化学

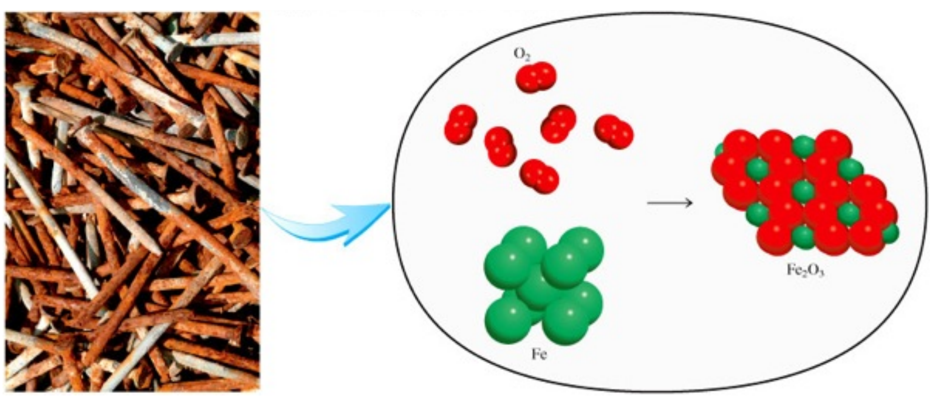
Introduction

Chapter 1

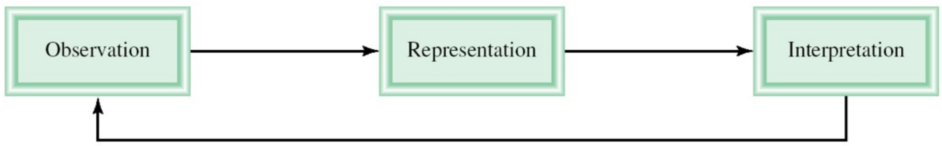
The Study of Chemistry

Macroscopic

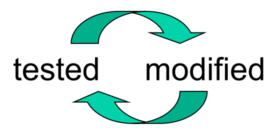
Microscopic



The ***scientific method*** is a systematic approach to research



A ***hypothesis*** is a tentative explanation for a set of observations



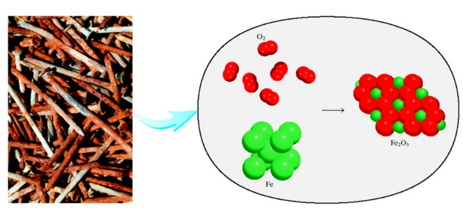
3

A ***law*** is a concise statement of a relationship between phenomena that is always the same under the same conditions.

Force = mass x acceleration

A ***theory*** is a unifying principle that explains a body of facts and/or those laws that are based on them.

Atomic Theory



4

Chemistry is the study of matter and the changes it undergoes

Matter is anything that occupies space and has mass.

A ***substance*** is a form of matter that has a definite composition and distinct properties.


liquid nitrogen


gold ingots

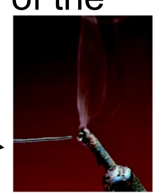

silicon crystals

5

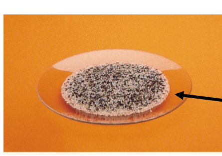
A ***mixture*** is a combination of two or more substances in which the substances retain their distinct identities.

1. ***Homogenous mixture*** – composition of the mixture is the same throughout.

   soft drink, milk, solder



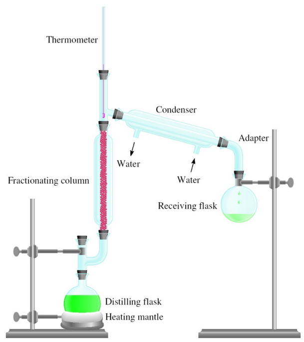
2. ***Heterogeneous mixture*** – composition is not uniform throughout.



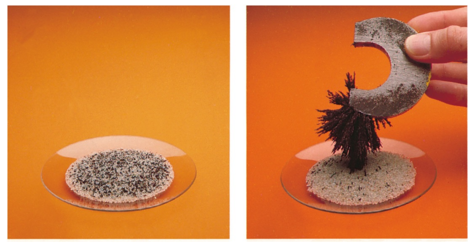
cement,
iron filings in sand

6

Physical means can be used to separate a mixture into its pure components.



distillation



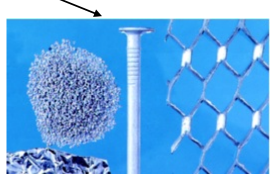
magnet

7

An ***element*** is a substance that cannot be separated into simpler substances by ***chemical means.***

- 117 elements have been identified
  - 82 elements occur naturally on Earth

    gold, aluminum, lead, oxygen, carbon, sulfur



  - 35 elements have been created by scientists

    technetium, americium, seaborgium

8

TABLE 1.1 Some Common Elements and Their Symbols					
Name	**Symbol**	**Name**	**Symbol**	**Name**	**Symbol**
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

9

A ***compound*** is a substance composed of atoms of two or more elements chemically united in fixed proportions.

Compounds can only be separated into their pure components (elements) by ***chemical*** means.

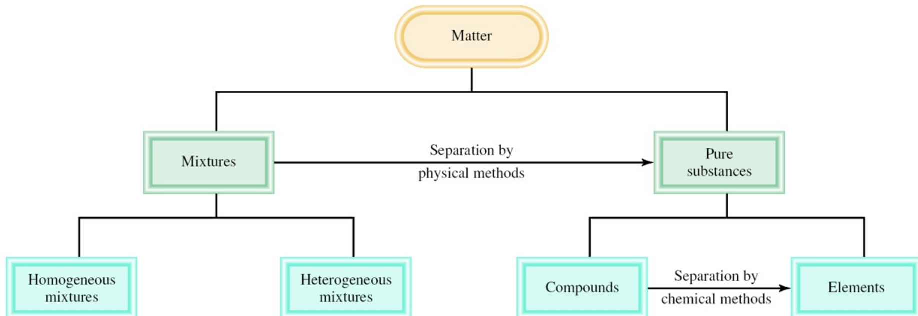
lithium fluoride

quartz

dry ice – carbon dioxide

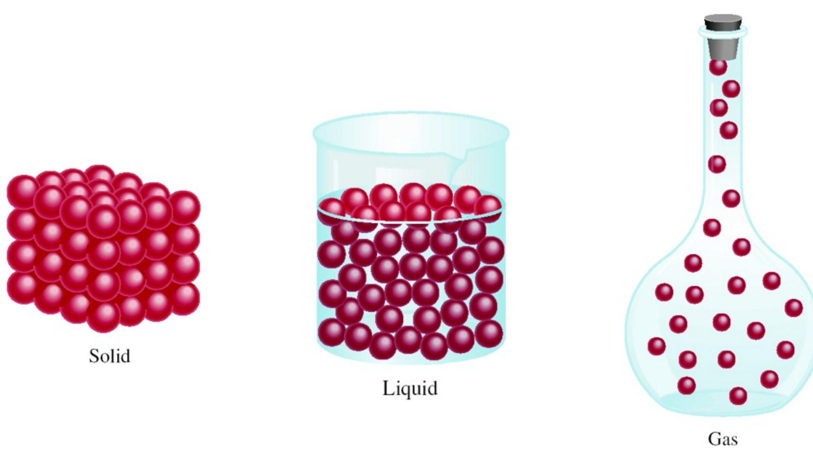
10

Classifications of Matter



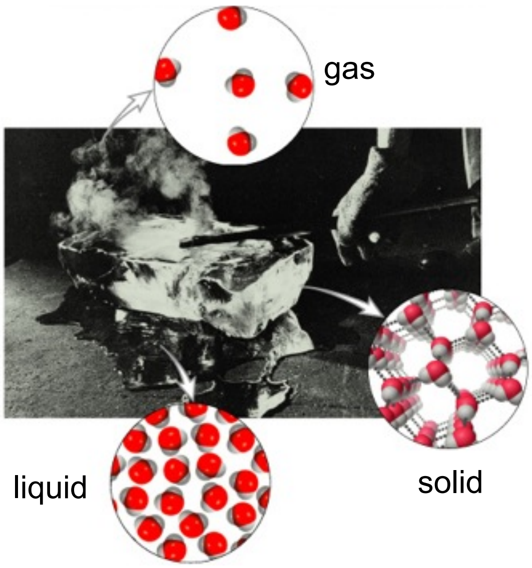
11

A Comparison: The Three States of Matter



12

The Three States of Matter: Effect of a Hot Poker on a Block of Ice



13

Types of Changes

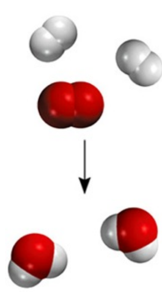
A ***physical change*** does not alter the composition or identity of a substance.

ice melting

sugar dissolving in water

A ***chemical change*** alters the composition or identity of the substance(s) involved.

hydrogen burns in air to form water

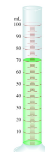


14

Extensive and Intensive Properties

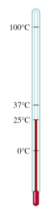
An ***extensive property*** of a material depends upon how much matter is is being considered.

- mass
- length
- volume



An ***intensive property*** of a material does not depend upon how much matter is being considered.

- density
- temperature
- color



15

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

mass – measure of the quantity of matter

SI unit of mass is the ***kilogram*** (kg)

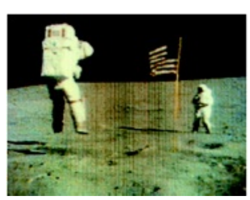
1 kg = 1000 g = 1 x 10^3 g

weight – force that gravity exerts on an object

weight = *c* x mass

on earth, *c* = 1.0

on moon, *c* ~ 0.1



A 1 kg bar will weigh

1 kg on earth

0.1 kg on moon

16

International System of Units (SI)

TABLE 1.2 SI Base Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

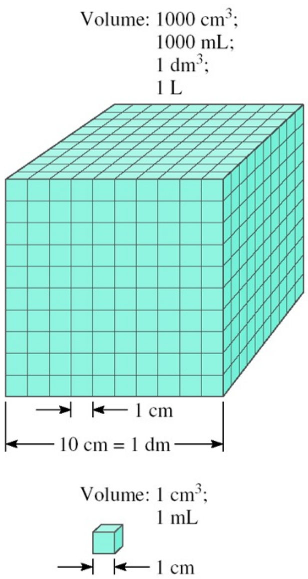
17

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

18

Volume – SI derived unit for volume is cubic meter (m^3)



$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$

$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$

$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$

$1\ mL = 1\ cm^3$

Density – SI derived unit for density is kg/m^3

$1\ g/cm^3 = 1\ g/mL = 1000\ kg/m^3$

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{m}{V}$$

A piece of platinum metal with a density of 21.5 g/ cm^3 has a volume of 4.49 cm^3. What is its mass?

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

$$m = d \times V = 21.5\ g/cm^3 \times 4.49\ cm^3 = 96.5\ g$$

20

TABLE 1.4

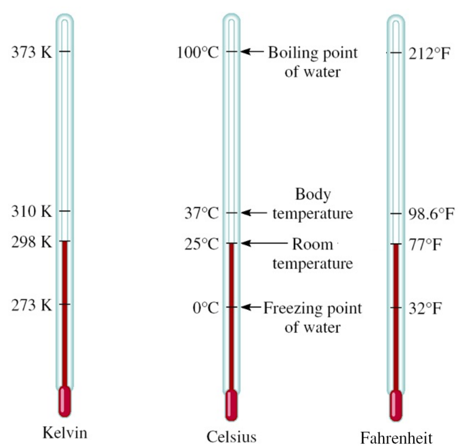
Densities of Some Substances at 25°C

Substance	Density (g/cm^3)
Air*	0.001
Ethanol	0.79
Water	1.00
Mercury	13.6
Table salt	2.2
Iron	7.9
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.

21

A Comparison of Temperature Scales



K = ˚C + 273.15

273 K = 0 ˚C
373 K = 100 ˚C

$$˚F = \frac{9}{5} \times ˚C + 32$$

32 ^{0}F = 0 ˚C
212 ^{0}F = 100 ˚C

22

Convert 172.9 ˚F to degrees Celsius.

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

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

$$\frac{5}{9} \times ({}^\circ F - 32) = {}^\circ C$$

$$^\circ C = \frac{5}{9} \times ({}^\circ F - 32)$$

$$^\circ C = \frac{5}{9} \times (172.9 - 32) = 78.3$$

23

Scientific Notation

The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

6.022×10^{23}

The mass of a single carbon atom in grams:

0.0000000000000000000000199

1.99×10^{-23}

$N \times 10^{n}$

N is a number between 1 and 10

n is a positive or negative integer

24

Scientific Notation

568.762

← move decimal left

n > 0

568.762 = 5.68762 x 10^{2}

0.00000772

→ move decimal right

n < 0

0.00000772 = 7.72 x 10^{-6}

Addition or Subtraction

1. Write each quantity with the same exponent *n*
2. Combine N_1 and N_2
3. The exponent, *n*, remains the same

4.31 x 10^{4} + 3.9 x 10^{3} =

4.31 x 10^{4} + 0.39 x 10^{4} =

4.70 x 10^{4}

25

Scientific Notation

Multiplication

1. Multiply N_1 and N_2
2. Add exponents n_1 and n_2

(4.0 x 10^{-5}) x (7.0 x 10^{3}) =

(4.0 x 7.0) x (10^{-5+3}) =

28 x 10^{-2} =

2.8 x 10^{-1}

Division

1. Divide N_1 and N_2
2. Subtract exponents n_1 and n_2

8.5 x 10^{4} ÷ 5.0 x 10^{9} =

(8.5 ÷ 5.0) x 10^{4-9} =

1.7 x 10^{-5}

26

Significant Figures

- Any digit that is not zero is significant

  1.234 kg 4 significant figures

- Zeros between nonzero digits are significant

  606 m 3 significant figures

- Zeros to the left of the first nonzero digit are **not** significant

  0.08 L 1 significant figure

- If a number is greater than 1, then all zeros to the right of the decimal point are significant

  2.0 mg 2 significant figures

- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

  0.00420 g 3 significant figures

27

How many significant figures are in each of the following measurements?

24 mL	2 significant figures
3001 g	4 significant figures
0.0320 m^3	3 significant figures
6.4 x 10^4 molecules	2 significant figures
560 kg	2 significant figures

28

Significant Figures

Addition or Subtraction

The answer cannot have more digits to the right of the decimal point than any of the original numbers.

```
  89.332
  +1.1    ←—— one significant figure after decimal point
 ———————
  90.432  ←—— round off to 90.4
```

```
   3.70   ←—— two significant figures after decimal point
  -2.9133
 ————————
   0.7867 ←—— round off to 0.79
```

29

Significant Figures

Multiplication or Division

The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures

4.51 x 3.6666 = 16.536366 = 16.5

- 4.51: 3 sig figs
- 16.536366: round to 3 sig figs

6.8 ÷ 112.04 = 0.0606926 = 0.061

- 6.8: 2 sig figs
- 0.0606926: round to 2 sig figs

30

Significant Figures

<u>Exact Numbers</u>

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures

The average of three measured lengths: 6.64, 6.68 and 6.70?

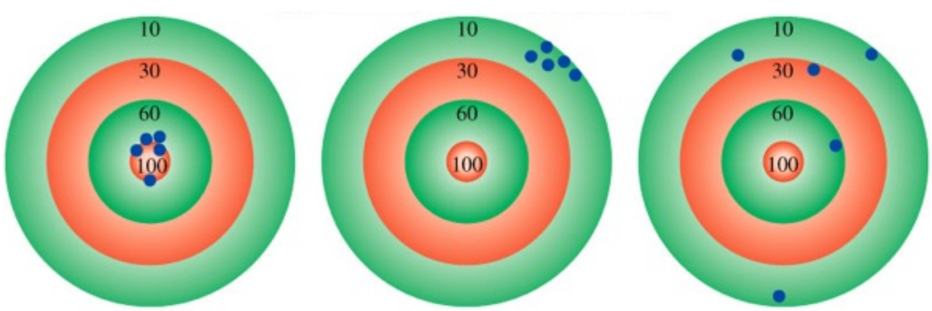
$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 = \cancel{7}$$

Because 3 is an ***exact number***

31

Accuracy – how close a measurement is to the *true* value

Precision – how close a set of measurements are to each other



accurate & precise	precise but not accurate	not accurate & not precise

32

Dimensional Analysis Method of Solving Problems

1. Determine which unit conversion factor(s) are needed
2. Carry units through calculation
3. If all units cancel except for the ***desired unit(s)***, then the problem was solved correctly.

given quantity x conversion factor = desired quantity

$$\text{given unit} \times \frac{\text{desired unit}}{\text{given unit}} = \text{desired unit}$$

33

Dimensional Analysis Method of Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

$$1.63\ \text{L} \times \frac{1000\ \text{mL}}{1\ \text{L}} = 1630\ \text{mL}$$

$$1.63\ \text{L} \times \frac{1\text{L}}{1000\ \text{mL}} = 0.001630\ \frac{\text{L}^2}{\text{mL}}$$

34

The speed of sound in air is about 343 m/s. What is this speed in miles per hour?

conversion units

meters to miles

seconds to hours

1 mi = 1609 m　　1 min = 60 s　　1 hour = 60 min

$$343 \frac{\cancel{m}}{\cancel{s}} \times \frac{1 \text{ mi}}{1609 \cancel{m}} \times \frac{60 \cancel{s}}{1 \cancel{min}} \times \frac{60 \cancel{min}}{1 \text{ hour}} = 767 \frac{\text{mi}}{\text{hour}}$$

35