mercury

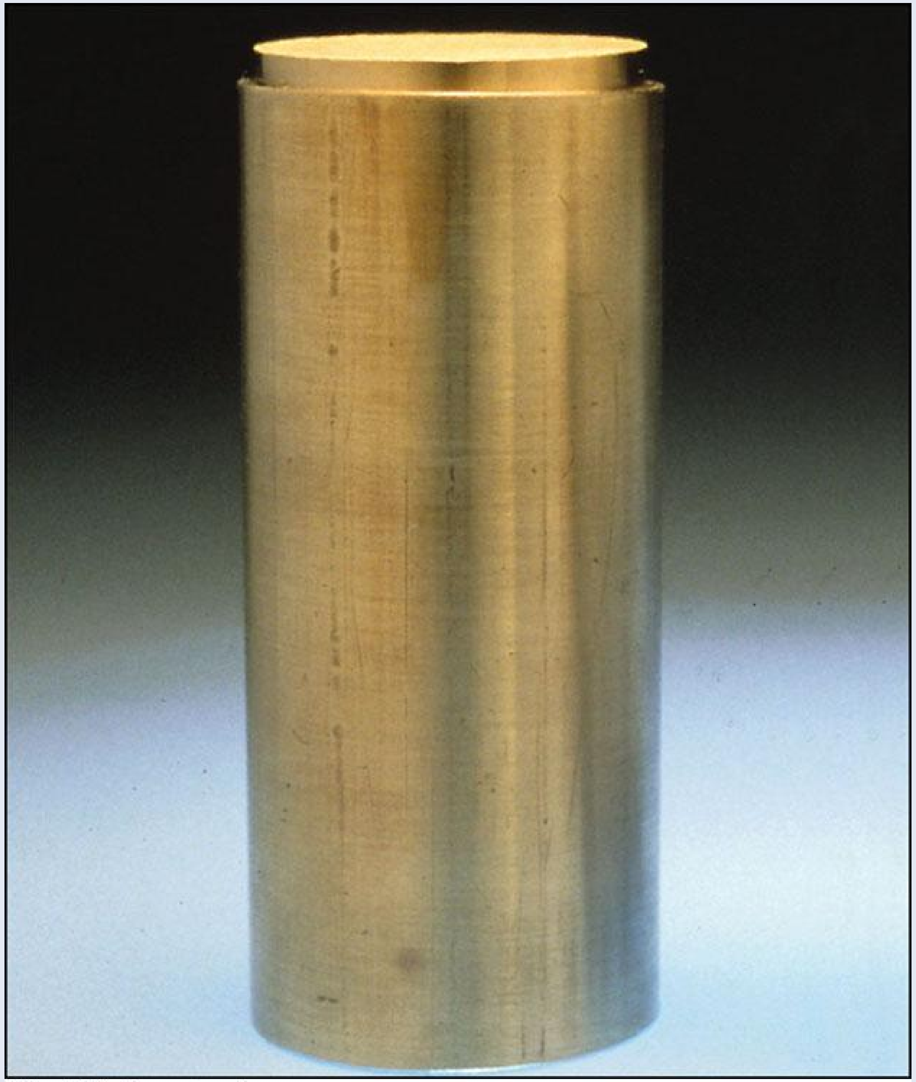
(a)



(b)

Figure 1.4 – Copper Sulfate and Sand

Figure 1.5 – Two Mixtures



Solutions

- Common heterogeneous mixture
- Components
  - Solvent
    - Most commonly a liquid
  - Solute
    - May be solid, liquid or gas
- Seawater
  - Water is the solvent
  - Solutes are 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

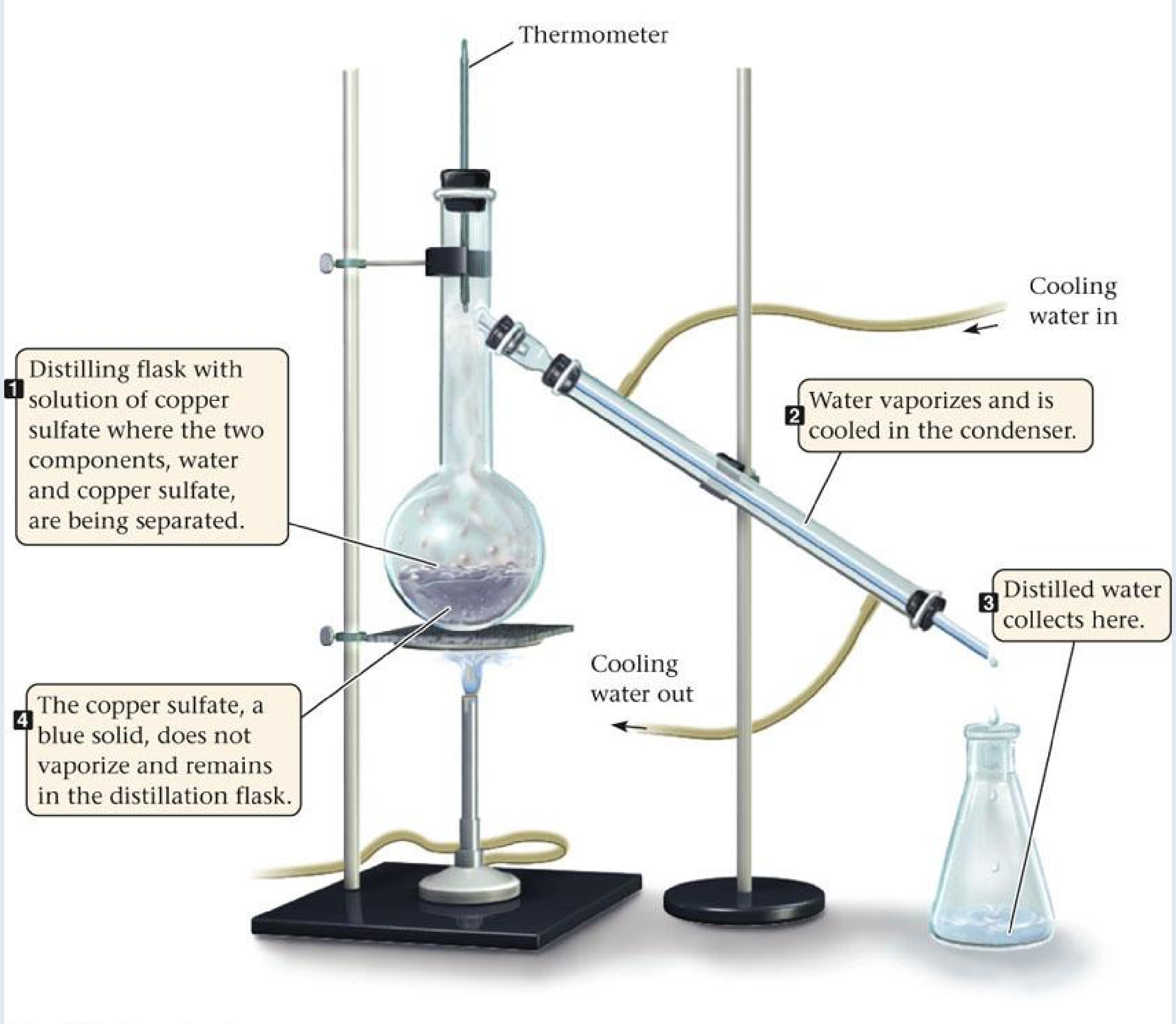
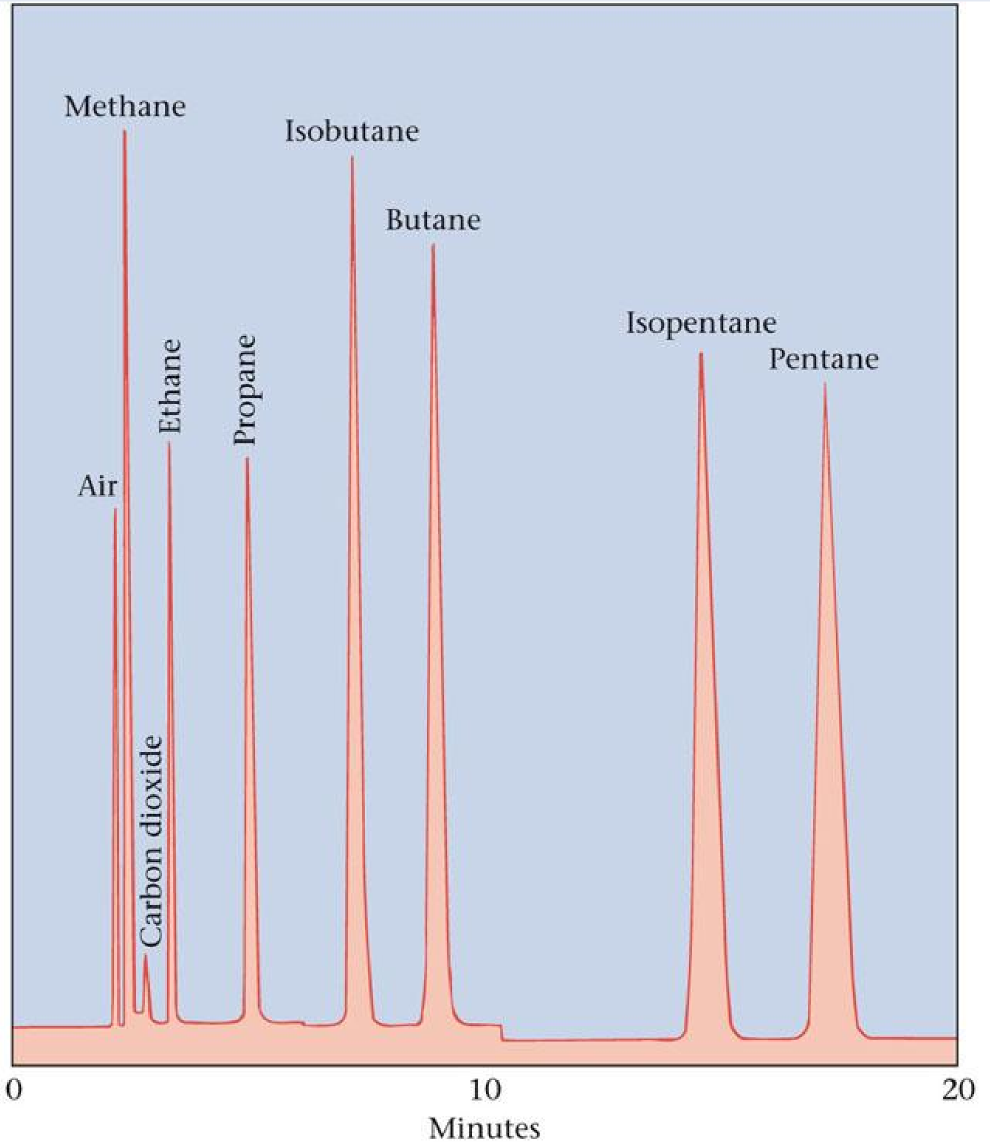


Figure 1.7 – Gas-Liquid Chromatogram



1.2 Measurements

- Quantitation
  - 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 1/299,272,248 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

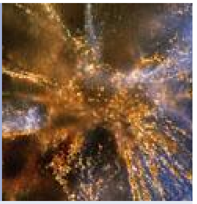
Figure 1.8 – Measuring Volume

Mass

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

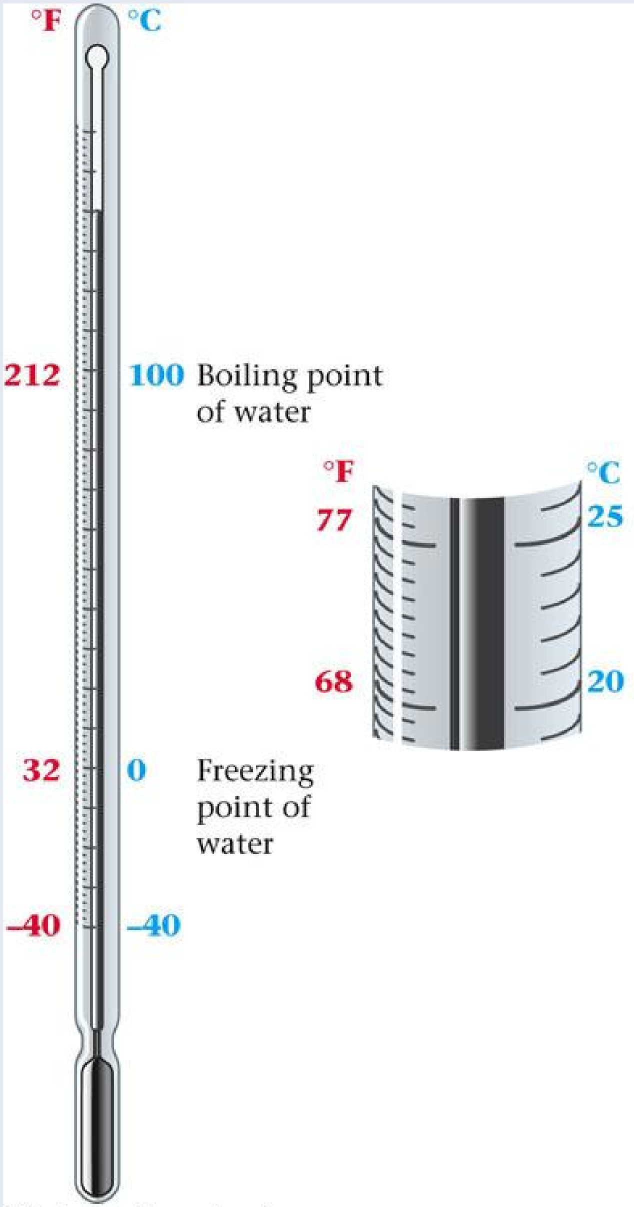
Figure 1.9 – Weighing a Solid

Temperature



- Factor that determines the direction of heat flow
- Temperature is measured indirectly
  - Observing its effect on the properties of a substance
  - Mercury in glass 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 $^{\circ}$ C is 1.8 times larger than each $^{\circ}$ 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 $^{\circ}$C)
  - Unlike the other two scales, no degree sign is used to express temperature in K

Relationships Between Temperature Scales

- Fahrenheit and Celsius

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

- Celsius and Kelvin

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

Example 1.1

Example 1.1 Mercury thermometers are being phased out because of the toxicity of mercury vapor. A common replacement for mercury is the organic liquid isoamyl benzoate, which boils at 262°C. What is its boiling point in °F? K?

Strategy Use Equations 1.1 and 1.2 above. Solve for the desired quantity, $t_{°F}$ in the first case, T_K in the second.

SOLUTION

$$t_{°F} = 1.8(262°) + 32° = \boxed{504°F}$$

$$T_K = 262 + 273.15 = \boxed{535\ K}$$

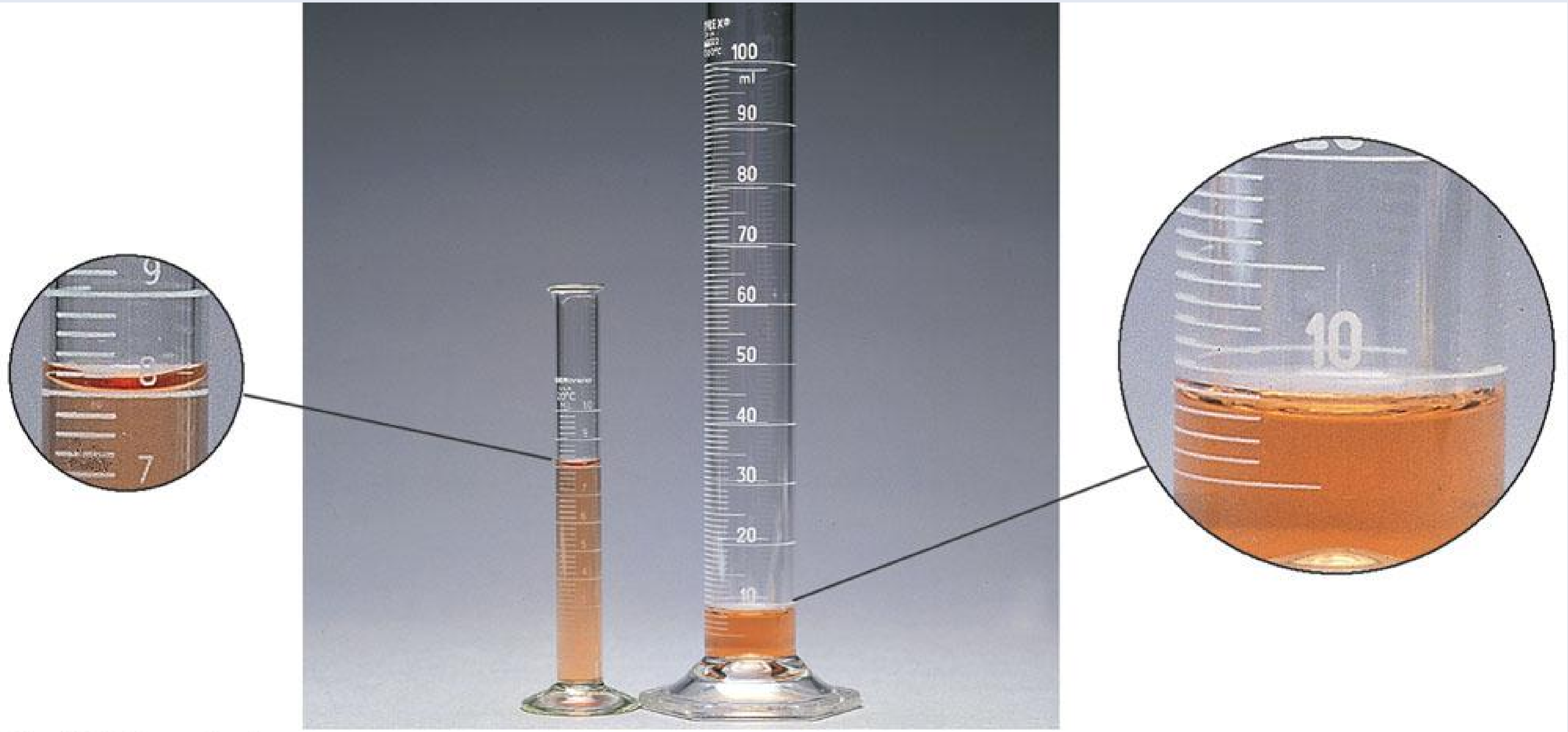
Uncertainties in Measurements

- 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 $\pm$ 1 mL
  - Small graduate cylinder, 8.0 $\pm$ 0.1 mL
  - Pipet or buret, 8.00 $\pm$ 0.01 mL
- Text convention
  - Uncertainty of $\pm$ 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 as to indicate the uncertainty associated with the measurement. Then follow a common sense approach.

SOLUTION

(a) 4

(b) 3 The zero after the decimal point is significant. It indicates that the object was weighed to the nearest 0.01 g.

(c) 4 The zeros at the left are not significant. They are there only because the mass was expressed in kilograms rather than grams. Note that 1.611 g and 0.001611 kg represent the same mass.

Reality Check 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: 4 in (a), 3 in (b), and 4 in (c).

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?

Strategy Calculate the average speed by taking the quotient

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

Count the number of significant figures in the numerator and in the denominator; the smaller of these two numbers is the number of significant figures in the quotient.

SOLUTION The average speed will appear on your calculator as

$$\frac{6.6 \times 10^3 \text{ km}}{8.05 \text{ h}} = 819.8757764 \text{ km/h}$$

There are three significant figures in the denominator and two in the numerator. The answer should have two significant figures; round off the average speed to be

$$8.2 \times 10^2 \text{ km/h.}$$

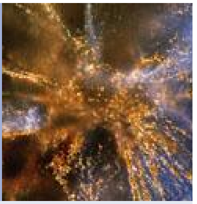
Rounding

- Rounding off numbers
  - If the first digit to be discarded is ***5 or greater***, ***round up***
  - If the first digit to be discarded is ***4 or smaller***, ***round down***

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 multiplying or dividing two numbers, the result is rounded to the number of significant figures in the less (or least in the case of three or more) measurements
- 2.40 X 2 = 5

Example

	Mass	Uncertainty	
Instant coffee	10.21 g	± 0.01 g	2 decimal places
Sugar	0.2 g	± 0.1 g	1 decimal place
Water	256 g	± 1 g	0 decimal places
Total mass	266 g		

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

Table 1.3 – Length, Volume and Mass Units

Table 1.3 Relations Between Length, Volume, and Mass Units

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Example 1.4

Example 1.4 Graded

A gasoline station in Manila, Philippines, charges 37.57 pesos per liter for super unleaded gasoline at a time when one U.S. dollar (USD) buys 47.15 Philippine pesos (PHP). The car you are driving has a capacity of 14.00 U.S. gallons, and gets 24 miles per gallon.

* (a) What is the cost of super unleaded gasoline in Manila in U.S. dollars per gallon?
** (b) How much would a complete tankful of super unleaded for your car cost in dollars?
*** (c) Suppose you have only 1255 pesos, and the car's tank is almost empty. How many kilometers can you expect to drive if you spend all your money on super unleaded gasoline?

SOLUTION

(a) Use Table 1.3 and the dollar-peso conversion rate. Start with 37.57 PHP/L

$$\frac{37.57\ \text{PHP}}{\text{L}} \times \frac{1\ \text{USD}}{47.15\ \text{PHP}} \times \frac{1\ \text{L}}{1.057\ \text{qt}} \times \frac{4\ \text{qt}}{1\ \text{gal}} = \boxed{3.015\ \text{USD/gal}}$$

(b) With the price in dollars per gallon from (a), this is a simple conversion:

$$14.00\ \text{gal} \times \frac{3.015\ \text{USD}}{1\ \text{gal}} = \boxed{42.22\ \text{USD}}$$

(c) First find how much gasoline (in gallons) you can buy. Then use the car's miles/gallon rating to find the number of miles you can travel. Finally, convert miles to kilometers (Table 1.3).

$$1255\ \text{PHP} \times \frac{1\ \text{USD}}{47.15\ \text{PHP}} \times \frac{1\ \text{gal}}{3.015\ \text{USD}} \times \frac{24\ \text{miles}}{1\ \text{gal}} \times \frac{1.609\ \text{km}}{1\ \text{mile}} = \boxed{3.4 \times 10^2\ \text{km}}$$

Reality Check On the day of the above pricing in Manila, super unleaded gasoline in Storrs, Connecticut, sold for $2.49 per gallon. However, 1255 PHP (≈$27) was about one day's wage for a teacher in the Philippines at that time.

Ex1.4： A certain filling station in Paris sells gas for 1.18euros per liter （Assume that one dollar＝1.15euro）

（a） How much will a thankful of gas （43L） cost in erros？

（b） Suppose you have 9L of gas in your tank. If you ask the attendant to filer-up How much should you expect to pay in dollars？

（c） What is the cost of the gasoline in dollars per gallon？

Sol:

- a). 43 x 1.18 = 50.74 = 51 euros
- b). (43 –9) x 1.18 ／1.15 = 34.89 dollars = 35 dollars
- c). 1L = 1.057 qt; 1qt =0.25 gallon
- $$\frac{1.18euro}{1L}\times\frac{1}{1.15euro}\times\frac{1L}{1.057qt}\times\frac{4qt}{1gal}=3.88/gal$$

1.3 Properties of Substances

- There are two fundamental types of property
  - Chemical properties
    - Require chemical change
  - Physical properties
    - No chemical change is required

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

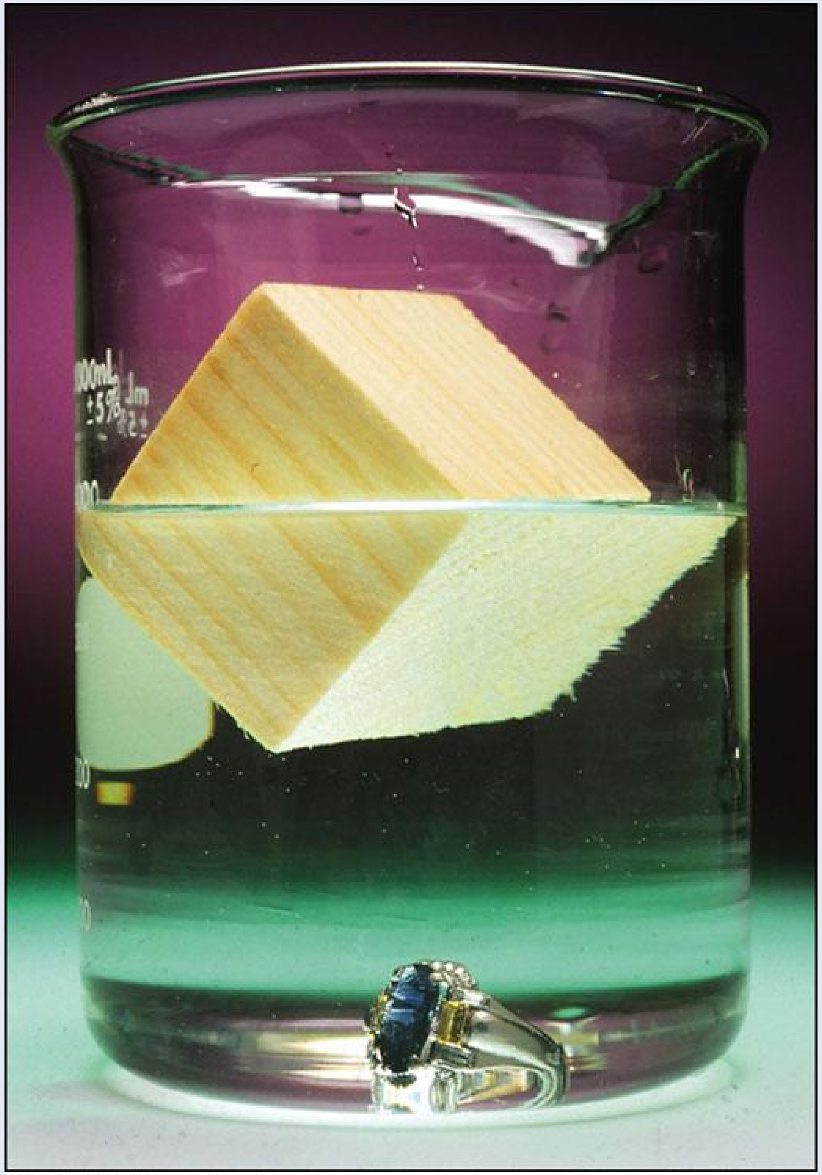
- Melting point
  - Temperature at which a solid changes to a liquid
- Boiling point
  - Temperature at which a liquid changes to a gas
- Both boiling and melting are reversible simply by changing the temperature

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 Palladium (Pd) is an element with properties similar to those of platinum. It is useful in eliminating harmful emissions produced by internal combustion engines. Two students were given identical cylindrical "palladium" bars with the following data:

$$\text{mass} = 96.03 \text{ g, length} = 10.7 \text{ cm, diameter} = 9.82 \text{ mm, density} = 12.02 \text{ g/cm}^3$$

(a) Student X was asked to determine whether her bar was made of pure palladium.
(b) Student Y was asked to determine how many grams of ethyl alcohol ($d = 0.789$ g/cm^3) his bar would displace.

Show the calculations that Students X and Y would do for the assigned tasks.

Strategy

(a) A pure bar of Pd should have density 12.02 g/cm^3. To determine the density of her bar, Student X would need its mass (given) and its volume. She could calculate the volume from the formula

$$\text{volume} = \pi(\text{radius})^2(\text{length})$$

(b) The bar will displace its volume of ethyl alcohol. So Student Y can use the density of alcohol to determine the mass of his bar.

Example 1.5 (cont'd)

SOLUTION

(a)

$$\text{mass of bar} = 96.03\text{ g}$$

$$\text{volume of bar} = \pi\left(\frac{9.82\text{ mm}}{2} \times \frac{1\text{ cm}}{10\text{ mm}}\right)^2 \times (10.7\text{ cm}) = 8.10\text{ cm}^3$$

$$\text{density of bar} = \frac{96.03\text{ g}}{8.10\text{ cm}^3} = 11.8\text{ g/cm}^3$$

Hence, the bar is not pure palladium.

(b) From (a), the volume of the bar is 8.10 cm^3. That is the volume of ethyl alcohol displaced by Student Y's bar. Thus,

$$\text{mass of ethyl alcohol displaced} = (\text{density of ethyl alcohol}) \times (\text{volume of ethyl alcohol displaced})$$

$$= \frac{0.789\text{ g}}{\text{cm}^3} \times 8.10\text{ cm}^3 = 6.39\text{ g}$$

Reality Check Since the density of palladium (12.02 g/cm^3) is larger than that of the metal bars, it would seem that they are slightly contaminated by a less dense metal.

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 1.13 – Sugar Crystals

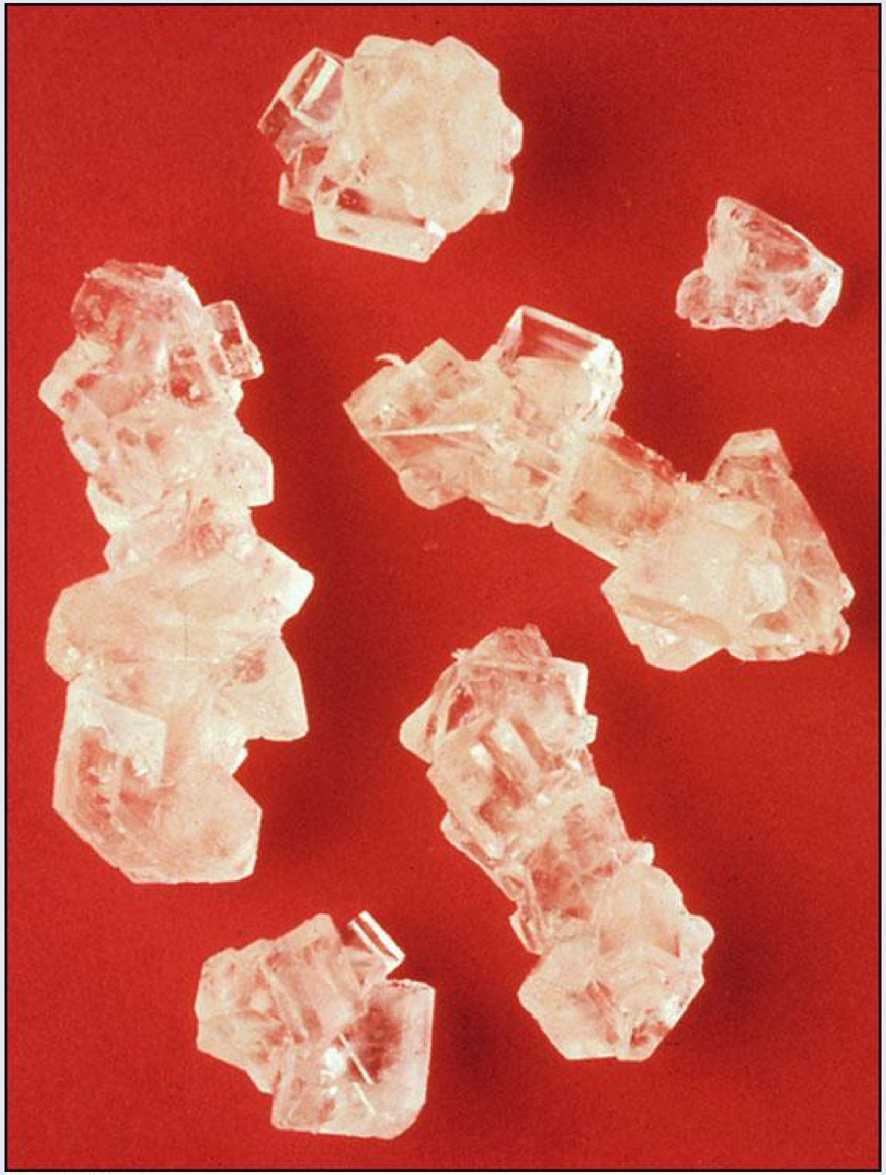
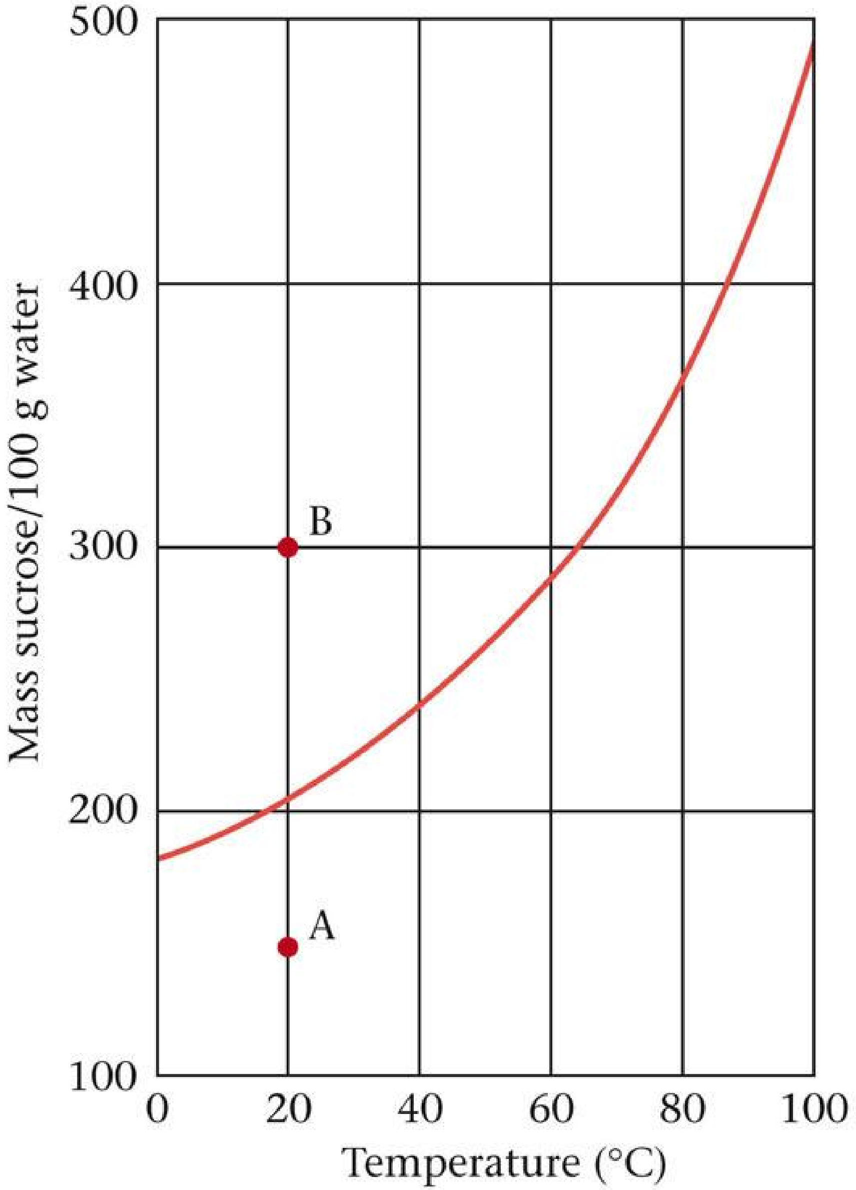


Figure 1.12 – Solubility and Temperature



Example 1.6

Example 1.6 Graded

Sucrose is the chemical name for the sugar we consume. Its solubility at 20°C is 204 g/100 g water, and at 100°C is 487 g/100 g water. A solution is prepared by mixing 139 g of sugar in 33.0 g of water at 100°C.

- * (a) What is the minimum amount of water required to dissolve the sugar at 100°C?
- ** (b) What is the maximum amount of sugar that can be dissolved in the water at 100°C?
- *** (c) The solution is cooled to 20°C. How much sugar (if any) will crystallize out?
- **** (d) How much more water is required to dissolve all the sugar at 20°C?

Strategy The solubility at a particular temperature gives a relationship between grams of sugar and grams of water. This in turn leads to a conversion factor to calculate either the mass of sugar or that of water.

Color

- Some substances can be identified by color
- Color arises from the absorption and transmission of specific wavelengths of light
  - Copper sulfate is blue
  - Potassium permanganate is deep violet

Visible Light

- Visible light ranges from 400 to 700 nm
  - Below 400 nm is the ultraviolet
    - Ultraviolet light leads to sunburn
  - Above 700 is the infrared
    - Heat
    - Absorption of infrared light leads to warming up
    - Global warming and carbon dioxide

Table 1.4 – Color and Wavelength

Table 1.4 Relation Between Color and Wavelength

Wavelength (nanometers)	Color Absorbed	Color Transmitted
<400 nm	Ultraviolet	Colorless
400–450 nm	Violet	Red, orange, yellow
450–500 nm	Blue	
500–550 nm	Green	Purple
550–580 nm	Yellow	
580–650 nm	Orange	Blue, green
650–700 nm	Red	
>700 nm	Infrared	Colorless

Figure 1.14-1.15

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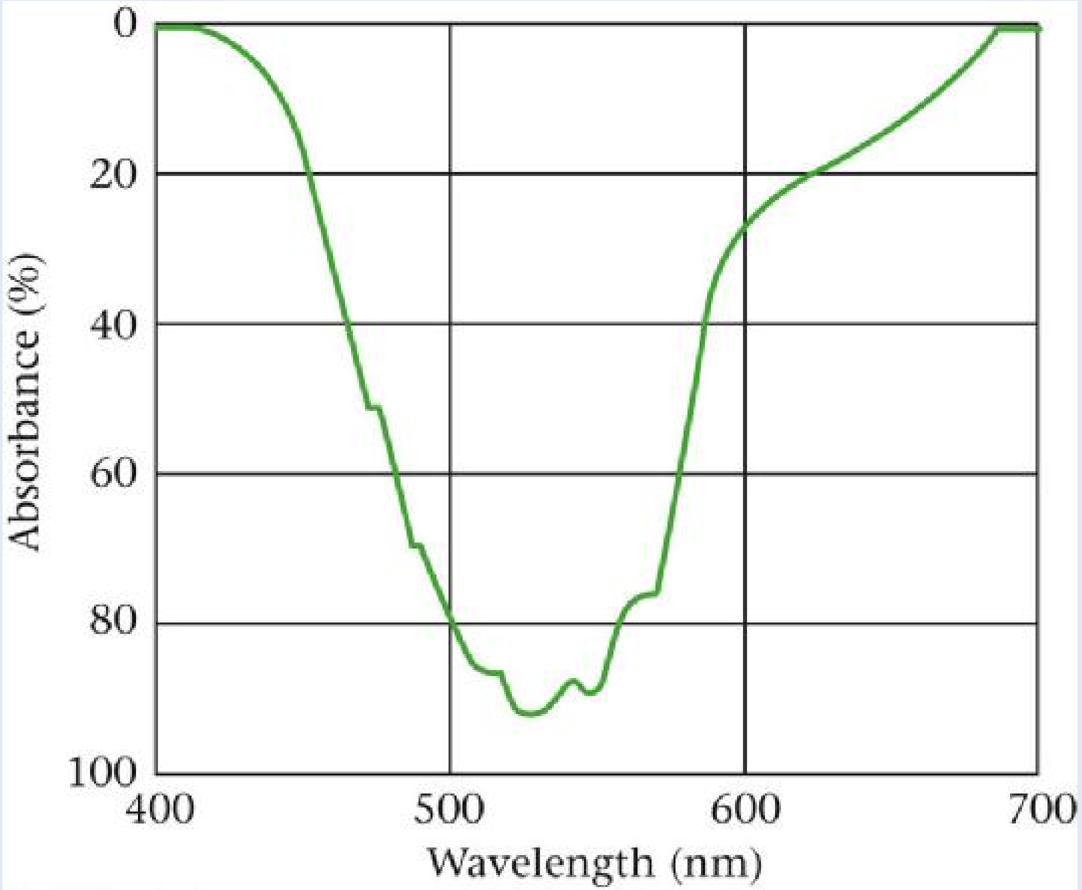

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Key Concepts

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.