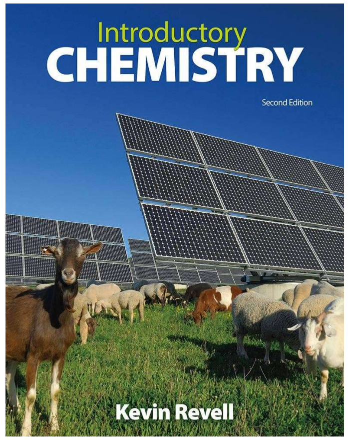
Introductory Chemistry
Chem 103

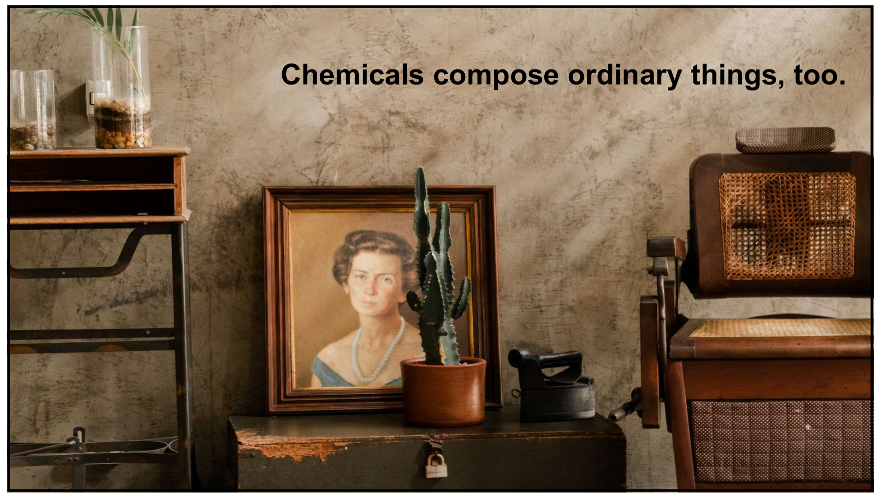
Chapter 1 – Foundations

Lecture Slides



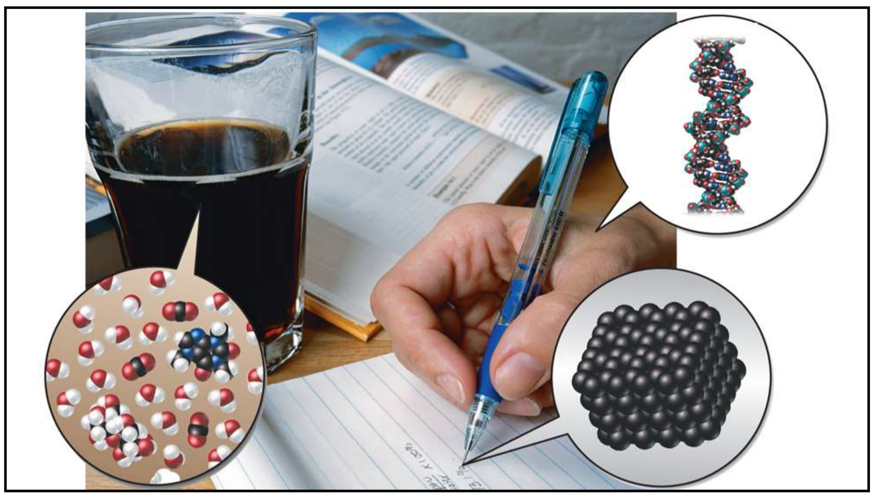
People often have a very narrow view of chemicals, thinking of them only as dangerous poisons or pollutants.

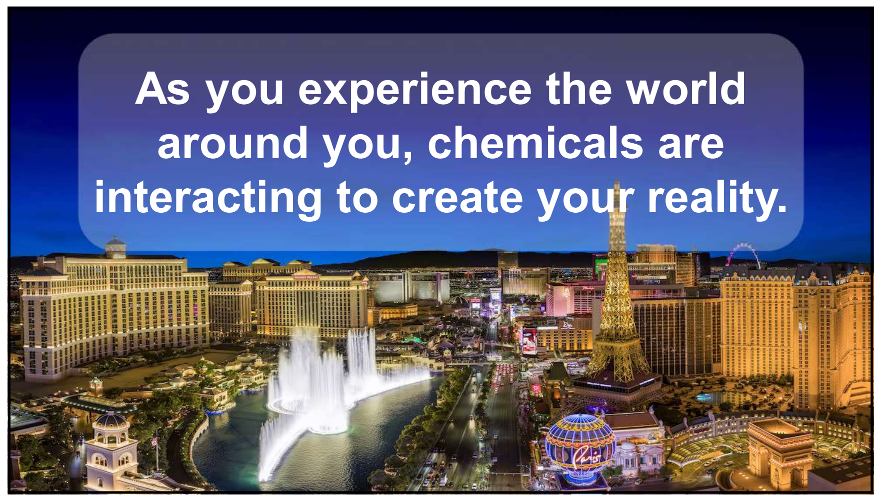
Are chemicals a good or bad thing?



Chemicals compose ordinary things, too.

Chemistry – Part of Everything You Do





CLASS ACTIVITY

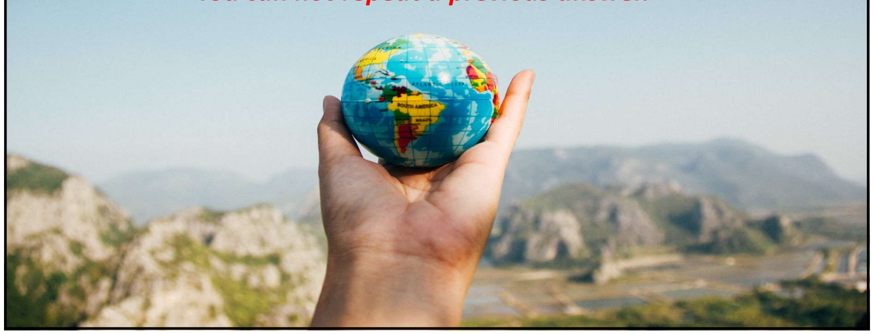
Describing Matter

Matter anything that has mass and takes up volume

Courtesy David Revell

Provide an example of chemistry in your everyday life.

You can not repeat a previous answer.



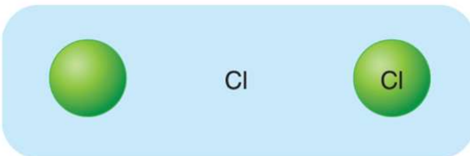
Composition and Structure

Composition	Structure
What something is made of	*What something is made of* *and* *How the components are arranged*

Left to Right: YinYang/iStock/Getty Images; Joel Blit/Shutterstock; David Lee/Shutterstock

Pure Substances: Elements and Compounds

Atom: the fundamental unit of matter

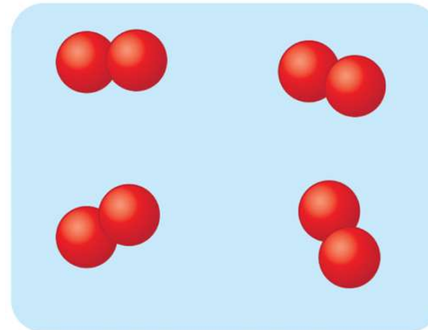


Revell, *Introductory Chemistry*, 2e, © 2021 W. H. Freeman and Company

Left and Right: Duncan Chard/Bloomberg via Getty Images; sumire8/Shutterstock

Element: made of only one type of atom

Diatomic Molecules

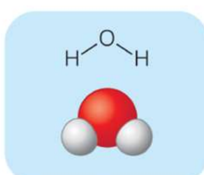


Some elements, such as hydrogen, nitrogen, and oxygen also exist as diatomic (two atom) molecules. For example, this image shows four molecules of oxygen. Each molecule contains two oxygen atoms bound together.

Compounds and Molecules

Compounds: composed of more than one element, bound in fixed ratios

Molecules: groups of atoms that bind tightly together, and behave as a single unit



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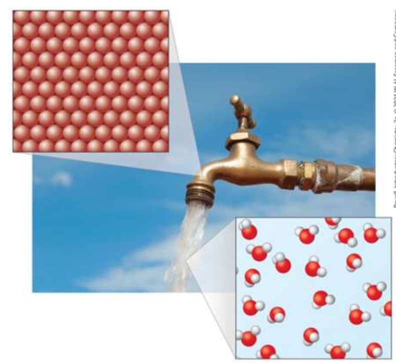
Composition of Materials

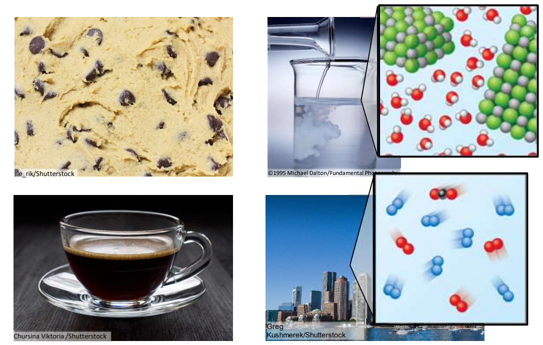


Mixtures

Contain more than one substance, not bound in a fixed ratio.

Bronze:
A mixture of copper and tin

Other mixtures…



Homogeneous and Heterogeneous Mixtures

Homogeneous mixtures – components mix evenly.

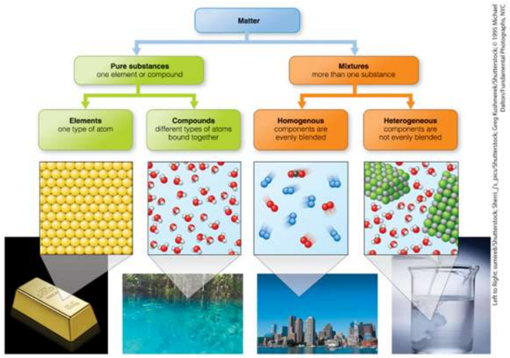
Heterogeneous mixtures – components do not mix evenly.

Homogeneous mixture
Salt mixes evenly with water

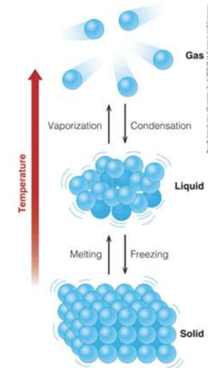
Heterogeneous mixture
Sand separates from water

Separating Mixtures:

Defining Matter



Transitions Between Three States of Matter



Three States of Matter

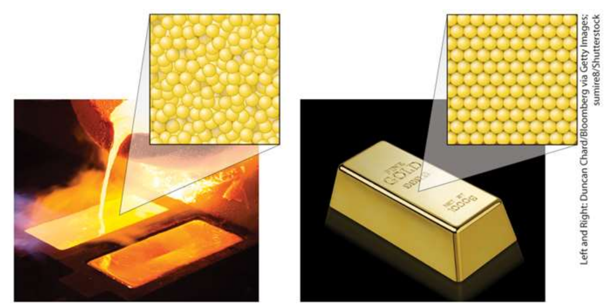
Solid Definite shape, definite volume

Liquid Definite volume, but no definite shape

Gas No definite shape or volume

Particle Arrangement

The behavior of any substance is determined by the arrangement of the particles that compose the substance.



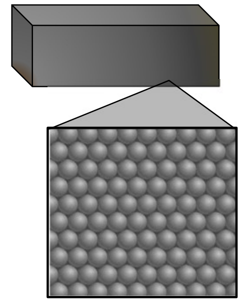
Properties and Changes, Part 1

Physical Properties Can be measured without changing the identity of the substance

- mass
- volume
- temperature
- color
- hardness

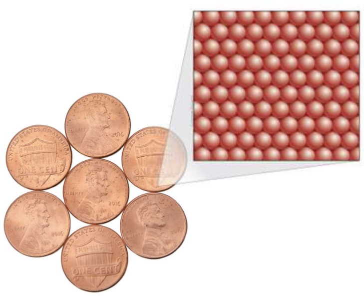
Physical Changes

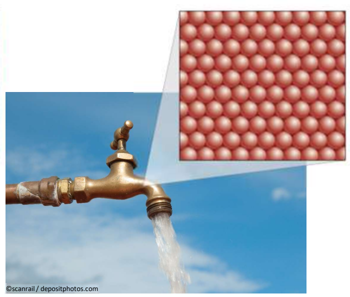
Phase changes are physical changes.



Properties and Changes, Part 2

Physical Changes Don’t change the identity of the substance

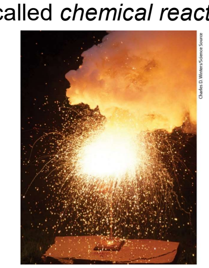




Properties and Changes

Chemical Properties: Can NOT be measured without changing the identity of the substance.

Chemical Changes: Change the identity of the substance - also called *chemical reactions.*



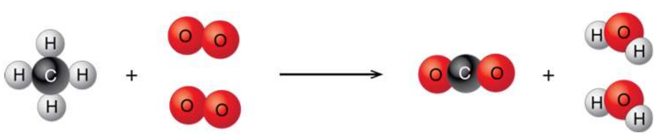
Elements combine to form compounds: a chemical change.

Properties and Changes

Chemical -
Change the identity of a substance.

Physical -
Do NOT change the identity of a substance

A change that forms new compounds: a chemical change.



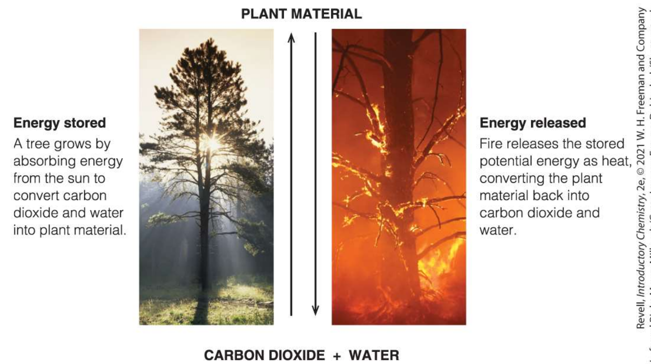
Energy and Change

Energy: The ability to do work
Potential energy: Energy that is stored
Kinetic energy: The energy of motion

Heat Energy

Heat energy: involves the kinetic energy of the particles in a substance

Energy Changes



Physical and chemical changes involve changes in energy.

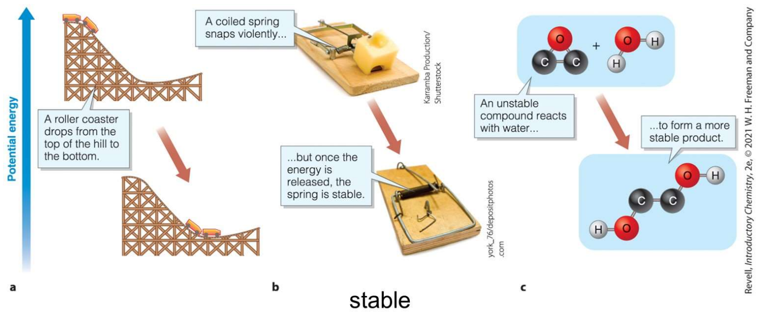
Moving from higher energy to lower energy

Moving from lower energy to higher energy

High Energy or Stable?

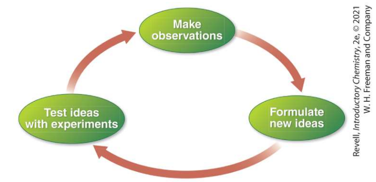
Potential Energy

high energy



stable

The Scientific Method



hypothesis: A tentative explanation that has not been tested.

theory: An idea supported by experimental evidence, or a *paradigm*, or way of thinking about a topic.

scientific law: A statement that describes observations that are true in widely varying circumstances.

Exothermic and Endothermic Change

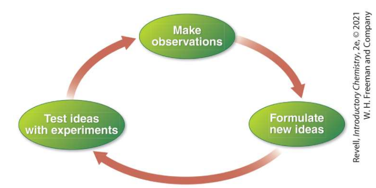
Exothermic change: releases heat energy

Endothermic change: absorbs heat energy

exothermic

endothermic

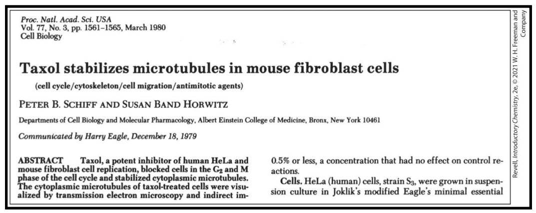
The Scientific Method, Continued



Theories	**Laws**
How or *Why* it happens	*What* happens

Scientific Communication

Scientists communicate findings through scientific papers.



Proc. Natl. Acad. Sci. USA
Vol. 77, No. 3, pp. 1561–1565, March 1980
Cell Biology

Taxol stabilizes microtubules in mouse fibroblast cells

(cell cycle/cytoskeleton/cell migration/antimitotic agents)

PETER B. SCHIFF AND SUSAN BAND HORWITZ

Departments of Cell Biology and Molecular Pharmacology, Albert Einstein College of Medicine, Bronx, New York 10461

Communicated by Harry Eagle, December 18, 1979

ABSTRACT Taxol, a potent inhibitor of human HeLa and mouse fibroblast cell replication, blocked cells in the G_2 and M phase of the cell cycle and stabilized cytoplasmic microtubules. The cytoplasmic microtubules of taxol-treated cells were visualized by transmission electron microscopy and indirect im- 0.5% or less, a concentration that had no effect on control reactions.

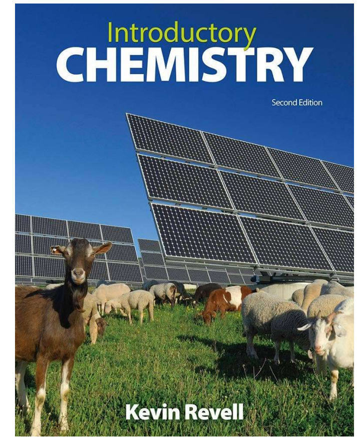
Cells. HeLa (human) cells, strain S_3, were grown in suspension culture in Joklik's modified Eagle's minimal essential

inga spence/Alamy age; Courtesy of Dr. Susan Band Horwitz

Introductory Chemistry
Chem 103

Chapter 2 – Measurement

Lecture Slides



Scientists

Large and Small Measurements

Photo credit: Tetra Images/Getty Images

Earth to the Sun:
149,600,000,000 meters

Photo credit: Philippe Crassous/FEI/REX/Shutterstock

Hydrothermal worm:
0.0005 m

Scientific Notation

$$2.14 \times 10^{-3}$$

exponent

coefficient

multiplier

Examples of Exponential Notation

5.1×10^{3}	=	5100.
5.1×10^{2}	=	510.
5.1×10^{1}	=	51.
5.1×10^{0}	=	5.1
5.1×10^{-1}	=	0.51
5.1×10^{-2}	=	0.051
5.1×10^{-3}	=	0.0051

Exponential Notation

→ 10^{3}	=	$10 \times 10 \times 10$	=	1,000.
→ 10^{2}	=	10×10	=	100.
→ 10^{1}	=	10	=	10.
→ 10^{0}	=	1	=	1.
→ 10^{-1}	=	$\frac{1}{10}$	=	0.1
→ 10^{-2}	=	$\frac{1}{10\times10}$	=	0.01
→ 10^{-3}	=	$\frac{1}{10\times10\times10}$	=	0.001

Going from Standard to Scientific Notation:

2,500,000 L $= 2.5 \times 10^{6}$ L

move 6 digits

137,000,000,000 J $= 1.37 \times 10^{11}$ J

move 11 digits

0.000000142 g $= 1.42 \times 10^{-7}$ g

move 7 digits (right)

0.000326 cm $= 3.26 \times 10^{-4}$ cm

move 4 digits (right)

Going from Scientific to Standard Notation:

→ 1.528×10^5 kg   1.52800   = 152,800 kg

→ 1.64×10^7 L   1.6400000   = 16,400,000 L

→ 1.35×10^{-5} m   00001.35   = 0.0000135 m

→ 8.28×10^{-3} g   008.28   = 0.00828 g

Calculations Involving Scientific Notation, Example 1

multiplication

add exponents

$3.1 \times 10^4 \times 2.0 \times 10^2 = 6.2 \times 10^6$

multiply coefficients

Calculations Involving Scientific Notation, Example 2

division

divide coefficients

$$\frac{8.4 \times 10^7}{2.0 \times 10^3} = 4.2 \times 10^4$$

subtract exponents

Calculations Involving Scientific Notation, Example 3

$2.5 \times 10^4 \times 6.0 \times 10^8 = 15. \times 10^{12}$

increase exponent

move 1 digit

$= 1.5 \times 10^{13}$

Using a Calculator For Scientific Notation:

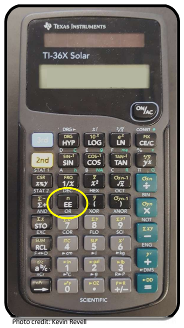


Photo credit: Kevin Revell

EE E Exp

"× 10⁻"

1.53 x 10^16

1.53 EE *16*

Measurement and Units, Continued

Photo credits clockwise from top left: Kevin Revell; Maisie Paterson/Getty Images; Charles Brutlag/Shutterstock; Kevin Revell

Measurement and Units

units of measurement Quantities with accepted values that can be communicated between people.

Units

Common English and Metric Units

Measurement	Metric Unit	English Unit	Relationship
Length	meter (m)	foot (ft) mile (mi)	1 m = 3.280 ft 1 km = 0.621 mi
Mass or Weight	kilogram (kg)	pound (lb)	1 kg = 2.204 lb
Volume	liter (L)	gallon (gal)	1 liter = 0.264 gal

Units, Continued

Fundamental Units

Measurement	Unit
Mass	kilogram (kg)
Length	meter (m)
Time	second (s)
Temperature	kelvin (K)
Light Intensity	candela (cd)
Electric current	ampere (A)
Amount	mole (mol)

Derived Units

Measurement	Units
Volume	m^3
Velocity	m/s
Density	kg/m^3

Using Common Metric Prefixes

1. *How many meters are in a kilometer?*

   1 km = 1,000 m

2. *How many A are in a MA?*

   1 MA = 1,000,000 A

3. *How many mg are in a g?*

   $1\ mg = \frac{1}{1,000}\ g$

   1,000 mg = 1 g

Table 2.5 Common Metric Prefixes

Prefix	Symbol	Meaning	
Mega-	M	10^6	1,000,000
Kilo-	k	10^3	1,000
Milli-	m	10^{-3}	$\frac{1}{1,000}$

Metric Prefixes

Table 2.5 Common Metric Prefixes

Prefix	Symbol	Meaning	
Tera-	T	10^{12}	1,000,000,000,000
Giga-	G	10^9	1,000,000,000
Mega-	M	10^6	1,000,000
Kilo-	k	10^3	1,000
Deci-	d	10^{-1}	$\frac{1}{10}$
Centi-	c	10^{-2}	$\frac{1}{100}$
Milli-	m	10^{-3}	$\frac{1}{1,000}$
Micro-	μ	10^{-6}	$\frac{1}{1,000,000}$
Nano-	n	10^{-9}	$\frac{1}{1,000,000,000}$
Pico-	p	10^{-12}	$\frac{1}{1,000,000,000,000}$

160,000,000 bits

= 160 megabits

0.0000032 grams

= 3.2×10^{-6} grams

= 3.2 micrograms

Describing the Quality of Measurements

Photo credit: James A. Prince/Science Source

Precision and Accuracy

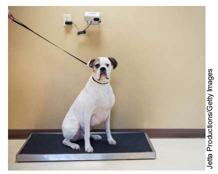
±0.0001 g

Accuracy

- How reliable are the measurements?
- Do they reflect the true value?

Precision

- How finely are the measurements made?
- How closely are they grouped together?

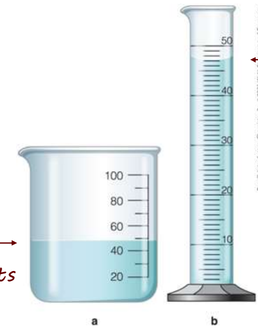


±0.1 kg

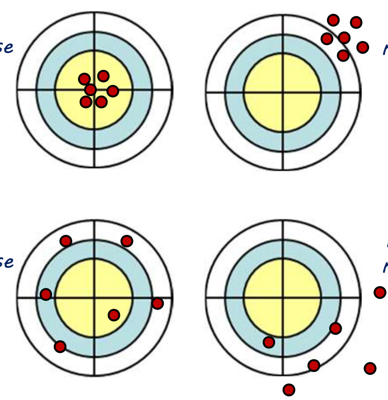
Significant Digits

We can estimate *one digit* between the marked values.

Significant digits: Indicate how precisely we know a measurement



Precision and Accuracy, Continued



Identifying Significant Digits, Part 1

1. All nonzero digits are significant, and all zeros between nonzero digits are significant.

   1.2571 g — *5 sig. digits* 1.1052 cm — *5 sig. digits*

2. If a decimal point is present, zeros to the right of the last nonzero digit are significant.

   5.01 g
   5.00 g *3 sig. digits*
   4.99 g

Identifying Significant Digits, Part 2

3. Zeros to the left of the nonzero numbers are never significant.

000012 kg — 000: *not significant*; 12: *2 sig. digits*

0.0045 m — 0.00: *not significant*; 45: *2 sig. digits*

How many significant digits are in 4.5 mm? 2

4.5 mm = 0.0045 m 2

Place holders for the decimal

Defining Significant Digits for Large Numbers

10,000 kg — *3 sig. digits*

10,000 ± 100 kg

1.00 x 10^4 kg

Identifying Significant Digits, Part 3

4. If there is no decimal point present, zeros to the right of the last nonzero *may* or *may not* be significant.

$ 11,000,000

$ 9,000,000

$ 10,000,000

?

$ 10,001,000

$ 9,999,000

Summary of Significant Digits

Significant digits show the precision of a measured quantity.

- Significant:
  - nonzeros 1.2571 g
  - zeros between nonzeros 1.1052 cm
  - zeros after the decimal point 1.100 mm
- Not Significant
  - zeros to the left of all nonzeros 000023 L, 0.0031 mg
- May be Significant
  - zeros to the right of nonzeros with no decimal 47,000,000 kg

Exact Numbers

Values for which there is no uncertainty

- Counted Values

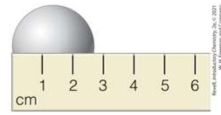
Exactly 7 pennies

- Defined Values

1,000 mg = 1 g

3 feet = 1 yard

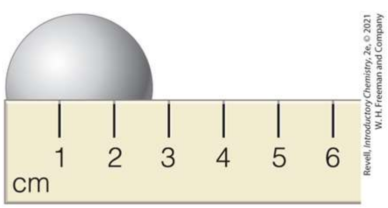
Example: What is the circumference of the ball?



Circumference = πd

Diameter	Calculated Circumference	
2.6 cm	8.16814090 cm	
2.7 cm	8.48230016 cm	8.5 cm
2.8 cm	8.79645943 cm	

Calculations with Significant Digits



2.6 cm

2.7 cm

2.8 cm

estimated

Multiplication and Division with Significant Digits

1. When multiplying or dividing, report the same number of digits as are in the least precise starting measurement.

A vehicle travels 315.3 miles in the span of 5.2 hours. What is its average speed, in miles per hour?

$$\frac{315.3 \text{ miles}}{5.2 \text{ hours}} = 60.\sout{63461538}$$

4 sig. digits (315.3); 2 sig. digits (5.2)

$$= 61 \text{ miles/hour}$$

Addition and Subtraction with Significant Digits

2. When adding or subtracting, round to the last decimal place of the least precise starting measurement.

While training for a triathlon, you swim 0.432 miles, then bike 18.1 miles. What was your total distance traveled?

Swim 0.432 mi.
\+ Bike 18.1 mi.
= 18.~~532~~ mi.

= 18.5 mi.

Example with Significant Digits

A chemist measures the mass of chloride in three water samples, as shown in the table. Together, the three samples have a volume of 2.31 liters. What is the average mass of chloride per liter of water? Answer to significant digits.

Sample	Mass of Chloride
A	15.21 mg
B	9.33 mg
C	11.329 mg

total mass chloride:

15.21 mg
9.33 mg
11.329 mg
35.869 mg

= 35.87 mg
4 sig. digits

$\frac{\text{total mass}}{\text{volume}}$

$= \frac{35.869\ \text{mg}}{2.31\ \text{L}}$

= 15.5~~2770563~~

= 15.5 mg/L

Use unrounded mass

4 sig. digits

3 sig. digits

Rounding Calculations with Significant Digits

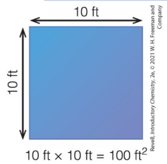
If a calculation involves multiple steps, wait until the end to round to significant digits.

Unit Conversions

Photo credit clockwise from top: nickichen/Shutterstock ; eye35 / Alamy

Unit Conversions: Dimensional Analysis, Example 1

Whatever we do to the number, we also do to the units.



multiply number → multiply units

$10 \text{ ft} \times 10 \text{ ft} = 100 \text{ ft}^2$

$15.0 \text{ mL} \times \frac{1.02 \text{ mg}}{1 \text{ mL}} = 15.3 \text{ mg}$

Unit Conversions: Dimensional Analysis, Example 2

A copper pellet has a mass of 0.281 kg. What is this mass in grams?

1 kg = 1,000 g

$\frac{1 \text{ kg}}{1{,}000 \text{ g}} = 1$ *or* $\frac{1{,}000 \text{ g}}{1 \text{ kg}} = 1$

conversion factors

$0.281 \text{ kg} \times \frac{1{,}000 \text{ g}}{1 \text{ kg}} = 281 \text{ g}$

starting unit (mg) — conversion factor — ending unit (g)

Unit Conversions: Dimensional Analysis, Example 3

How many inches are in 326 cm?

2.54 cm = 1 inch

$\frac{2.54 \text{ cm}}{1 \text{ inch}}$ or $\frac{1 \text{ inch}}{2.54 \text{ cm}}$

starting unit (cm) x conversion factor = ending unit (inches)

$326 \text{ cm} \times \frac{1 \text{ inch}}{2.54 \text{ cm}} = 128 \text{ inches}$

$326 \text{ cm} \times \frac{2.54 \text{ cm}}{1 \text{ inch}} = 828 \frac{\text{cm}^2}{\text{inch}}$ X wrong units

Unit Conversions: Dimensional Analysis, Example 4

The speed of light in a vacuum is 3.00 × 10⁸ m/s. What is this speed in miles per hour?

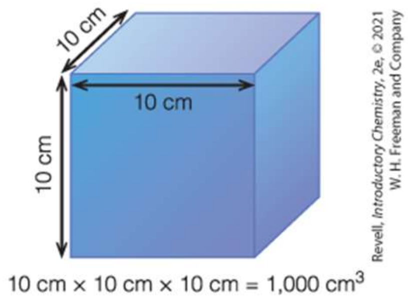
1 mile = 1609.3 meters

1 hour = 3600 seconds

$3.00 \times 10^8 \frac{\text{m}}{\text{s}} \times \frac{1 \text{ mi}}{1609.3 \text{ m}} \times \frac{3600 \text{ s}}{1 \text{ hr}} = 6.71 \times 10^8 \frac{\text{mi}}{\text{hr}}$

meters to miles — seconds to hours

Units of Volume



Volume Sizes

m^3

dm^3

cm^3

Volume Sizes, Continued

liter (L):

$1\ L = 1\ dm^3$

milliliter (mL)

1 milliliter = 1 cubic centimeter

$1\ mL = 1\ cm^3$

Cubic Decimeters and Cubic Meters

How many cubic decimeters are in one cubic meter?

$1\ m = 10\ dm$

$(1\ m)^3 = (10\ dm)^3$

$1\ m^3 = 1{,}000\ dm^3$

Cubic Centimeters and Cubic Meters

How many cubic centimeters are in one cubic meter?

$$1\ m = 100\ cm$$

$$(1\ m)^3 = (100\ cm)^3$$

$$1\ m^3 = 1{,}000{,}000\ cm^3$$

Relating mass and volume: density

Example, Multiple Unit Conversions

A hospital administers an IV fluid at a rate of 95.0 cm³ per hour. How many liters of this fluid does the patient receive per day?

Volume

$1\ cm^3 = 1\ mL$

$1{,}000\ mL = 1\ L$

$1{,}000\ cm^3 = 1\ L$

Time

$24\ hr = 1\ day$

$$95.0\ \frac{\cancel{cm^3}}{\cancel{hr}} \times \frac{1\ L}{1000\ \cancel{cm^3}} \times \frac{24\ \cancel{hr}}{1\ day} = 2.28\ \frac{L}{day}$$

cm^3 to liters

hours to days

Density

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

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

Density Examples

Density, Example 1

A saltwater solution has a mass of 11.29 g, and a volume of 10.4 mL. What is the density of this solution?

$$d = \frac{m}{V} = \frac{11.29\ g}{10.4\ mL} = 1.09\ g/mL$$

Density, Example 2

An antifreeze mixture has a density of 1.06 g/mL. If you measure out 600.0 g of this solution, what volume will it occupy?

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

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

$$V = \frac{600.0\ \cancel{g}}{1.06\ \frac{\cancel{g}}{mL}} = 566\ mL$$

Density, Example 3

Aluminum has a density of 2.70 g/cm³. What is the mass of a block of aluminum with a volume of 1.32 L?

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

$$1\ cm^3 = 1\ mL$$

$$1{,}000\ cm^3 = 1\ L$$

$$m = dV$$

$$1.32\ \cancel{L} \times \frac{1000\ cm^3}{1\ \cancel{L}} = 1320\ cm^3$$

$$m = (2.70\ \frac{g}{\cancel{cm^3}})(1320\ \cancel{cm^3})$$

$$m = 3{,}560\ g$$

Densities of Common Materials

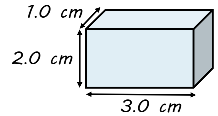
Table 2.6 Densities of Common Materials

Material	Density (g/cm^3)
Aluminum	2.70
Titanium	4.51
Iron	7.87
Copper	8.96
Lead	11.34
Gold	19.31
Water*	1.00
Seawater*	1.02
Air*	0.001

*At 25°C and standard atmospheric pressure

Density, Example 4

A rectangular object measures 3.0 cm x 2.0 cm x 1.0 cm and has a mass of 7.2 g. What is the density of this object? Will it float in water?



V = length x width x height

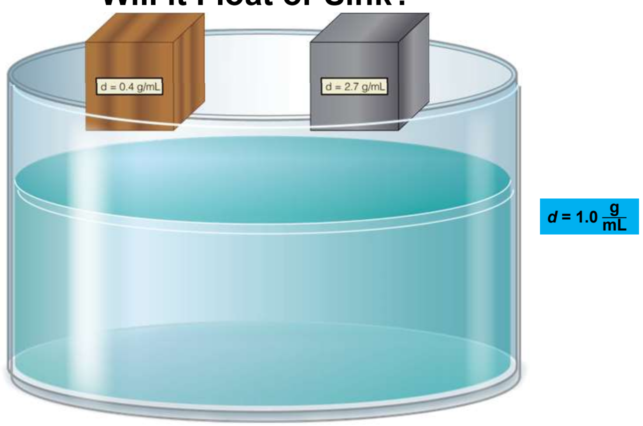
= 3.0 cm x 2.0 cm x 1.0 cm

= 6.0 cm^3

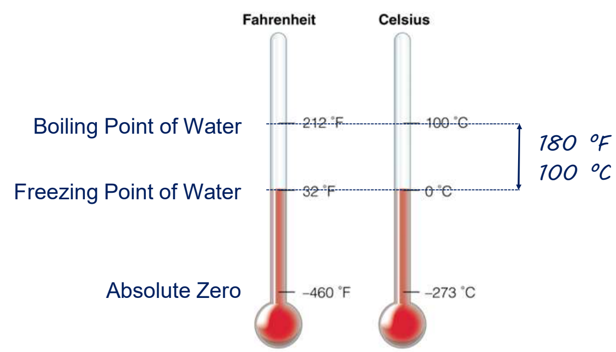
$$d = \frac{m}{V} = \frac{7.2\ g}{6.0\ cm^3} = 1.2\ g/cm^3$$

more dense than water – will not float

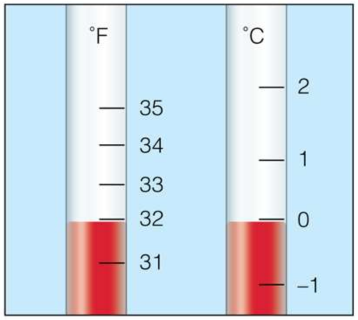
Will it Float or Sink?



Converting Between Temperature Units



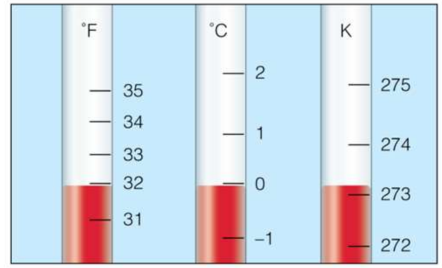
Fahrenheit and Celsius



$$°F = \frac{9}{5}°C + 32$$

$$°C = \frac{5}{9}(°F - 32)$$

Freezing Point on the Three Temperature Scales



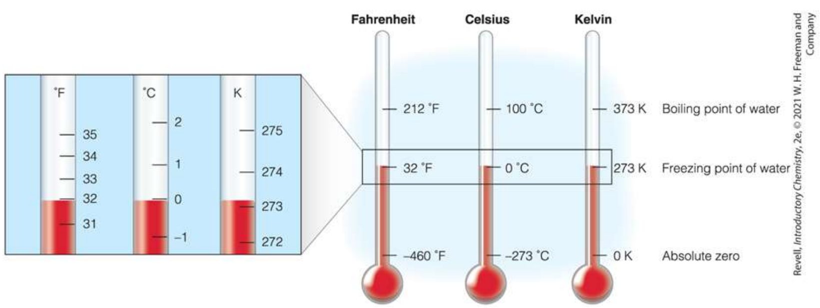
$$K = °C + 273.15$$

"32 degrees Fahrenheit"

"0 degrees Celsius"

"273 kelvins"

The Three Temperature Scales



Temperature Calculation

A refrigerator maintains an inside temperature of 42 °F. Express this temperature in Celsius and in kelvins.

$$°C = \frac{5}{9}(°F - 32)$$

$$°C = \frac{5}{9}(42 - 32) = 5.6\ °C$$

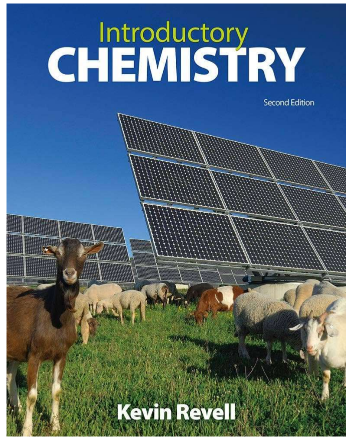
$$K = °C + 273.15$$

$$K = 5.6 + 273.15 = 278.75\ K = 278.8\ K$$

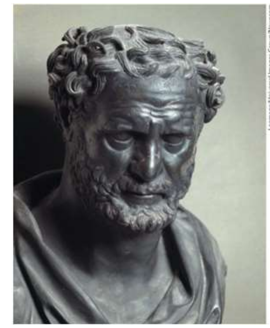
Introductory Chemistry
Chem 103

Chapter 3 – Atoms

Lecture Slides



400 B.C.E. - Democritus



atomos – “indivisible”

Atoms

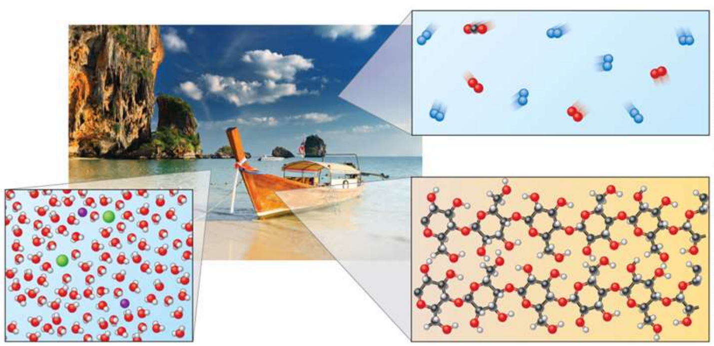


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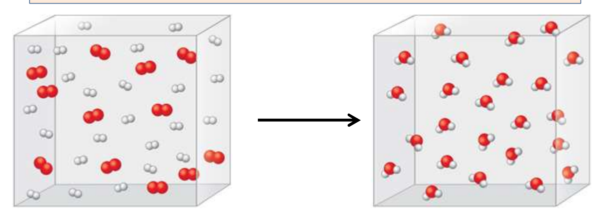
Photo: efired/depositphotos.com

Law of Conservation of Mass

Antoine Lavoisier (1743-1794)

In chemical reactions,
matter is neither created or destroyed.



Hydrogen	+	Oxygen	→	Water
4.0 g	+	32.0 g	=	36.0 g

Example of the Law of Conservation of Mass

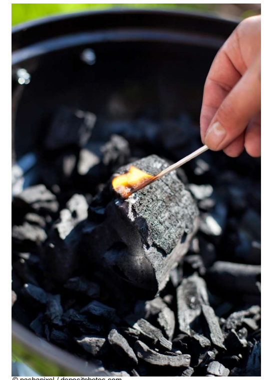
If 16.0 grams of methane react with 64.0 grams of oxygen, 36.0 grams of water are produced. How many grams of carbon dioxide are produced in this reaction?

methane + oxygen ⟶ carbon dioxide + water

16.0 g + 64.0 g = 44.0 g + 36.0 g

Chepko Danil Vitalevich/Shutterstock

Understanding Atomic Theory



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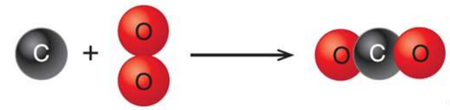

$C + O_2 \longrightarrow CO_2$

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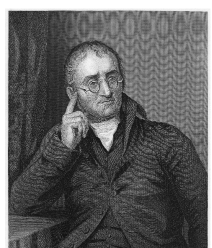
Origins of Atomic Theory

John Dalton (1766-1844)



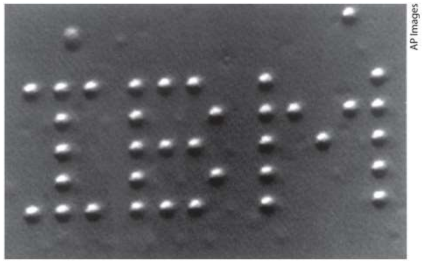
Photos.com/Getty Images

- Elements are made of tiny, indivisible particles called atoms
- The atoms of each element are unique.
- Atoms can join together in whole-number ratios to form compounds.
- Atoms are unchanged in chemical reactions.

Three Foundational Ideas

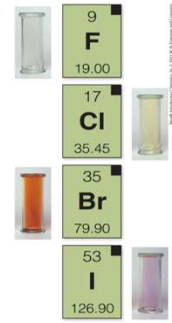
1. All matter is composed of atoms.
2. The atoms of each element have unique characteristics and properties.
3. In chemical reactions, atoms are not changed, but combine in whole-number ratios to form compounds.

Can we see atoms?

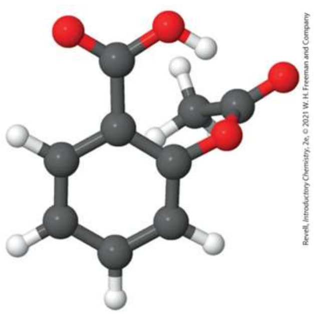


Periodic Table of the Elements

Mendeleev

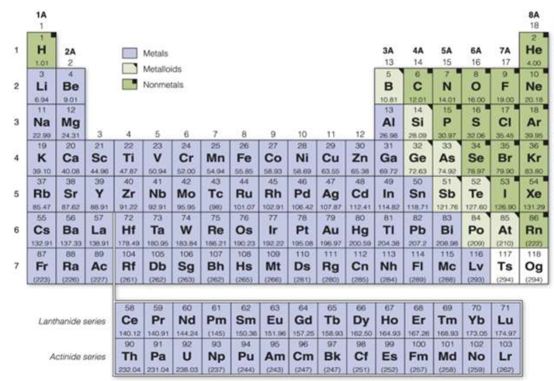


Scientists use X-ray crystallography to visualize the arrangement of atoms



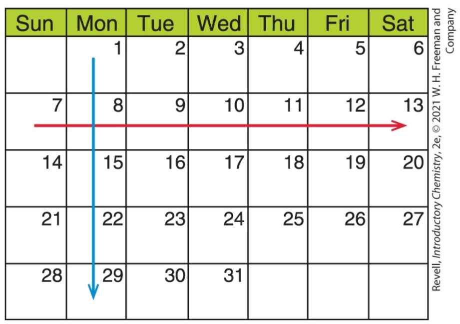
PDB ID: 1GZX
Paoli *et al*, 2002

Periodic Table of the Elements, Continued

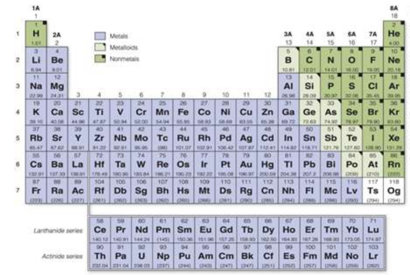


The Meaning of Periodic

A calendar is <u>periodic</u>…



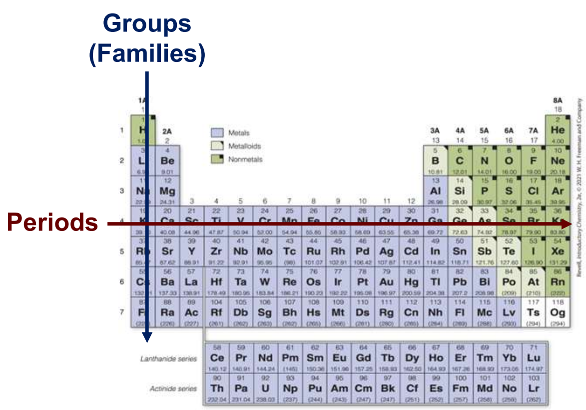
Abbreviations for the Elements



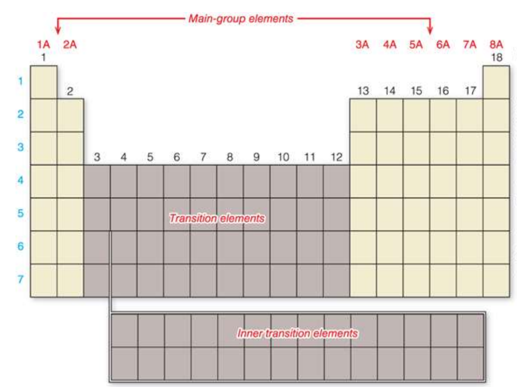
Name	Symbol
carbon	C
hydrogen	H
magnesium	Mg
calcium	Ca

Name	Symbol	Latin Name
sodium	Na	*natrium*
iron	Fe	*ferrum*
copper	Cu	*cuprum*
lead	Pb	*plumbum*

Groups and Periods

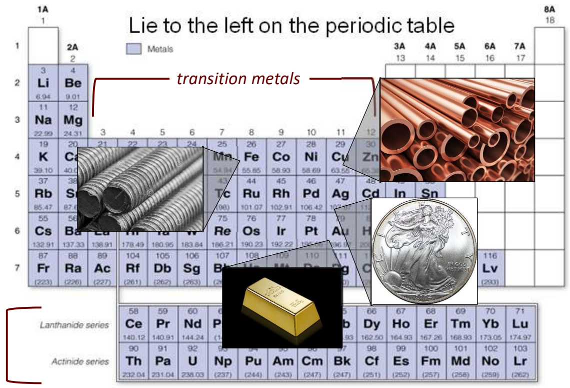


Blocks of Elements



Metals

Lie to the left on the periodic table



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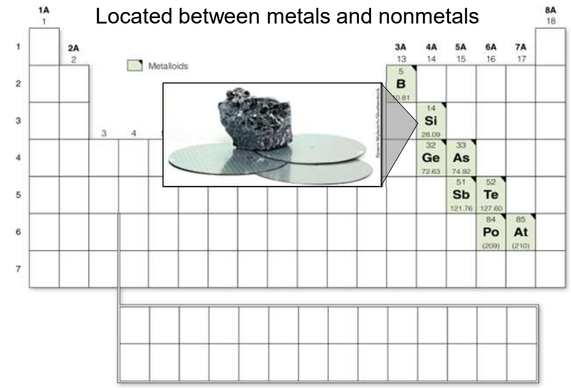
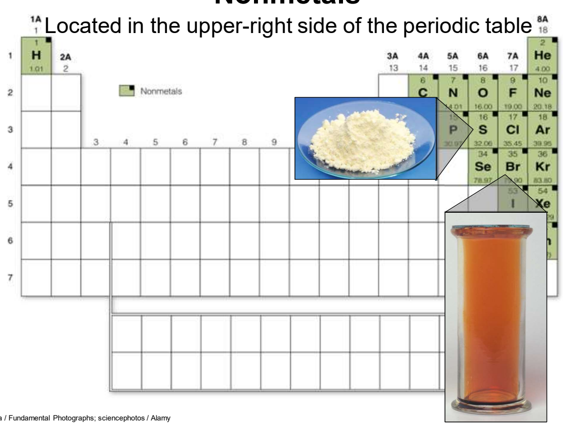
Metalloids

Located between metals and nonmetals



Nonmetals

Located in the upper-right side of the periodic table



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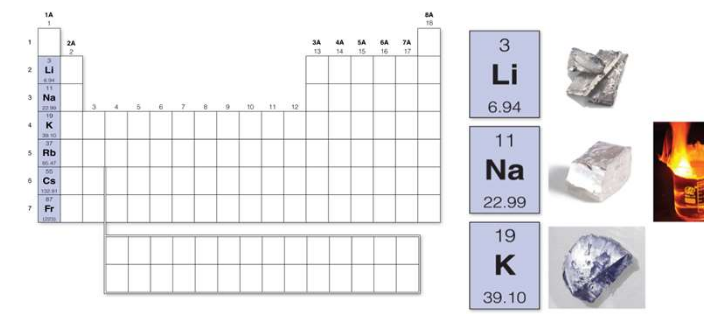
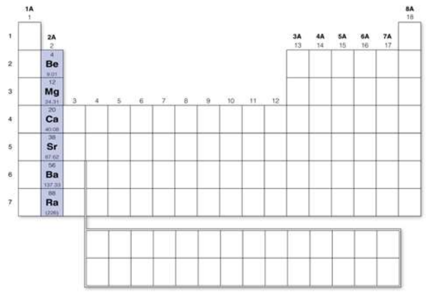
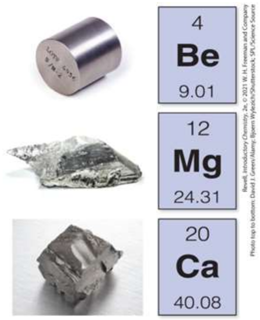
Group 1A: Alkali Metals



- Soft metals
- React violently with water

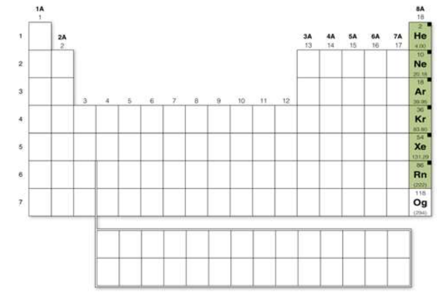
Group 2A: Alkaline Earth Metals

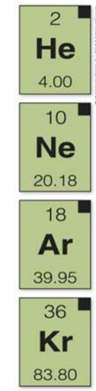




- Less reactive than group 1A
- burn brightly

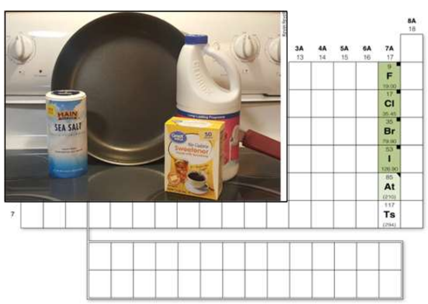
Group 8A: Noble Gases

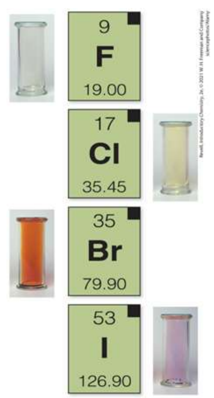




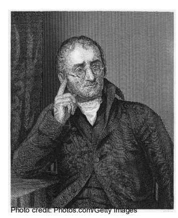
- generally do not form compounds
- gases at room temperature

Group 7A: Halogens





- diatomic molecules in elemental form
- Form many different compounds



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Uncovering Atomic Structure

- The atoms of each element are unique.
- Atoms combine in whole-number ratios to form compounds.
- Atoms are not created or destroyed in chemical reactions.

subatomic particles particles that make up atoms

Describing particles

Mass

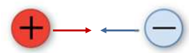
atomic mass unit (u)

1 u = 1.66 × 10^{-27} kg

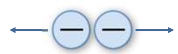
hydrogen atom:
mass = 1.0 u

Charge

opposite charges attract



like charges repel



Separating Elements from Molecules

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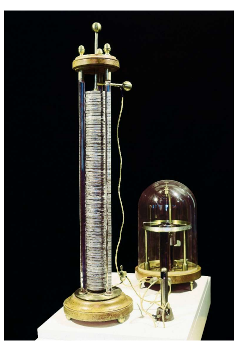
Charles D. Winters/Science Source

Volta

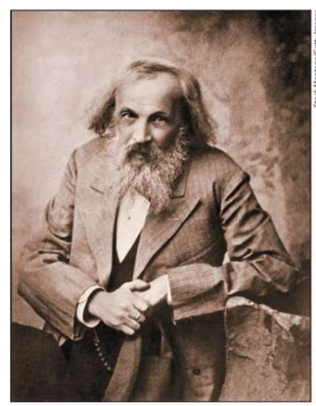
1800: The year that changed chemistry



BeBa/Iberfoto/The Image Works

Volta invents electrochemical cell (battery)

Mendeleev's Periodic Table



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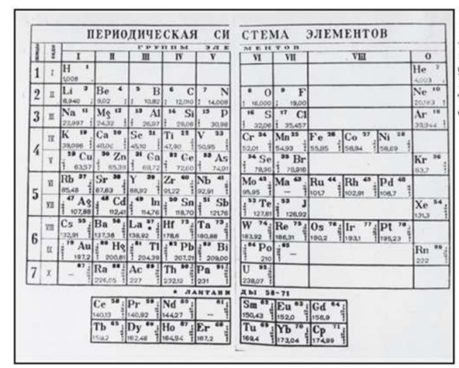

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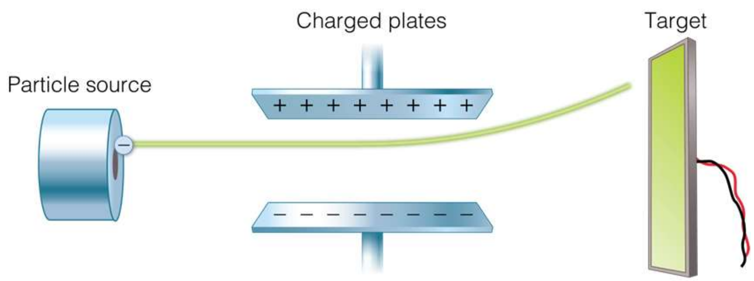
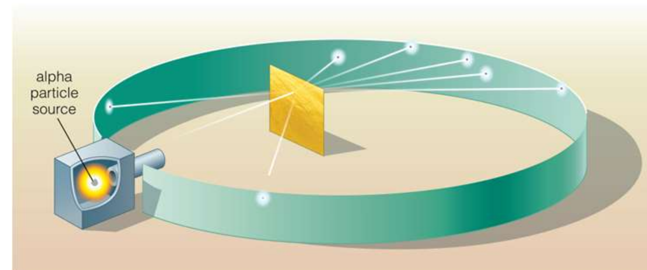
Identification of Charged Particles



Electron: a tiny, negatively-charged particle

Ernest Rutherford



Plum Pudding Model

atoms

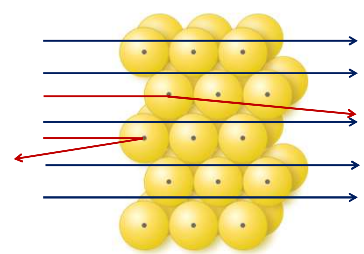
negative electrons | positive ?

Plum pudding model

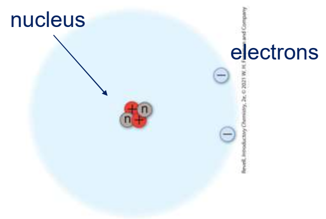
envisioned negative electrons spread throughout a positively-charged material.

Rutherford's Conclusions

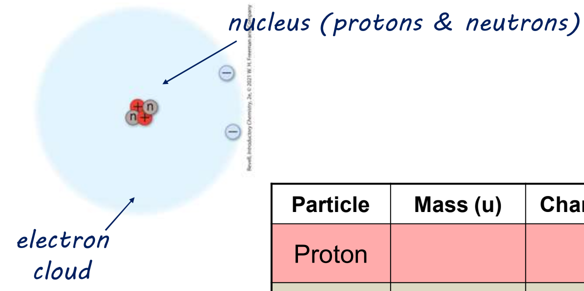
The atom is mostly empty space, with a dense nucleus at the center.



Model of an Atom



Atomic Particles

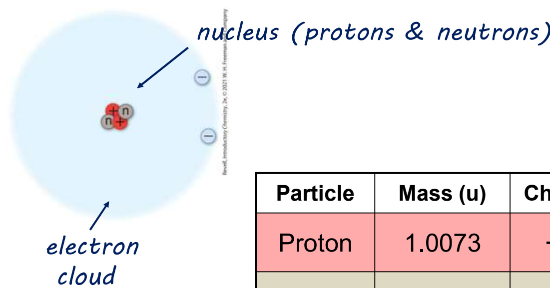


Particle	Mass (u)	Charge
Proton		
Neutron		
Electron		

The Volume of an Atom

Photo credit: Jim West/ Alamy

Atomic Particles, Continued



Particle	Mass (u)	Charge
Proton	1.0073	+1
Neutron	1.0087	--
Electron	0.0005	−1

Atomic Identity

The number of protons determines the identity of the atom.

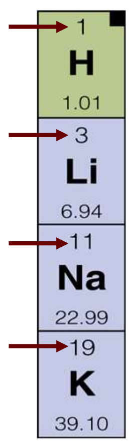
1 proton – hydrogen
2 protons – helium
3 protons – lithium
4 protons – beryllium

Isotopes

Have the same atomic number, but different mass number

Three isotopes of hydrogen:

Atomic Number and Mass Number



Atomic number

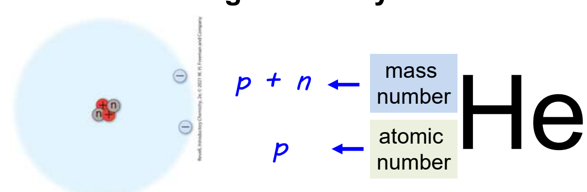
The number of protons in an atom

Also the number of electrons in a neutral atom

Mass number

The number of protons + neutrons

Writing Atomic Symbols



Helium:
2 protons
2 neutrons

$^{4}_{2}He$

Example of Writing Atomic Symbols

An important isotope of uranium has 92 protons and 143 neutrons. Write the symbol with the atomic and mass numbers.

$92 + 143 = 235$

$^{235}_{92}U$

92
U
238.03

average atomic mass

Example of Weighted Average

We have a large number of poker chips. 10% of the chips are $1 chips, and 90% are $2 chips. What is the average value of the chips?

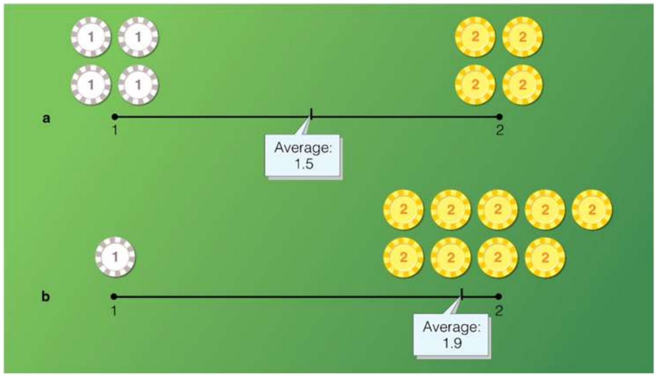
average value = (value A × fraction A) + (value B × fraction B)

average value of chips = ($1 x 0.10) + ($2 x 0.90)

= $1.9

Average Atomic Mass

A weighted average of the different isotopes of an element.



Example of Average Atomic Mass of Carbon

Carbon atoms exist primarily as two isotopes:

^{12}C: mass = 12.0000 u (98.93%)

^{13}C: mass = 13.0034 u (1.07%)

What is the average atomic mass for carbon?

Average mass of carbon

= (12.0000 u)(0.9893) + (13.0034 u)(0.0107)

= 12.01 u

6
C

Summary of Atoms and Elements

- The protons determine the identity of the atoms
- **atomic number**: protons
- **mass number:** protons + neutrons
- **isotopes:** same number of protons, different neutrons
- The periodic table: atomic number and the average atomic mass.

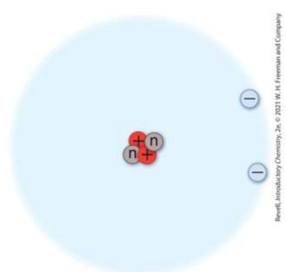
mass number / atomic number He

78
Pt
195.08

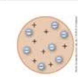
History of the Atom, Part 1

Model	Year		Key Ideas
Dalton's Atomic Theory	1808		Atoms are indivisible particles.

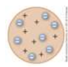
Electrons – A Preview

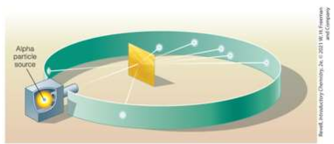


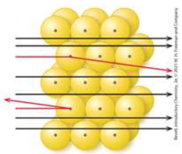
History of the Atom, Part 2

Model	Year		Key Ideas
Dalton's Atomic Theory	1808		Atoms are indivisible particles.
Plum Pudding Model	1904	 	Negative electrons are spread through a positive matrix.

History of the Atom, Part 3

Model	Year		Key Ideas
Dalton's Atomic Theory	1808		Atoms are indivisible particles.
Plum Pudding Model	1904		Negative electrons are spread through a positive matrix.

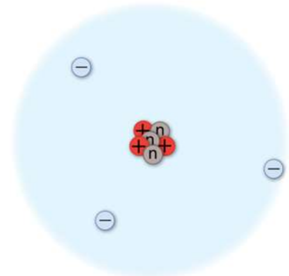




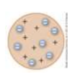
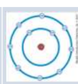
Ions, Part 1

Atoms gain or lose electrons to form *ions*.

Ion: An atom or group of atoms with an overall charge.



History of the Atom, Part 4

Model	Year		Key Ideas
Dalton's Atomic Theory	1808		Atoms are indivisible particles.
Plum Pudding Model	1904		Negative electrons are spread through a positive matrix.
Bohr Model	1913		Electrons orbit the nucleus like planets orbit the sun.
Quantum Model	1920s		Electrons behave both as particles and as waves.

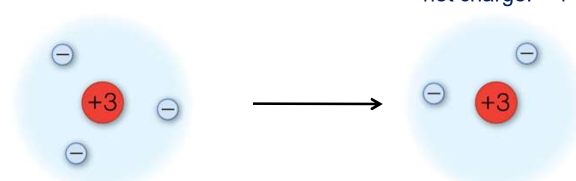
Ions, Part 2

Atoms gain or lose electrons to form *ions*.

Ion: An atom or group of atoms with an overall charge.

lithium atom
3 protons, 3 electrons

lithium ion:
3 protons, 2 electrons
net charge: +1



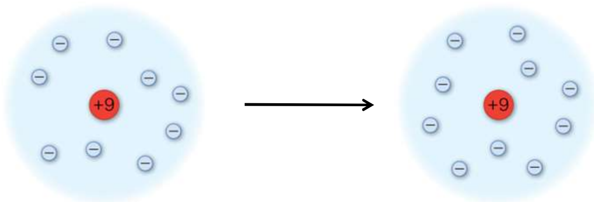
Ions, Part 3

Atoms gain or lose electrons to form *ions*.

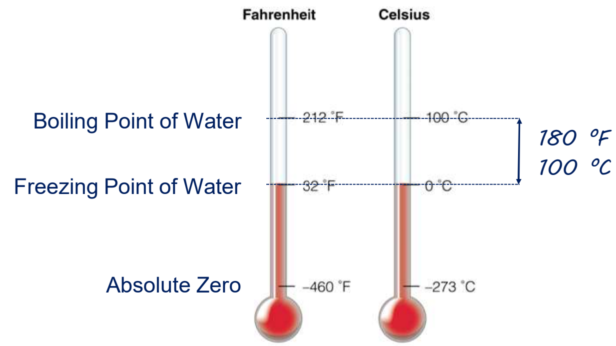
Ion: An atom or group of atoms with an overall charge.

fluorine atom
9 protons, 9 electrons

fluoride ion:
9 protons, 10 electrons
net charge: −1



Converting Between Temperature Units



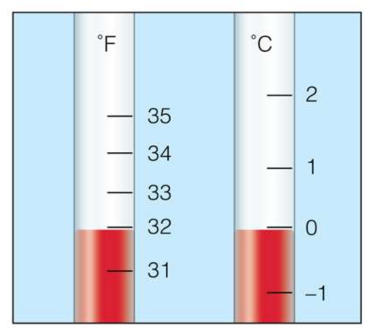
Example of Ions

Sulfur is atomic number 16. Sulfur atoms commonly form sulfide ions, which have a charge of −2. How many electrons are in the electron cloud of a sulfide ion?

sulfur atom:	sulfide ion: (-2)
16 protons	16 protons
16 electrons	18 electrons

16
S
32.06

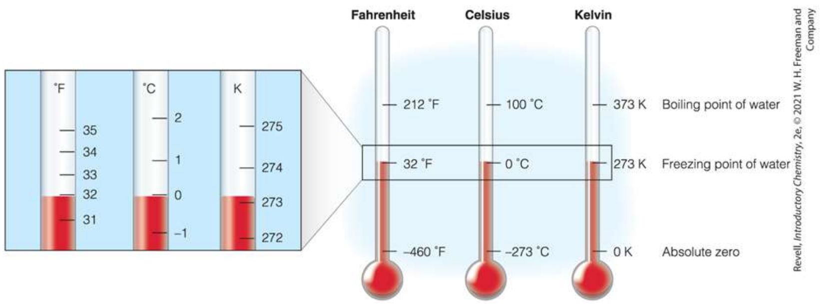
Fahrenheit and Celsius



$$°F = \frac{9}{5}°C + 32$$

$$°C = \frac{5}{9}(°F - 32)$$

The Three Temperature Scales



Temperature Calculation

A refrigerator maintains an inside temperature of 42 °F. Express this temperature in Celsius and in kelvins.

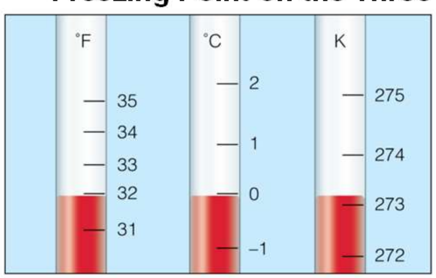
$$°C = \frac{5}{9}(°F - 32)$$

$$°C = \frac{5}{9}(42 - 32) = 5.6\ °C$$

$$K = °C + 273.15$$

$$K = 5.6 + 273.15 = 278.75\ K = 278.8\ K$$

Freezing Point on the Three Temperature Scales



$$K = °C + 273.15$$

"32 degrees Fahrenheit"

"0 degrees Celsius"

"273 kelvins"