CHEM 118 Syllabus

Course Information

Instructor

Instructor Name: Mr. Michael Stogsdill (he, him)

Email Address: michael.stogsdill@mcc.edu

Phone: (567) 343-1425

Important Dates

Event	Date
First Day of Class	09/03/2024
Registration Closes	09/09/2024
Midterm Grades Due	11/01/2024

Event	Date
Thanksgiving Break - College Closed	11/28/2024 - 12/01/2024
Last Day to Drop with W	12/11/2024
Last Day of Class	12/18/2024
Final Grades Due	12/20/2024

Required Materials

A non-programmable scientific calculator, such as, Texas Instrument TI-30X IIS or Casio FX-115ESPLUS.

Course Description

This course is designed for students who need a review of, an introduction to, or additional experience in the basic mathematical and chemical concepts and skills necessary for satisfactory academic work in CHEM-131.

Prerequisites

- Placement into RDNG-030 (ACSR-030), completion of ELAP-110 with a minimum grade of 2.0, or successful completion of RDNG-016 or ACLT-074 with a minimum grade of 3.0.
- A 2.0 or better in MATH-082.

Student Learning Outcomes

Provide an introduction to or a review of basic chemical concepts and associated mathematical operations.

- Emphasize Chemistry topics known to be troublesome to freshman students.
- Provide the necessary foundation for the chemistry courses designed for both science and science-related majors.

Course Content

Topics

- The Chemical World
- Measurement and Problem Solving
- Matter and Energy

- Atoms and Elements
- Molecules and Compounds
- Chemical Composition
- Chemical Reactions
- Electrons and the Periodic Table
- Chemical Bonding

Textbook

Introductory Chemistry, 6th Edition
Nivaldo J. Tro, Santa Barbara City College
© 2018 |Pearson
ISBN-13: 9780134564074

Description

Builds 21st century and problem-solving skills, preparing students for success. The best-selling Introductory Chemistry text continues to encourage student interest by showing how chemistry manifests in students' daily lives. Author Nivaldo Tro draws upon his classroom experience as an award-winning instructor to extend chemistry from the laboratory to the student's world, capturing student attention with relevant applications and an engaging writing style. The text provides a superior teaching and learning experience,

enabling deep conceptual understanding, fostering the development of problem-solving skills, and encouraging interest in chemistry with concrete examples. Extending chemistry from the lab to the student's world, the text reveals that anyone can master chemistry.

Attendance

Attendance will be taken at each meeting of this course. Students who regularly attend their courses typically perform much better.

Exams

There will be two exams in this course. A midterm exam and a final exam. These exams will be cumulative. Meaning they will include questions on all of the material covered up to that point. The dates of these exams can be found in the schedule below.

Quizzes

Quizzes will be due following the completion of each chapter. Quizzes can be completed outside of class. The must be completed individually.

Grades

Student assessment for the course will be derived from the quizzes and exams.

Assessment	Overall Percentage
Quizzes	40%
Midterm Exam	25%
Final Exam	35%

GPA	Course Grade
4.0	92 - 100%
3.5	88 - 91%
3.0	82 - 87%
2.5	76 - 81%
2.0	70 - 75%
1.5	65 - 69%

GPA	Course Grade
1.0	60 - 64%
0.0	<60%

Plagiarism

Basically, copying anyone's work is considered plagiarism unless that person has been appropriately cited. The copying of another student's work will be considered plagiarism and will not be tolerated.
I am very adept at detecting cheating. Students caught cheating or plagiarizing will be reprimanded as follows: 1st offense results in zero score for the assignment, 2nd offense referred to academic standards. The repercussions for being referred to academic standards are very severe. For more information please refer to the academic integrity statement posted in the Online Learning Resources page within Canvas, the relevant section in the college catalog, and the MCC Academic Integrity Website.

Mott Community College Policies and Resources

Tutoring Information

Mott Community College offers tutoring services. You can learn more about tutoring services **here**.

MCC Tutoring Center

Telephone: 810-232-4696
Email: Peer.Tutoring@mcc.edu

I will be tutoring in the Math and Science Study Center on the second floor of the Gorman Science Center Building from 9:00 am - 12:00 pm on Tuesdays. Please feel free to stop by for additional on-on-one instruction.

Title IX of the Education Amendments of 1972:

Consistent with Title IX of the Education Amendments of 1972, Mott Community College will provide appropriate adjustments and/or support to pregnant and parenting students. With medical documentation from a physician, the adjustments and accommodations may include providing a larger classroom desk or a designated lactation room, allowing frequent trips to the restroom, permitting temporary access to elevators, providing opportunities for making up missed work, allowing the student to submit work

after a missed deadline due to absence related to pregnancy or childbirth, or providing a leave of absence for a period of time.

For questions about your rights pursuant to Title IX and other non-discrimination related compliance matters, contact the **Title IX Coordinator Wanda Brown** at titleix@mcc.edu.

Care Team

If you or someone you know needs support, is distressed, or exhibits concerning behavior, help by making a referral to the Care Team. The Mott Community College Care Team is committed to improving the safety and well-being of the college community through proactive and supportive interventions. As your instructor, I may contact the Care Team to seek support for you. I encourage you to fill out a referral if you or a classmate are in need of help. **You may contact the Care Team by reporting a concern at www.mcc.edu/care. You can also contact the Care Team by calling (810) 762-0331*****.**

In case of an emergency, please call 9-1-1.

Civility Statement

Mott Community College is committed to the highest standards of academic and ethical integrity, acknowledging that civility and respect of self -and others - are the foundation of educational excellence.

Accommodations Policy

Mott Community College is committed to providing equal opportunity for participation in all programs, services, and activities and adheres to Section 504 of the Rehabilitation Act (1973) and the Americans with Disabilities Act, as amended (2008) to provide effective auxiliary aids and services for qualified students with documented disabilities. Requests for accommodations by students with disabilities may be made by contacting Disability Services at 810-232-9181 or on the **website** . Once your eligibility for an accommodation has been determined, you will be issued an Instructor Notification Letter.

Please present Instructor Notification Letters to instructors at the start of the semester and/or two weeks prior to the accommodation date (test, project, etc.). Requests received after this date will be honored whenever possible.

Incomplete Grade Information

A student must initiate a request for an "I" (Incomplete Grade) from an instructor. The "I" will be given at the sole discretion of the instructor. Typically, an "I" will be given only when the student:

- Has completed at least 75% of the class (excluding the final exam) but is unable to complete the class work because of extraordinary, unusual, or unforeseen circumstances or other compelling reasons,
- Has done passing work in the course and in the instructor's judgment, can complete the required work without repeating the course.
- Is eligible under Title IX - Supporting the Academic Success of Pregnant and Parenting Students. Please view **this link** for more information.

Withdrawal Policy

Dropping a class during the semester is also referred to as Withdrawal from a course. This is the responsibility of the student. A student may withdraw from a class at any time prior to the 90% point of the semester or session. A withdrawal before the “Date of Record” means there is no record of enrollment. A withdrawal after the “Date of Record” but before the last week of the semester means a “W” grade. A "W" grade will be listed on the student’s grade report/transcript if the withdrawal occurs after the Date of Record as listed on the Academic/Registration Calendar.

Students must follow this procedure to officially discontinue enrollment in a particular course. To withdraw, a student must fill out the Class Schedule Worksheet form and file it with the Career & Enrollment Center.

Students who decide not to attend classes must drop these classes themselves. Classes are not automatically dropped for non-attendance.

To be eligible for a tuition refund classes must be dropped within the tuition refund dates, which are listed in the Academic/Registration Calendar.

Additional Contact Information

Science and Math division office

Telephone: 810-762-0279

MCC e-Learning office

Telephone: 810-762-0224
Email: elearning@mcc.edu

MCC 2-HELP technical support

Telephone: 810-762-4357
Email: 2help@mcc.edu

School Closings

Up-to-date closing information is available from the following sources:

MCC's Public Safety Webite

Emergency Notification System

As a member of the MCC Community you are initially set to receive (opted-in) notifications as both a voice and text message. In the event of any on-campus emergency a voice or SMS (text message) will be sent. Notifications can be managed through MCC4me > My Mott Apps > Account Manager.

Student Email Notification

Students will be notified through their MCC Webmail address of specific class cancellation, site-specific closings, or college-wide closings. Call (810) 232-8989 for a recorded message announcing a school closing. Area radio and television stations will be alerted whenever a decision is made to close the school.

Schedule

Date	Reading Sections	Assessment
09/05/2024	1.1 - 1.5	
09/12/2024	2.1 - 2.11	Ch. 1 Quiz
09/19/2024	3.1 - 3.7	Ch. 2 Quiz
09/26/2024	3.8 - 3.12	
10/03/2024	4.1 -4.6	Ch. 3 Quiz
10/10/2024	4.7 - 4.9	
10/17/2024	5.1 - 5.5	Ch. 4 Quiz
10/24/2024	5.6 - 5.11	
10/31/2024		Midterm
11/07/2024	6.1 - 6.4	
11/14/2024	6.5 - 6.9	
11/21/2024	7.1 - 7.5	Ch. 6 Quiz

Date	Reading Sections	Assessment
11/28/2024	Thanksgiving	Break
12/05/2024	7.6 - 7.10	
12/12/2024		Final Exam

Chapter 1

Science vs. Philosophy

Science was born out of philosophy. It was developed after it became obvious that philosophy alone could not fully explain the physical universe.

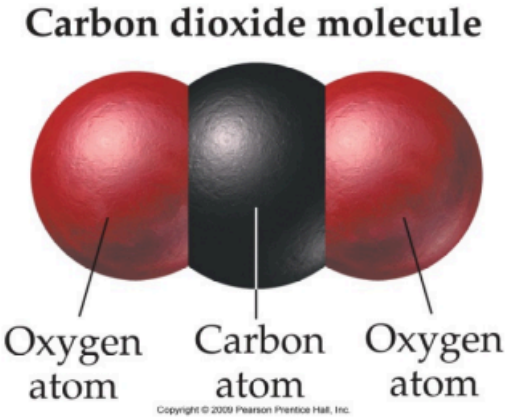
Philosophers	Scientists
Observe Nature	Observe Nature
Explain the Behavior of Nature	Explain the Behavior of Nature
Communicate and Debate Ideas with other Philosophers.	Communicate and Debate Ideas with other Scientists.
Truth is revealed through Logic and Debate	Truth is revealed through Experimentation

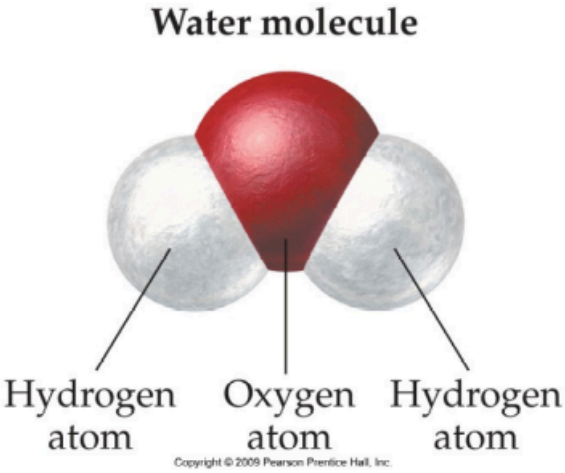
What Is Chemistry?

- Chemists use the **Scientific Method** to discover the relationships between the particle structure of matter and the properties of matter we observe.
    - How does the microscopic properties of matter effect the macroscopic properties?

Structure Determines Properties

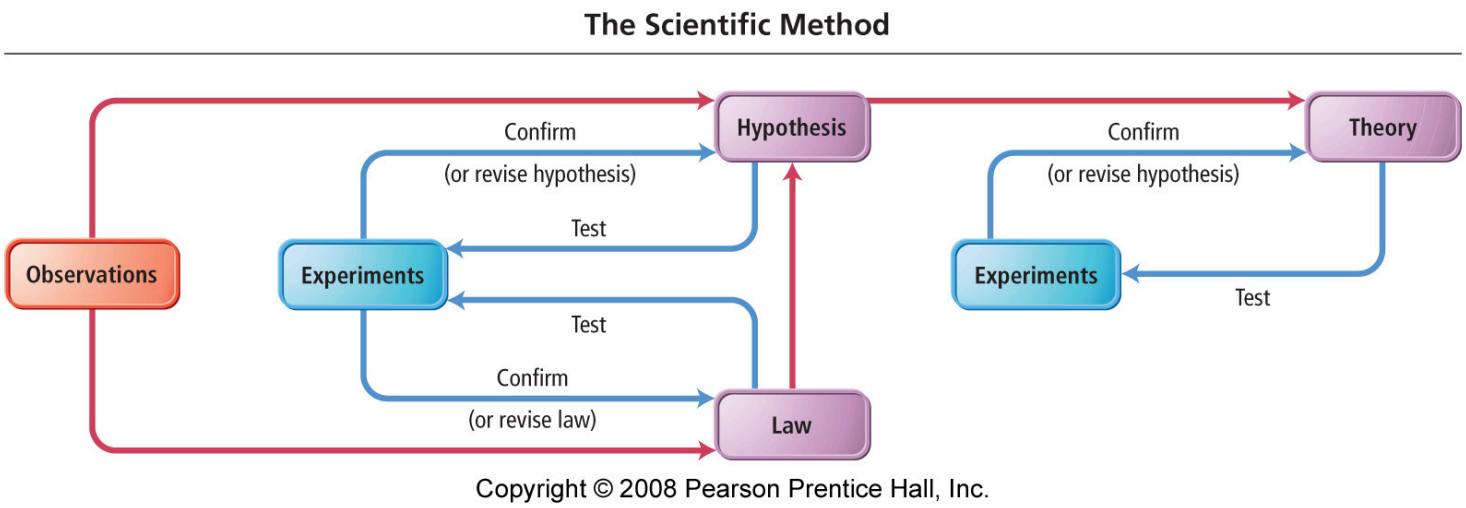
- Everything is made of tiny particles called **atoms** and **molecules**.
- Properties of a substance are determined by the type, amount, and interactions between these pieces.





The Scientific Method

- A process for trying to understand nature by observing it and analyzing the way it behaves.
- **Observations** are made to identify phenomenon to study and collect *data*.
- **Hypotheses** are formed and tested through **experimentation**
- **Conclusions** are drawn by analyzing data obtained from experiments.
- These conclusions are used to *confirm* or *reject* the hypothesis



Observation

- A way of acquiring information about nature.

- The information obtained from observation is known as Data.
- Some observations are simple descriptions about the characteristics or behavior of nature.
    - This is called *qualitative* data

“The soda pop is a liquid with a brown color and a sweet taste. Bubbles are seen floating up through it.”

- Some observations compare a characteristic to a standard numerical scale.
    - This is called *quantitative* data

“A 240-mL serving of soda pop contains 27 g of sugar.”

Hypothesis

- A tentative interpretation or explanation of your observations.

“The sweet taste of soda pop is due to the presence of sugar.”

- A good hypothesis is one that can be tested to be proven wrong.
    - One test should be able to invalidate your hypothesis.

Experiments

- Tests of hypotheses, laws, or theories.
  - Can you think of a way to test whether the sweet taste of soda pop is due to the presence of sugar?
- Results either validate (confirm) or invalidate (deny) your ideas.
  - Invalidate = Discard or Modify
    - Many times experiments invalidate only parts of the hypothesis or theory, in which case the idea is modified.
  - Validate ≠ Proof your idea will always hold

Laws

- Summary of observations that combines all past observations into one general statement.
  - Allows you to predict future observations.
  - Law of Conservation of Mass— “In a chemical reaction matter is neither created nor destroyed.”
- What’s the Difference Between an Observation and a Law?
  - An observation tells you what happened in a single event.

- A law summarizes all the observations, effectively telling you what you will observe in future events.

Theories

- General explanation for the characteristics and behavior of nature.
  - Models of nature.
  - Ex. Dalton's Atomic Theory, Theory of Gravity, Germ Theory
  - Can be used to predict future observations.
- What's the Difference Between a Hypothesis and a Theory?
  - A hypothesis is an explanation of a single or small number of observations.
  - A theory is an explanation that extends beyond individual observations to an understanding of the underlying causes for the way nature is or behaves.

What's the Difference Between a Law and a Theory?

- Laws answer the question "What" will happen.

- Theories tell us "What" will happen but also “Why” it happens.
    - Theories allow to extend your predictions to a wider set of circumstances.

	Applies to single or **small number** of events	Applies to **all** events
Describes ***what*** happens	observation	law
Explains ***why*** things happen	hypothesis	theory

Example from History

Why Do Some Things Burn?

Observations

- Things would stop burning when placed in a closed container.
- Many metals burn to form a white powder called calx.
- Metals can be recovered from their calx by roasting it with charcoal.

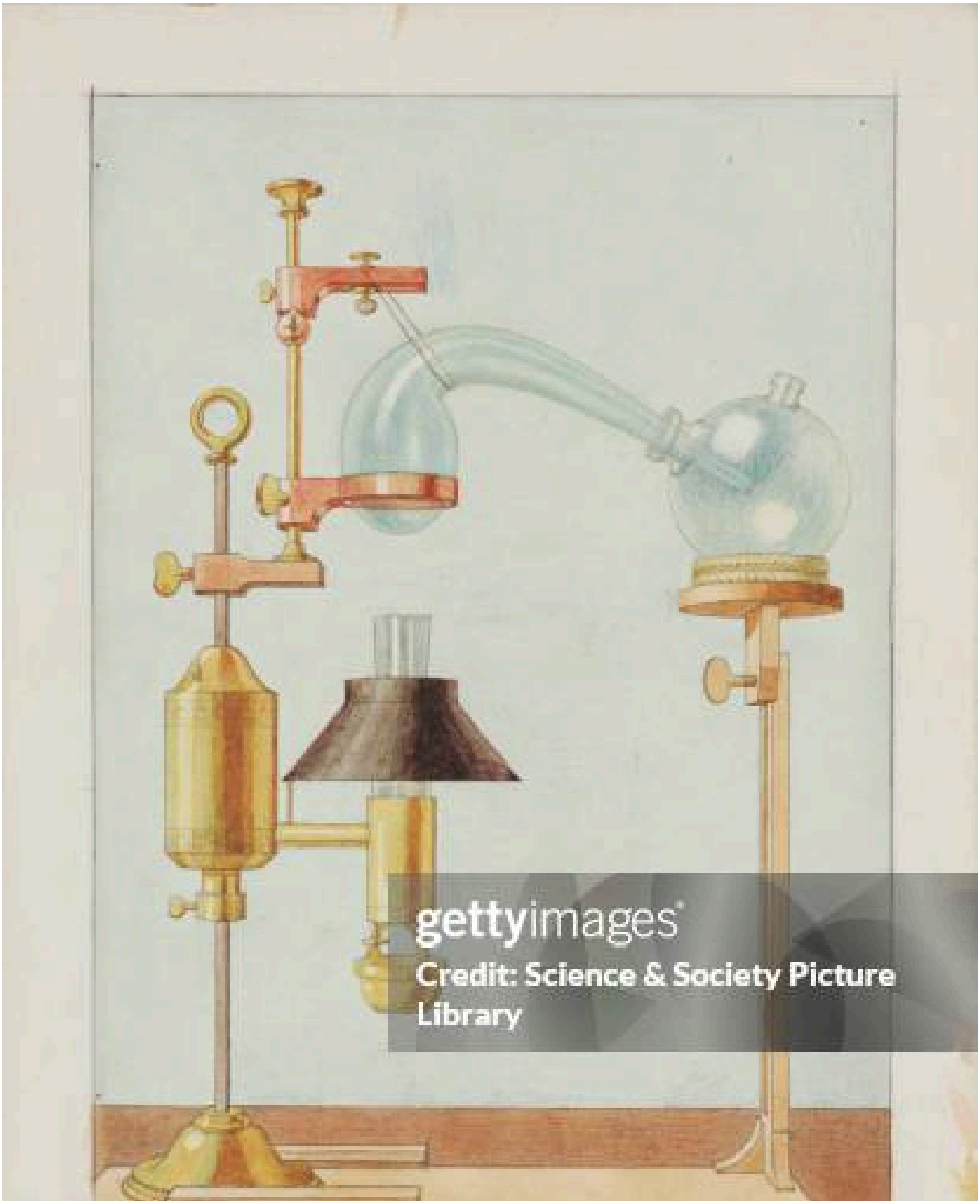
Hypothesis

- Phlogiston Theory is an Explanation of combustion proposed in early/mid-1700s.
  - Combustible substances contained a substance they called phlogiston.
  - When a substance burned it released all or some of its phlogiston into the air .
- How Does Phlogiston Theory Explain the Observations?
  - When a substance is burned in the open, all the phlogiston is released.
  - When a substance is burned in a closed container, the phlogiston is released until it saturates the container, at which point the combustion stops.
  - A metal's calx is what is left after it releases all its phlogiston.
  - When roasted with charcoal the calx reacquires phlogiston from the charcoal.
    - Charcoal is rich in phlogiston, that's why charcoal burns.

Experiment

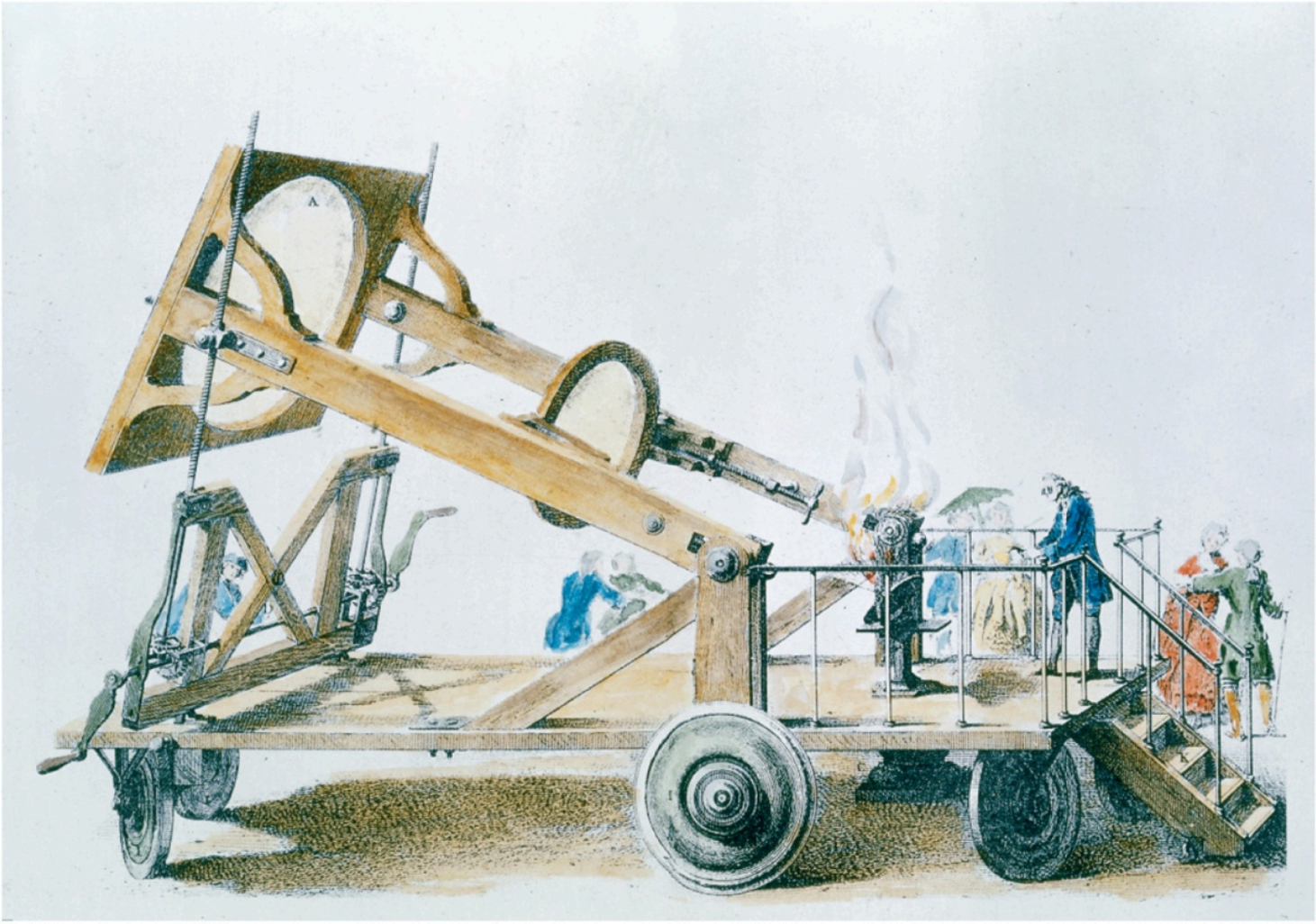
If phlogiston is lost when metals burn, then the metals should lose weight when burned.

- Morveau's experiments showed that when a piece of metal burned, the resulting calx weighed more than the original metal.
    - Do Morveau's observations validate or invalidate the Phlogiston Theory?


gettyimages
Credit: Science & Society Picture Library

If a calx is heated, it should remove phlogiston from the air as the calx is converted to the metal.

- Lavoisier roasted many calx with a large lens and observed that material he called “fixed air” was released into the air.
    - Do Lavoisier’s observations validate or invalidate the Phlogiston Theory?



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A Better Theory of Combustion

- Lavoisier proposed an alternative theory of combustion.
  - When materials burn, they remove and combine with "fixed air" from the air.
- Does Lavoisier’s idea explain all the previous observations?
- How could you test Lavoisier’s idea?

How to Succeed in Chemistry

- Be curious and use your imagination.
- Explore and investigate.
- Quantify and calculate
  - Even small differences can be important!
- Commitment
  - Work regularly and carefully.

Chapter 1 Quiz

1. Chemicals make up everything around you, including your own body. (1 point)

   True False

2. Chemists study the properties of substances and the particles that compose those substances. (1 point)

   True False

3. The properties of water do not depend on how the atoms are bonded together in the molecule. (1 point)

   True False

4. Laws are stronger statements than theories. (1 point)

   True False

5. The scientific method emphasizes logic and debate as the way to understand the world. (1 point)

   True False

6. A hypothesis can be proven wrong. (1 point)

True False

7. The first step in acquiring scientific knowledge is often the1 observation or measurement of some aspect of nature. (1 point)

True False

8. Which of the statements below is NOT part of the scientific method? (2 point)

A) observation and measurement

B) formation of a hypothesis

C) testing of a hypothesis by experimentation

D) refinement of a hypothesis as needed

E) All of the above steps are part of the scientific method.

9. What term best describes a brief statement that summarizes many past observations and predicts new ones? (2 point)

A) experiment

B) hypothesis

C) scientific law

D) theory

E) none of the above

10. A sample of wood with a mass of 3.0 grams was burned in an open dish. The ashes weighed 1.2 grams.

What happened to the rest of the wood? (2 point)

A) The remaining mass was destroyed.

B) The remaining 1.8 grams was converted into gaseous compounds.

C) Nothing, the burned wood simply has a mass of 1.2 grams.

D) The remaining 1.8 grams was converted into heat.

E) none of the above

11. Which statement accurately describes the purpose of experiments? (2 point)

A) Experiments are designed to produce the results predicted by a theory.

B) Experiments can be replaced by a simple, logical reasoning of known facts.

C) Experiments are designed to produce unexplainable results for further investigation.

D) Experiments are designed to validate or invalidate a hypothesis, law, or theory.

E) none of the above

Chapter 2

What Is a Measurement?

- Quantitative observation.
- Comparison to an agreed upon standard.
- Every measurement has a number, a unit, and an indicated degree of uncertainty.
- The unit tells you to what standard you are comparing your object.
- The number tells you:
  - What multiple of the standard the object measures.
  - The uncertainty in the measurement.

Scientists have measured the average global temperature rise over the past century to be 0.6 °C

- **°C** tells you that the temperature is being compared to the Celsius temperature scale.
- **0.6** tells you that:

- The average temperature rise is 0.6 times the standard unit of 1 degree Celsius.
- The confidence in the measurement is such that we are certain the measurement is between 0.5 and 0.7 °C.

Scientific Notation

- We commonly measure objects that are many times larger or smaller than our standard of comparison.
- Writing large numbers of zeros is tricky and confusing.
  - Not to mention there's the 8-digit limit of your calculator

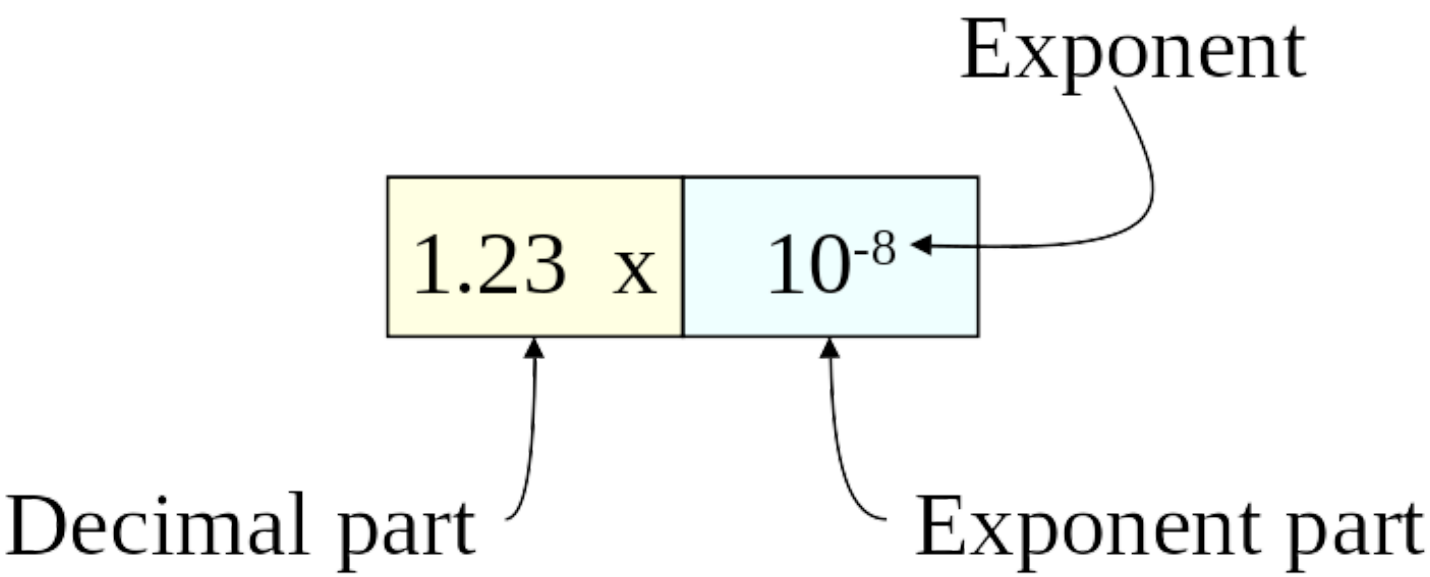
The sun’s diameter is 1,392,000,000 m.

An atom’s average diameter is 0.000 000 000 3 m.

- Each decimal place in our number system represents a different power of 10.
- Scientific notation writes the numbers so they are easily comparable by looking at the power of 10.

The sun's diameter is 1.392×10^9 m.

An atom's average diameter is 3×10^{-10} m.



Exponents

- When the exponent on 10 is positive, it means the number is that many powers of 10 larger.
  - Sun's diameter = 1.392×10^9 m = 1.392×1000000000 m = 1,392,000,000 m.
- When the exponent on 10 is negative, it means the number is that many powers of 10 smaller.
  - Average atom's diameter = 3×10^{-10} m = 3×0.0000000001 m = 0.0000000003 m.
- To compare numbers written in scientific notation:
  - First compare exponents on 10.
  - If exponents are equal, then compare decimal numbers

$$1.23 \text{ x } 10^5 > 4.56 \text{ x } 10^2$$

$$4.56 \text{ x } 10^{-2} > 7.89 \text{ x } 10^{-5}$$

$$7.89 \text{ x } 10^{10} > 1.23 \text{ x } 10^{10}$$

Writing Numbers in Scientific Notation

1. Locate the decimal point.
2. Move the decimal point to obtain a number between 1 and 10.
3. Multiply the new number by 10^n. Where n is the number of places you moved the decimal point.
4. If you moved the decimal point to the left, then n is positive; if you moved it to the right, then n is negative.

- If the original number is 1 or larger, then n is positive .
- If the original number is less than 1, then n is negative .

Example

12340

1. Locate the decimal point.
   12340.
2. Move the decimal point to obtain a number between 1 and 10.
   1.234
3. Multiply the new number by 10^n . Where n is the number of places you moved the decimal point.
   1.234×10^4

4. If you moved the decimal point to the left, then n is positive; if you moved it to the right, then n is negative .
   1.234×10^4

Example

1. 0.00012340
2. Locate the decimal point.
   0.00012340
3. Move the decimal point to obtain a number between 1 and 10.
   1.2340
4. Multiply the new number by 10^n . Where *n* is the number of places you moved the decimal point.
   1.2340×10^4
5. If you moved the decimal point to the left, then *n* is postive; if you moved it to the right, then *n* is negative .
   1.2340×10^{-4}

Writing a Number in Standard Form

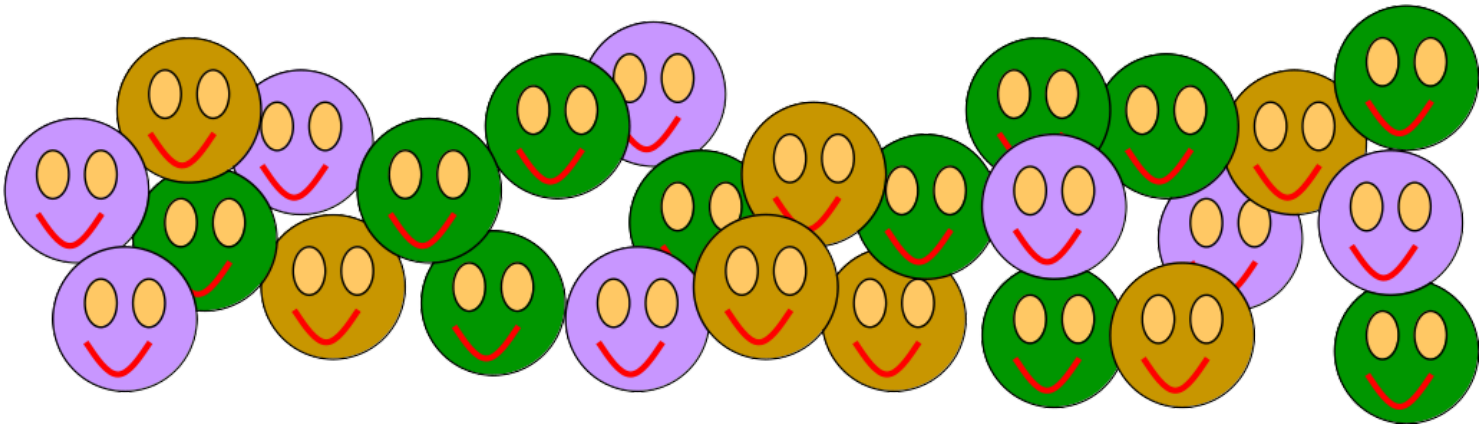
1.234×10^{-6}

- Since the exponent is -6, make the number smaller by moving the decimal point to the left 6 places.
    - When you run out of digits to move around, add zeros.
    - Add a zero in front of the decimal point for decimal numbers.

000 001,234

0.000 001 234

The U.S. population in 2007 was estimated to be 301,786,000 people. Express this number in scientific notation.



Write the Following Numbers in Scientific Notation

123.4

8.0012

145000

0.00234

25.25

0.0123

1.45

0.000008706

Write the Following Numbers in Standard Form

2.1×10^{3}

4.02×10^{0}

9.66×10^{-4}

3.3×10^{1}

6.04×10^{-2}

1.2×10^{5}

Inputting Scientific Notation into a Calculator

We're going to practice inputting the following into your calculator.

- The number -1.23×10^{-3}
- $\frac{-1.23\times10^{-3}}{4.78\times10^{5}}$
- $(-1.37 \times 10^{-4})(5.92 \times 10^{2})$

Significant Figures

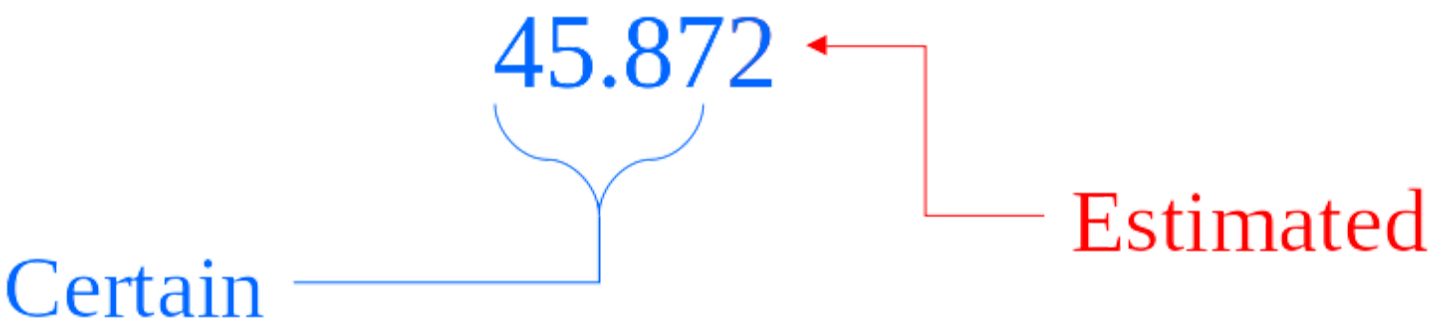
Writing numbers to reflect precision.

Exact Numbers vs. Measurements

- Sometimes you can determine an exact value for a quality of an object.
  - Often by counting.
    - Pennies in a pile.
  - Sometimes by definition
    - 1 ounce is exactly 1/16 pounds.
- Whenever you use an instrument to compare a quality of an object to a standard, there is uncertainty in the comparison.

Reporting Measurements

- Measurements are written to indicate the uncertainty in the measurement.
- The system of writing measurements we use is called **significant figures**.
- When writing measurements, all the digits written are known with certainty except the last one, which is an estimate.



Estimating the Last Digit

- For instruments marked with a scale, you get the last digit by estimating between the marks.
    - If possible.
- Mentally divide the space into 10 equal spaces, then estimate how many spaces over the indicator is.

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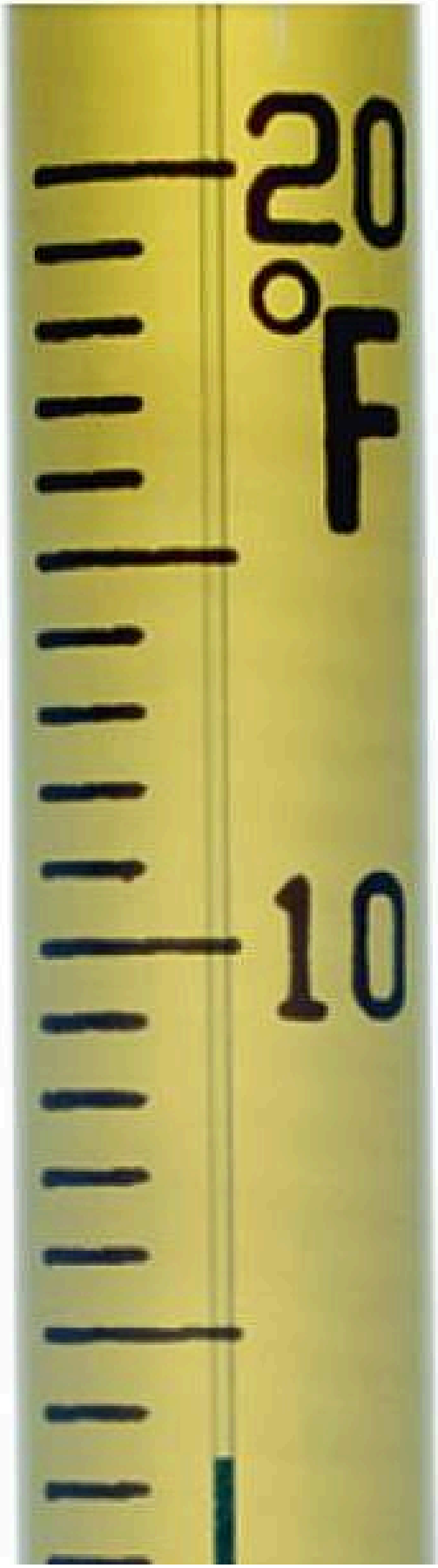
1.2 grams
the “1” is certain;
the “2” is an estimate.

What is the temperature reading on the thermometer to the correct number of digits?


20
°F
10

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Which Digits are Significant?

- The non-placeholding digits in a reported measurement are called significant figures.
  - Some zeros in a written number are only there to help you locate the decimal point.
- Significant figures tell us the range of values to expect for repeated measurements.
  - The more significant figures there are in a measurement, the smaller the range of values. Therefore, the measurement is more precise.

12.3 cm
has 3 significant figures
and its range is
12.2 to 12.4 cm.

12.30 cm
has 4 significant figures
and its range is
12.29 to 12.31 cm.

Counting Significant Figures

- All non-zero digits are significant.
  - 1.5 has 2 significant figures.

- Interior zeros are significant.
  - 1.05 has 3 significant figures.
- Trailing zeros after a decimal point are significant.
  - 1.050 has 4 significant figures.
- Leading zeros are **NOT** significant.
  - 0.001050 has 4 significant figures.
- Zeros at the end of a number without a written decimal point are ambiguous and should be avoided by using scientific notation.
  - If 150 has 2 significant figures, then 1.5×10^2, but if 150 has 3 significant figures, then 1.50×10^2.

Scientific Numbers are only written with Significant Digits. This is how you avoid ambiguity

Significant Figures and Exact Numbers

- Exact numbers have an unlimited number of significant figures.
- A number whose value is known with complete certainty is exact.
  - From counting individual objects.
  - From definitions.

    - 1 cm is exactly equal to 0.01 m.
- From integer values in equations.
    - In the equation for the radius of a circle, the 2 is exact.

$$\bar{r} = \frac{Diameter}{2}$$

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

0.0035

1.080

2371

2.97×10^5

1 dozen = 12

100,000

Determine the Number of Significant Figures, the Expected Range of Precision, and Indicate the Last Significant Figure

12000

120.

12.00

1.20×10^3

0.0012

0.00120

1201

1201000

Rounding

- When rounding to the correct number of significant figures, if the number after the place of the last significant figure is:
  - 0 to 4, round down.
    - Drop all digits after the last significant figure and leave the last significant figure alone.
    - Add insignificant zeros to keep the value, if necessary.
  - 5 to 9, round up.
    - Drop all digits after the last significat figure and increase the last significant figure by one.
    - Add insignificant zeros to keep the value, if necessary.
      Example
- Rounding to 2 significant figures.
  - 2.34 rounds to 2.3
  - 2.37 rounds to 2.4
  - 2.349865 rounds to 2.3
  - 0.0234 rounds to 0.023 or 2.3×10^{-2}
  - 0.0237 rounds to 0.024 or 2.4×10^{-2}

- 0.02349865 rounds to 0.023 or 2.3×10^{-2}
- 234 rounds to 230 or 2.3×10^{2}
- 237 rounds to 240 or 2.4×10^{2}
- 234.9865 rounds to 230 or 2.3×10^{2}

Multiplication and Division with Significant Figures

When multiplying or dividing measurements with significant figures, the result has the same number of significant figures as the measurement with the fewest number of significant figures.

$$\underset{\text{3 sig. figs.}}{5.02} \times \underset{\text{5 sig. figs.}}{89{,}665} \times \underset{\text{2 sig. figs.}}{0.10} = 45.0118 = \underset{\text{2 sig. figs.}}{45}$$

$$\underset{\text{4 sig. figs.}}{5.892} \div \underset{\text{3 sig. figs.}}{6.10} = 0.96590 = \underset{\text{3 sig. figs.}}{0.966}$$

Determine the Correct Number of Significant Figures for Each Calculation. Round and Report the Result.

$$\frac{1.01 \times 0.12 \times 53.51}{96} =$$

$$\frac{56.55 \times 0.920}{34.2585} =$$

Addition and Subtraction with Significant Figures

When adding or subtracting measurements with significant figures, the result has the same number of decimal places as the measurement with the fewest number of decimal places.

5.74	+	0.823	+	2.651	=	9.214	=	9.21
2 dec. pl.		3 dec. pl.		3 dec. pl.				2 dec. pl.
4.8	-	3.965			=	0.835	=	0.8
1 dec. pl		3 dec. pl.						1 dec. pl.

Determine the Correct Number of Significant Figures for Each Calculation. Round and Report the Result.

$$0.987 + 125.1 - 1.22 =$$

$$0.764 - 3.449 - 5.98 =$$

Both Multiplication/Division and Addition/Subtraction with Significant Figures

- When doing different kinds of operations with measurements with significant figures, evaluate the significant figures in the intermediate answer, then do the remaining steps.
- Follow the standard order of operations.
  - Please Excuse My Dear Aunt Sally.

    $()\rightarrow^{n}\rightarrow\times\div\rightarrow+\ -$

$$3.489 \times (5.67 - 2.3) =$$

2 dp 1 dp

$$3.489 \times 3.\underline{3}7 = 12$$

4 sf 1 dp & 2 sf 2 sf

Perform the Following Calculations to the Correct Number of Significant Figures

$1.10 \times 0.5120 \times 4.0015 - 3.4555 =$

$\frac{4.562 \times 3.99870}{452.6755 - 452.33} =$

$(14.84 \times 0.55) - 8.02 =$

Units

- Units tell the standard quantity to which we are comparing the measured property.
  - Without an associated unit, a measurement is without meaning.
- Scientists use a set of standard units for comparing all our measurements.
  - So we can easily compare our results.
- Each of the units is defined as precisely as possible.
- Scientists generally report results in an agreed upon International System.
- The SI System
  - Système International

Quantity	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K

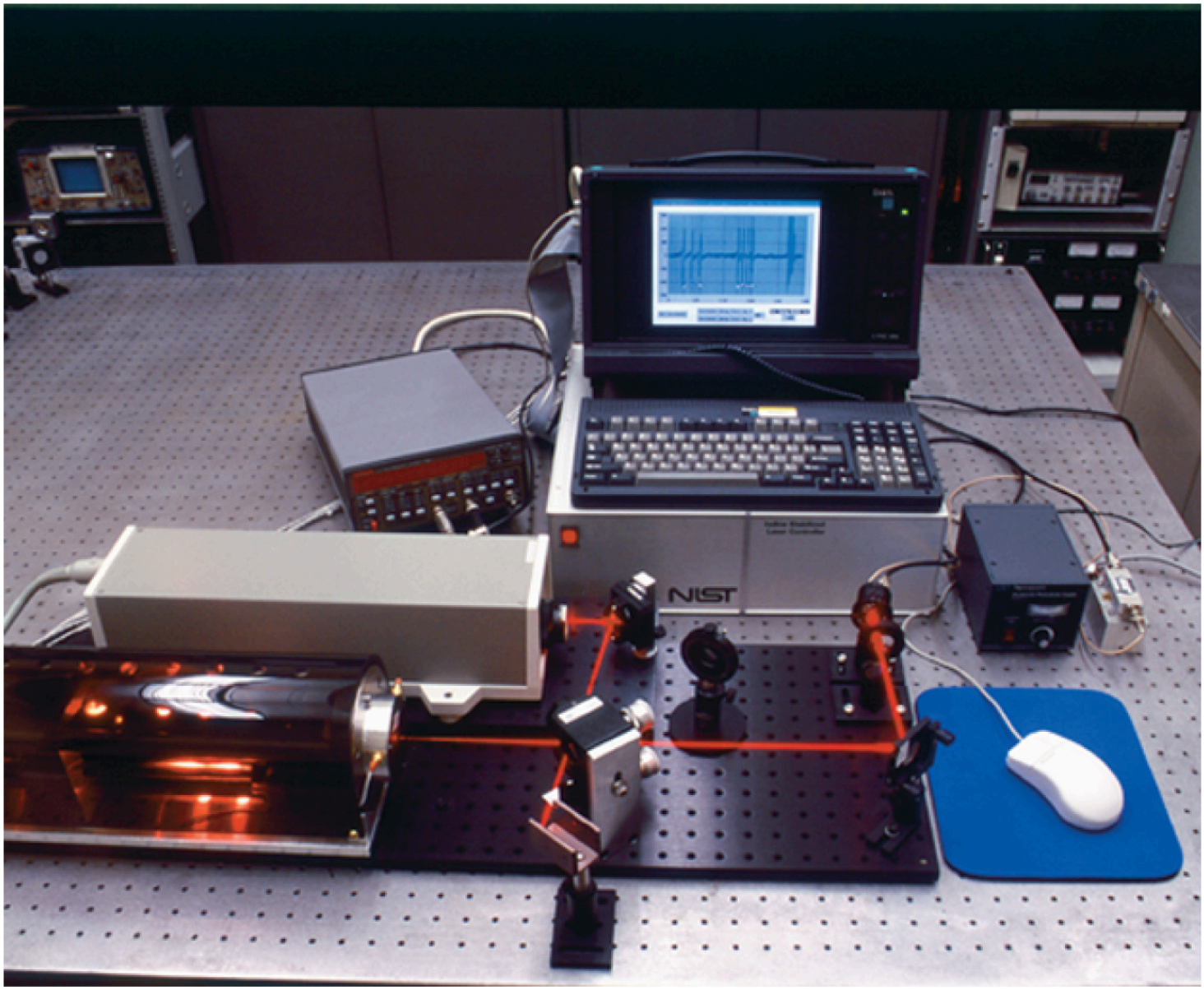
Quantity	Unit	Symbol
Volume	liter (litre)	L

Length

- Measure of the two-dimensional distance an object covers.
- The SI unit for length is a meter, about 3½ inches longer than a yard.
  - One ten-millionth the distance from the North Pole to the Equator
  - Distance between marks on standard metal rod in a Paris vault
  - Distance covered by a certain number of wavelengths of a special color of light

Yardstick

Meterstick



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Mass

- Measure of the amount of matter present in an object.
- The SI unit is the kilogram (kg), about 2 lbs. 3 oz. It's important to note the base unit is **not** the gram.

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Time

- Measure of the duration of an event.
- The SI units is the second (s)
  - 1 s is defined as the period of time it takes for a specific number of radiation events of a specific transition from cesium-133.

Temperature

- Measure of the average amount of kinetic energy.
    - The higher the temperature the greater the average kinetic energy
- Heat flows from the matter that has high thermal energy into matter that has low thermal energy.

    - Until they reach the same temperature.
- Heat is exchanged through molecular collisions between the two materials.
  Related Units in the

SI System

- All units in the SI system are related to the standard unit by a power of 10.
- The power of 10 is indicated by a prefix.
- The prefixes are always the same, regardless of the standard unit.
- It is usually best to measure a property in a unit close to the size of the property.
    - It reduces the number of confusing zeros.
      Common Prefixes in the

Prefix	Symbol	Decimal Equivalent	Power of 10
mega-	M	1,000,000	Base x 10^{6}
kilo-	k	1,000	Base x 10^{3}
deci-	d	0.1	Base x 10^{-1}
centi-	c	0.01	Base x 10^{-2}
milli-	m	0.001	Base x 10^{-3}
micro-	μ or mc	0.000 001	Base x 10^{-6}
nano-	n	0.000 000 001	Base x 10^{-9}

Metric Units												
mega (M)	*	*	kilo (k)	hector (h)	deka (da)	unit (m, g, L)	deci (d)	centi (c)	milli (m)	*	*	micro (mc) (u)
When going from larger unit to smaller unit move decimal to the right →												
When going from smaller unit to larger unit move decimal to the left ←												

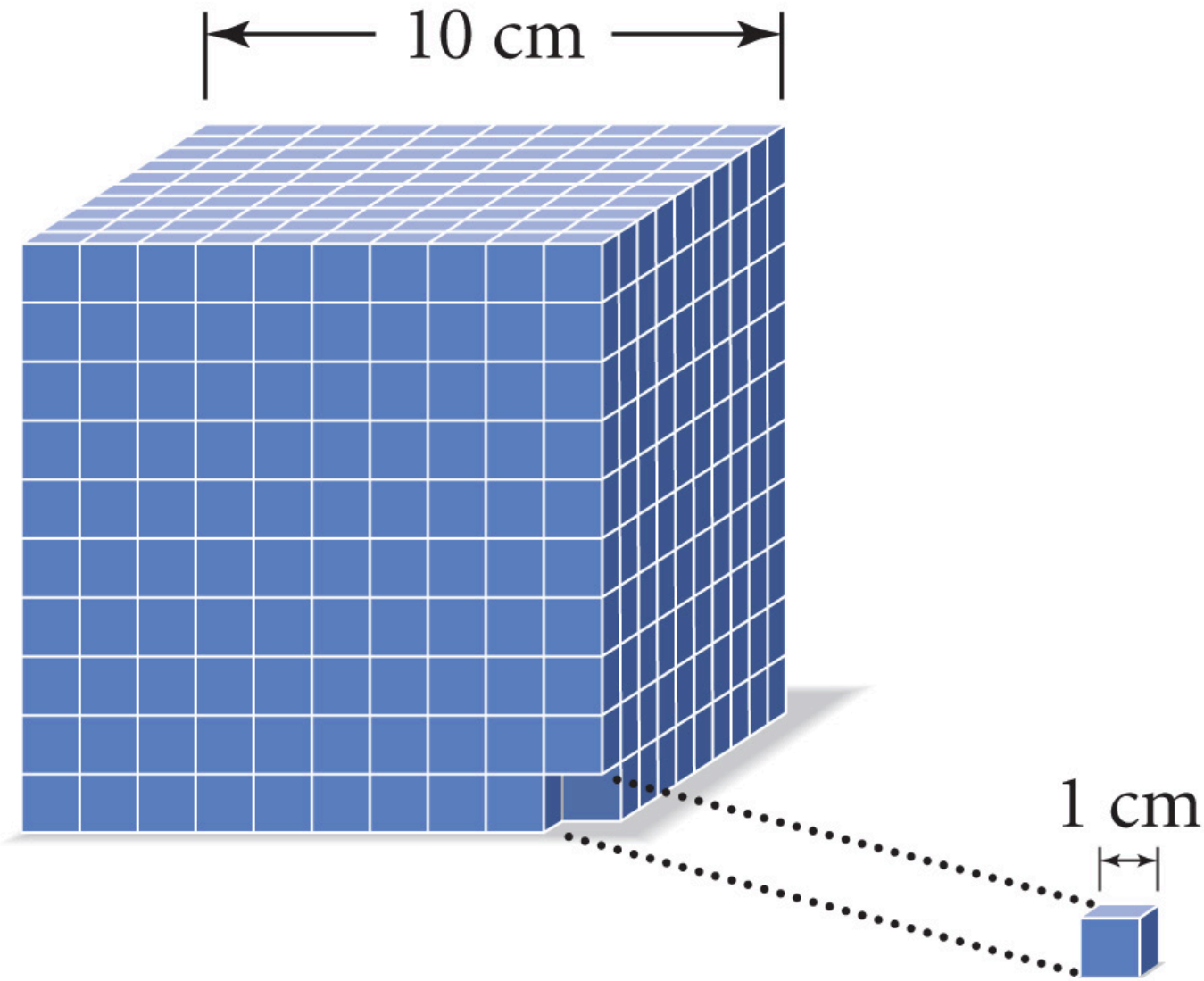
Which of the Following Units Would Be Best Used for Measuring the Diameter of a Quarter?

- kilometer
- meter

- centimeter
- micrometer
- megameters

Volume

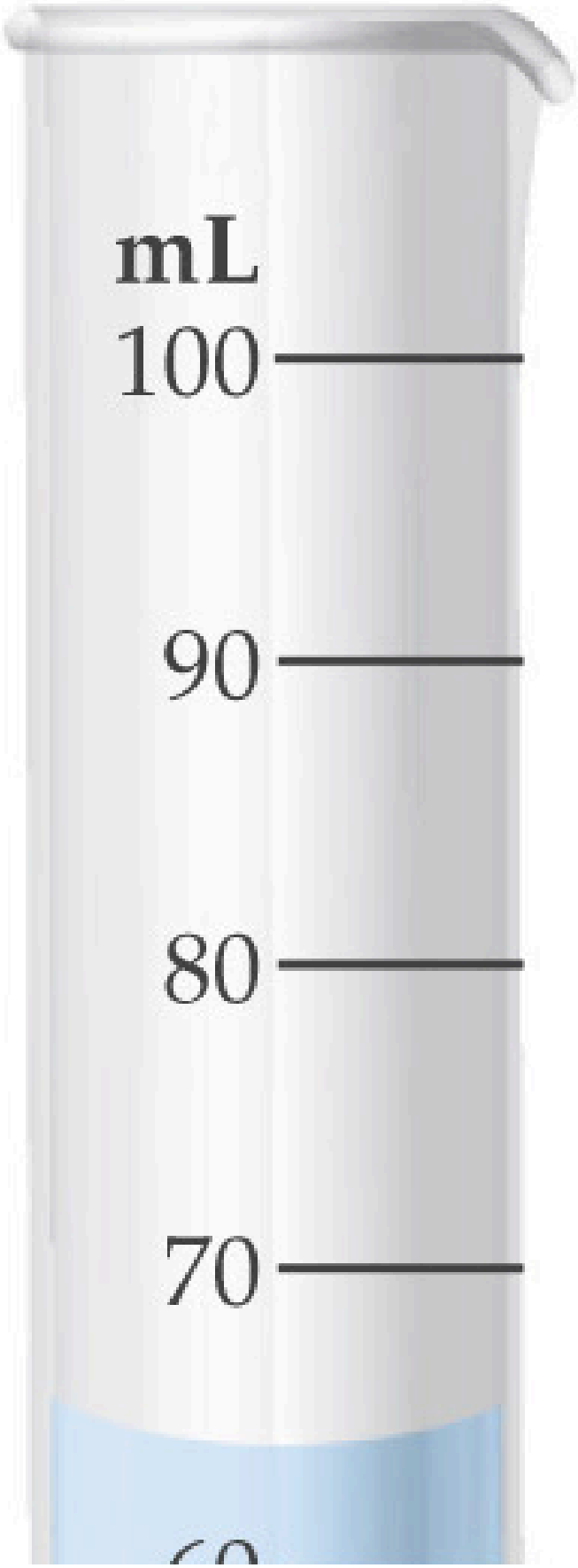
- Derived unit.
    - Any length unit cubed.

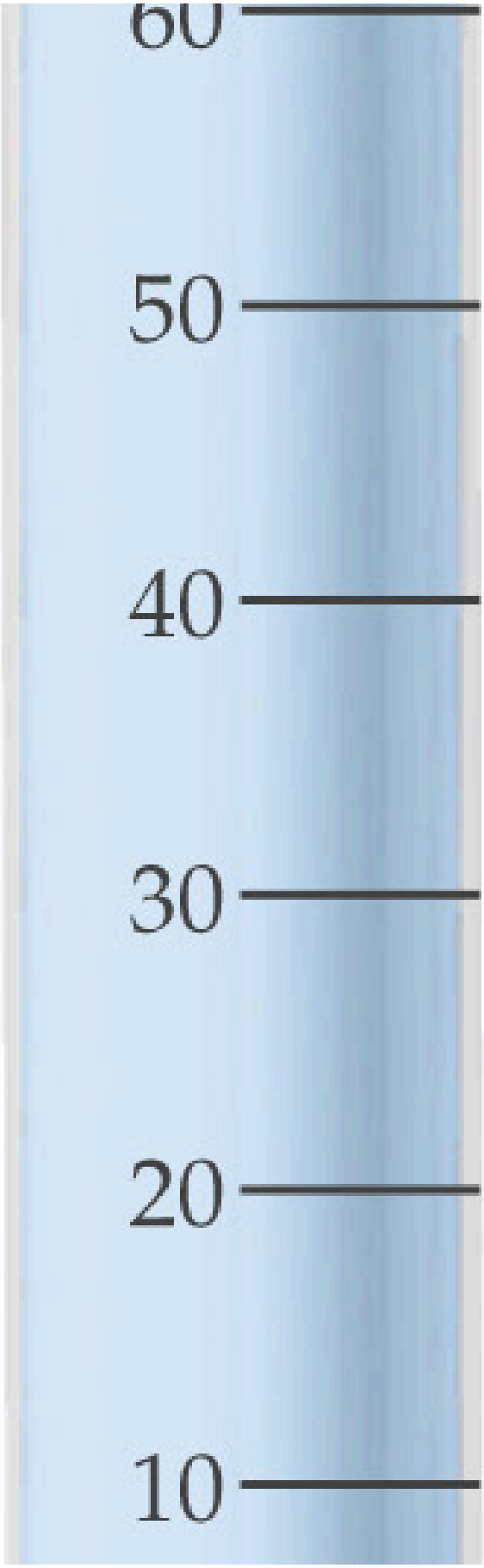


A 10-cm cube contains 1000 1-cm cubes.

- Measure of the amount of space occupied.
- SI unit = cubic meter (m^3)
- Commonly measure liquid or gas volume in milliliters (mL).
  - 1 L is slightly larger than 1 quart.

- 1 L = 1 dm^3 = 1000 ml = 10^3 mL
- 1 ml = 0.001 L = 10^{-3} L
- 1 mL = 1 cm^3


mL
100
90
80
70


60
50
40
30
20
10

Which Is Larger?

- 1 yard or 1 meter?
- 1 mile or 1 km?
- 1 cm or 1 inch?
- 1 kg or 1 lb?
- 1 mg or 1 μg?
- 1 qt or 1 L?
- 1 L or 1 gal?
- 1 gal or 1000 cm^3?

How to use Units

- Always write every number with its associated unit.
- Always include units in your calculations.

- You can do the same kind of operations on units as you can with numbers.
  - cm × cm = cm^2
  - cm + cm = cm
  - cm ÷ cm = 1
- You can use units as a guide to problem solving
  - This is called dimensional analysis.

Conversions

- Many problems in chemistry involve using relationships to convert one unit of measurement to another.
- Conversion factors are relationships between two units.
  - May be exact or measured.
  - Both parts of the conversion factor have the same number of significant figures.
- Conversion factors generated from equivalence statements.

$$1 \text{ inch} = 2.54 \text{ cm}$$

$$\frac{2.54 \text{ cm}}{1 \text{ in}} \qquad \frac{1 \text{ in}}{2.54 \text{ cm}}$$

- Arrange conversion factors so the starting unit cancels.
  - The starting unit should in the denominator of the conversion factor.

$$2 \cancel{\text{ hr}} \left(\frac{30 \text{ mi}}{1 \cancel{\text{ hr}}} \right) = 60 \text{ miles}$$

- May string conversion factors.
  - So we do not need to know every relationship, as long as we can find something else the starting and desired units are related to :

$$5 \cancel{\text{ cups}} \left(\frac{1 \cancel{\text{ pint}}}{2 \cancel{\text{ cups}}} \right) \left(\frac{1 \text{ quart}}{2 \cancel{\text{ pints}}} \right) = 1.25 \text{ quarts}$$

Convert 7.8 km to miles

Convert 30.0 g to Ounces

An Italian recipe for making creamy pasta sauce calls for 0.75 L of cream. Your measuring cup measures only in cups. How many cups should you use?

Convert 30.0 mL to Quarts

Convert 2,659 cm^2 to Square Meters

Mass and Volume

- Two main characteristics of matter.
- Cannot be used to identify what type of matter something is.
  - If you are given a large glass containing 100 g of a clear, colorless liquid and a small glass containing 25 g of a clear, colorless liquid, are both liquids the same stuff?
- Even though mass and volume are individual properties, for a given type of matter they are related to each other!

Density

- Ratio of mass to volume.

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

- Its value depends on the kind of material, not the amount.
  - Solids = g/cm^3 (g/ml)
  - Liquids = g/mL

    - Gases = g/L
- Volume of a solid can be determined by water displacement—Archimedes Principle.
- Density : solids > liquids > gases
    - Water is an exception
- When volumes are equal, the more dense substance will be heavier .
- When the mass of two samples is equal, the more dense substance will have smaller volume.
- Heating causes objects causes objects to expand. Lowering their density.
    - Volume will increase
    - Mass will remain the same
- In a heterogeneous mixture, the more dense object sinks.

Solve the density equation for mass and volume.

Platinum has become a popular metal for fine jewelry. A man gives a woman an engagement ring and tells her that it is made of platinum. Noting that the ring felt a little light, the woman decides to perform a test to determine the ring's density before giving him an answer about marriage. She places the ring on a balance and finds it has a mass of 5.84 grams. She then finds that the ring displaces 0.556 cm^3 of water. Is the ring made of platinum? (Density Pt = 21.4 g/cm^3)

What Is the Density of Metal if a 100.0 g Sample Added to a Cylinder of Water Causes the Water Level to Rise from 25.0 mL to 37.8 mL?

How much does 4.0 cm^3 of lead (11.3 g/cm^3) weigh?

The gasoline in an automobile gas tank has a mass of 60.0 kg and a density of 0.752 g/cm^3. What is the volume?

A 55.9 kg person displaces 57.2 L of water when submerged in a water tank. What is the density of the person in g/cm^3?

Chapter 2 Quiz

1. Numbers are usually written so that the uncertainty is in the last reported digit. (1 Point)

   True False

2. If you count 7 pennies, you can only report one significant figure in that measurement. (1 Point)

   True False

3. Measurements have an unlimited number of significant figures. (1 Point)

   True False

4. Zeros located between two numbers are not significant. (1 Point)

   True False

5. A trailing zero after a decimal point is significant. (1 Point)

   True False

6. Trailing zeros at the end of a number, but before an implied decimal point, are ambiguous. (1 Point)

True False

7. In multiplication and division calculations, the answer will have the same number of decimal places as the number carrying the fewest decimal places. (1 Point)

True False

8. In addition and subtraction calculations, the answer will have the same number of significant digits as the number carrying the most decimal places. (1 Point)

True False

9. In addition or subtraction, the result carries the same number of decimal places as the quantity carrying the fewest decimal places. (1 Point)

True False

10. Scientific numbers are reported so that every digit is certain except the last, which is estimated. (1 Point)

True False

11. The mass of an object depends on gravity. (1 Point)

True False

12. The correct scientific notation for the number 0.07805 is: (2 Point)

A) 7.805×10^{2}

B) 7.805 × 102

C) 7.805×10^{-2}

D) 0.7805×10^{-2}

E) none of the above

13. The correct scientific notation for the number 500.0 is: (2 Point)

A) 5×10^{2}

B) 5.00×10^{2}

C) 5.000×10^{2}

D) 5×10^{-2}

E) none of the above

14. The correct decimal representation of 1.709×10^{-5} is: (2 Point)

A) 170900

B) 0.0001709

C) 0.00001709

D) 1709

E) none of the above

15. The correct decimal representation of 6.453×10^{3} is: (2 Point)

A) 6,453

B) 0.006453

C) 6.5×10^{3}

D) 6.453

E) none of the above

16. The correct number of significant figures in the number 12501 is: (2 Point)

A) 5

B) 7

C) 4

D) ambiguous

E) none of the above

17. The correct number of significant figures in the number 9.08×10^{-4} is (2 Point)

A) 3

B) 4

C) 5

D) ambiguous

E) none of the above

18. When the value 4.449 is rounded to two significant figures, the number should be reported as: (2 Point)

A) 4.4

B) 4.5

C) 4.44

D) 4.45

E) none of the above

19. Determine the answer for the equation below with correct number of significant figures:

3.2 × 1.729 = ____ (2 Point)

A) 5.5328

B) 5.533

C) 5.54

D) 5.5

E) none of the above

20. Determine the answer to the following equation with correct number of significant figures:

2.02 + 8.102 - 0.0297 = ___ (2 Point)

A) 10.0923

B) 10.09

C) 10.1

D) 10.092

E) none of the above

21. The correct prefix for the multiplier 1,000,000 is: (2 Point)

A) mega (M)

B) milli (m)

C) micro (μ)

D) nano (n)

E) none of the above

22. What is the standard SI unit for mass? (2 Point)

A) kilogram

B) gram

C) pound

D) ton

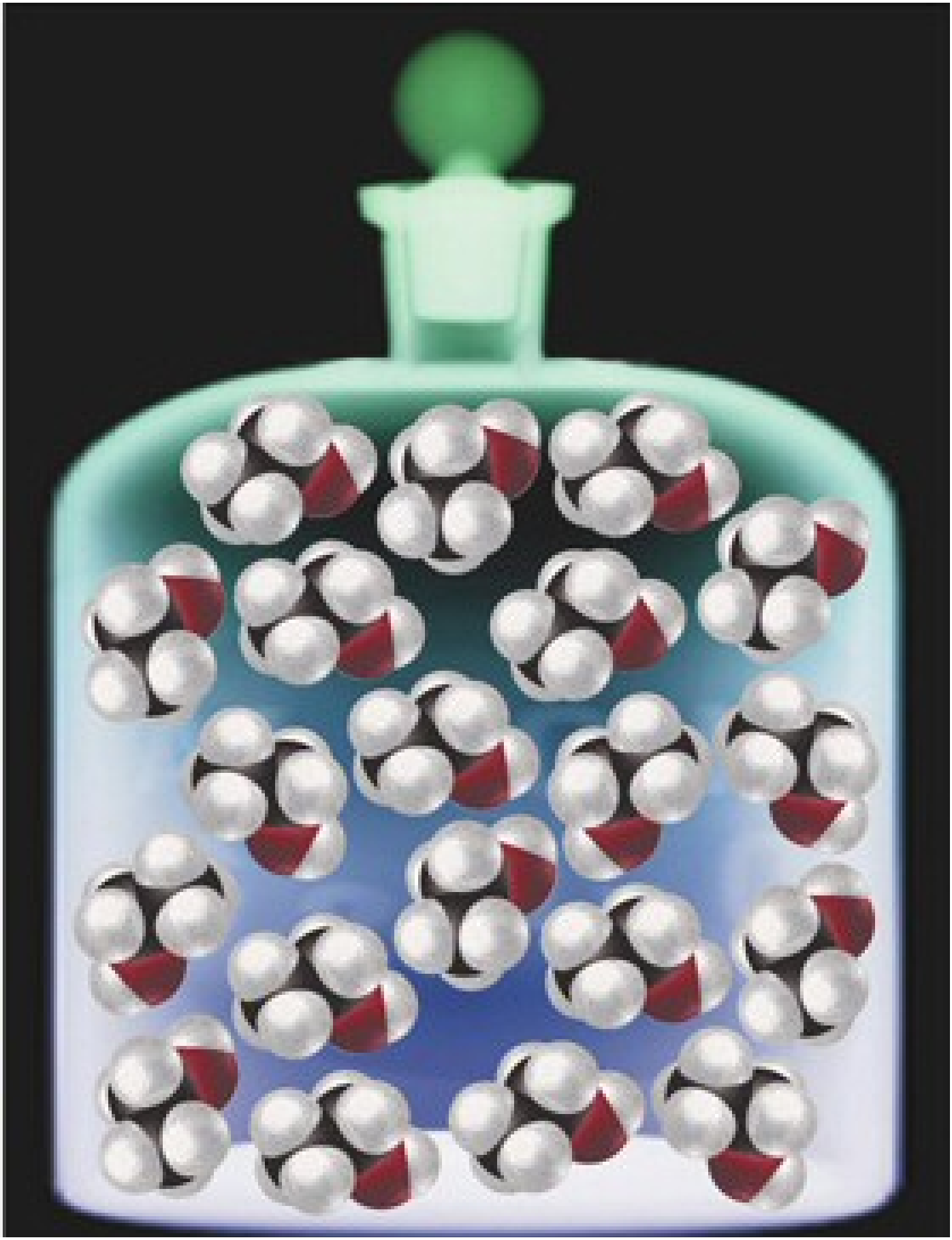
E) none of the above

Chapter 3

What is Matter?

- Matter is defined as anything that occupies space and has mass
- Even though it appears to be smooth and continuous, matter is actually composed of a lot of tiny little pieces we call **atoms** and **molecules**
- Atoms are the tiny particles that make up all matter.

- In most substances, the atoms are joined together in units called molecules

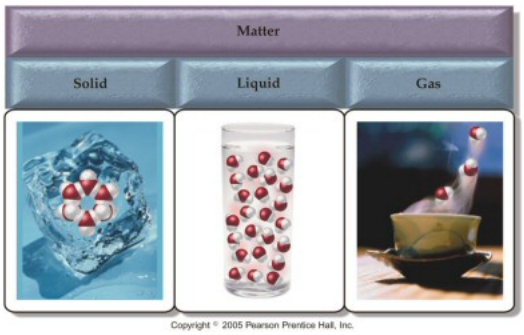


Classifying Matter by Physical State

- Matter can be classified as solid, liquid or gas based on what properties it exhibits

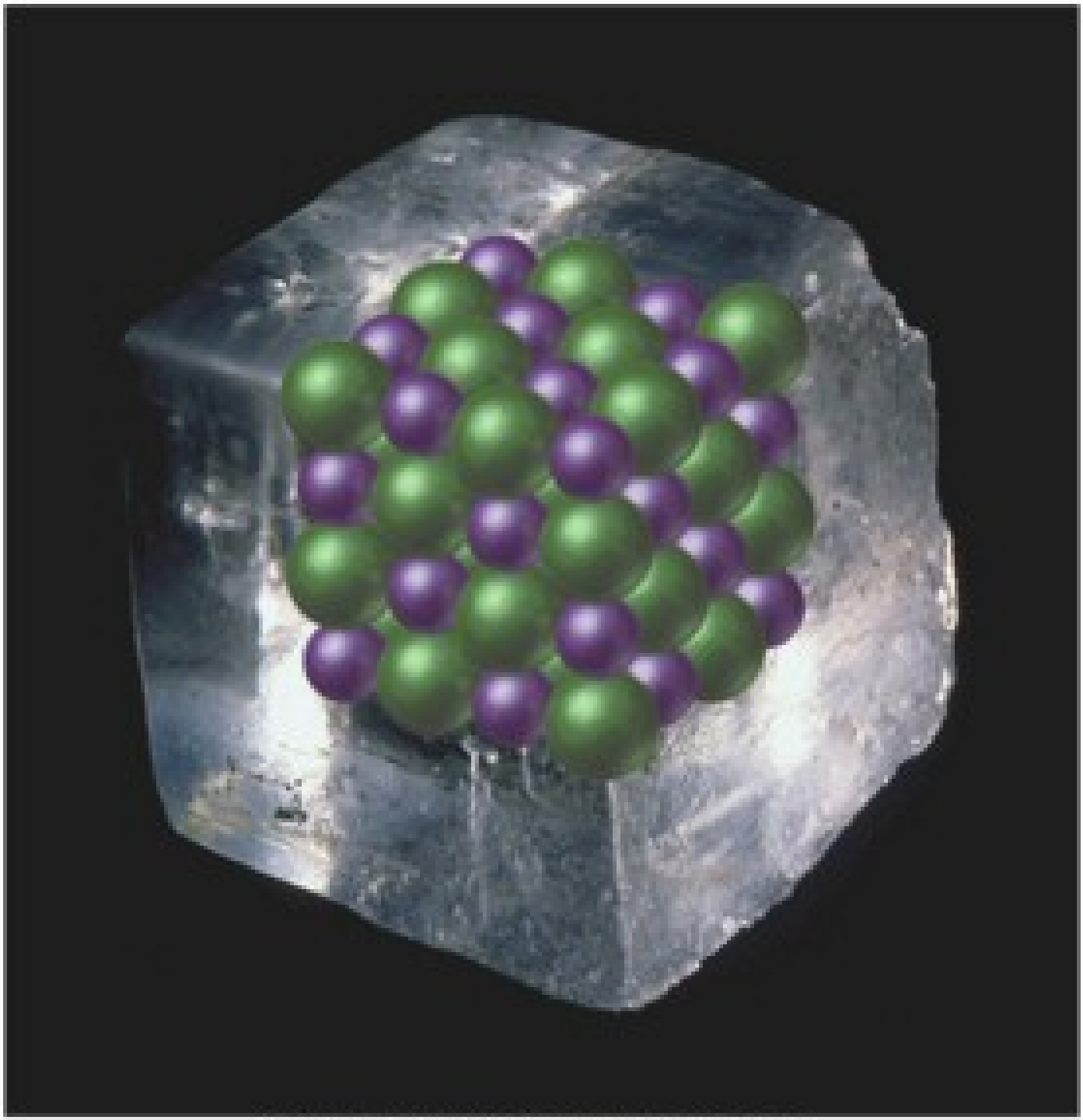
Phase	Shape	Volume	Compress	Flow
Solid	Fixed	Fixed	No	No
Liquid	Indefinite	Fixed	No	Yes
Gas	Indefinite	Indefinite	Yes	Yes

- These properties are the result of the arrangement of the atoms and molecules comprising a sample of matter.



Solid

- The particles in a solid are packed close together and are fixed in position
  - though they may vibrate
- The close packing of the particles results in solids being incompressible
- The inability of the particles to move around results in solids retaining their shape and volume when placed in a new container; and prevents the particles from flowing



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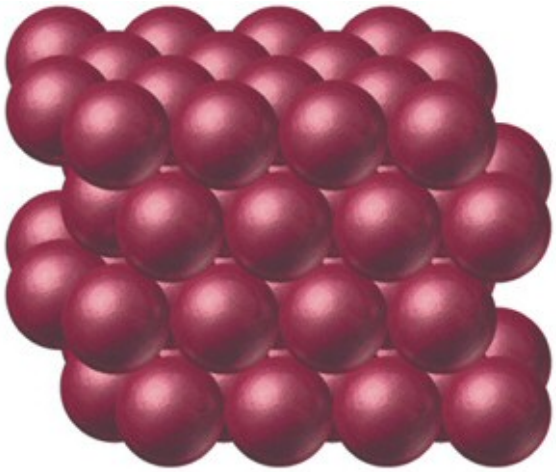
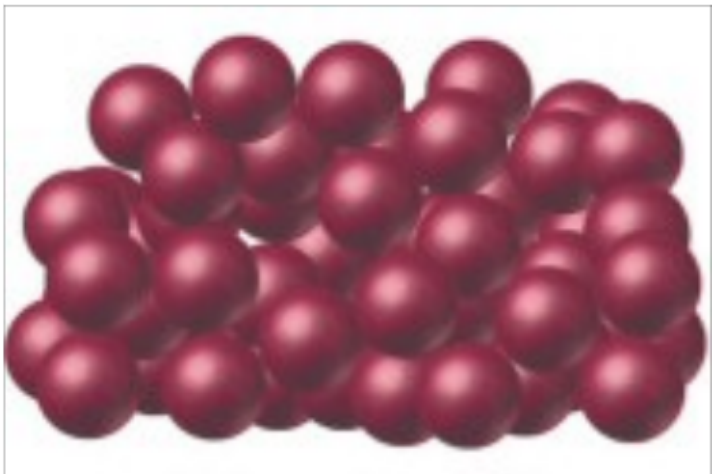
- Some solids have their particles arranged in an orderly geometric pattern – we call these **crystalline solids**.
    - Salt and Diamonds are examples

- Other solids have particles that do not show a regular geometric pattern over a long range. They are called **amorphous solids**
    - Plastic and Glass are examples



(a) Crystalline solid



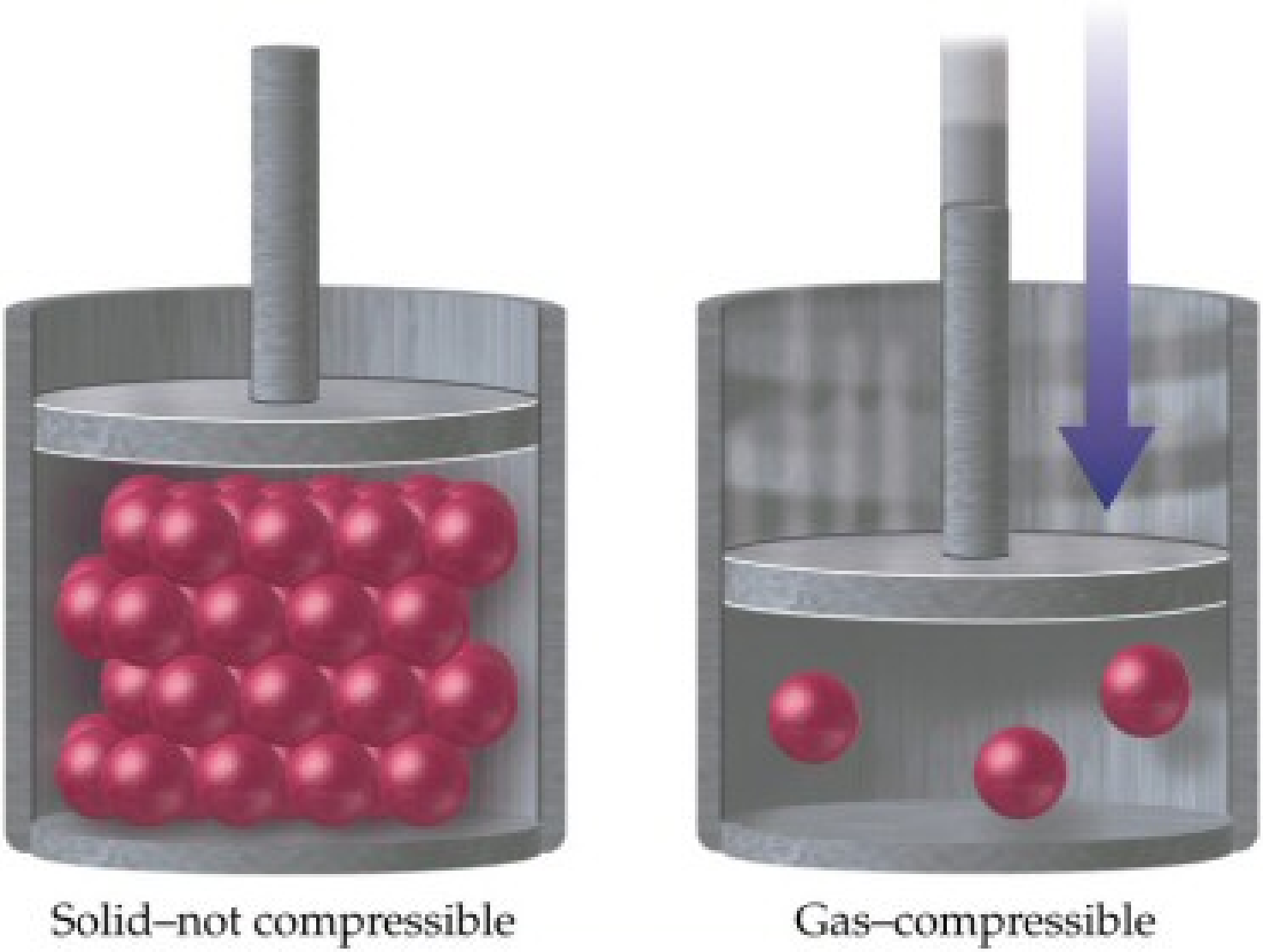
(b) Amorphous solid

Liquids

- The particles in a liquid are closely packed, but they have some ability to move around
- the close packing results in liquids being incompressible
- The ability of the particles to move allows liquids to take the shape of their container and to flow. However they don’t have enough freedom to escape and expand to fill the container.

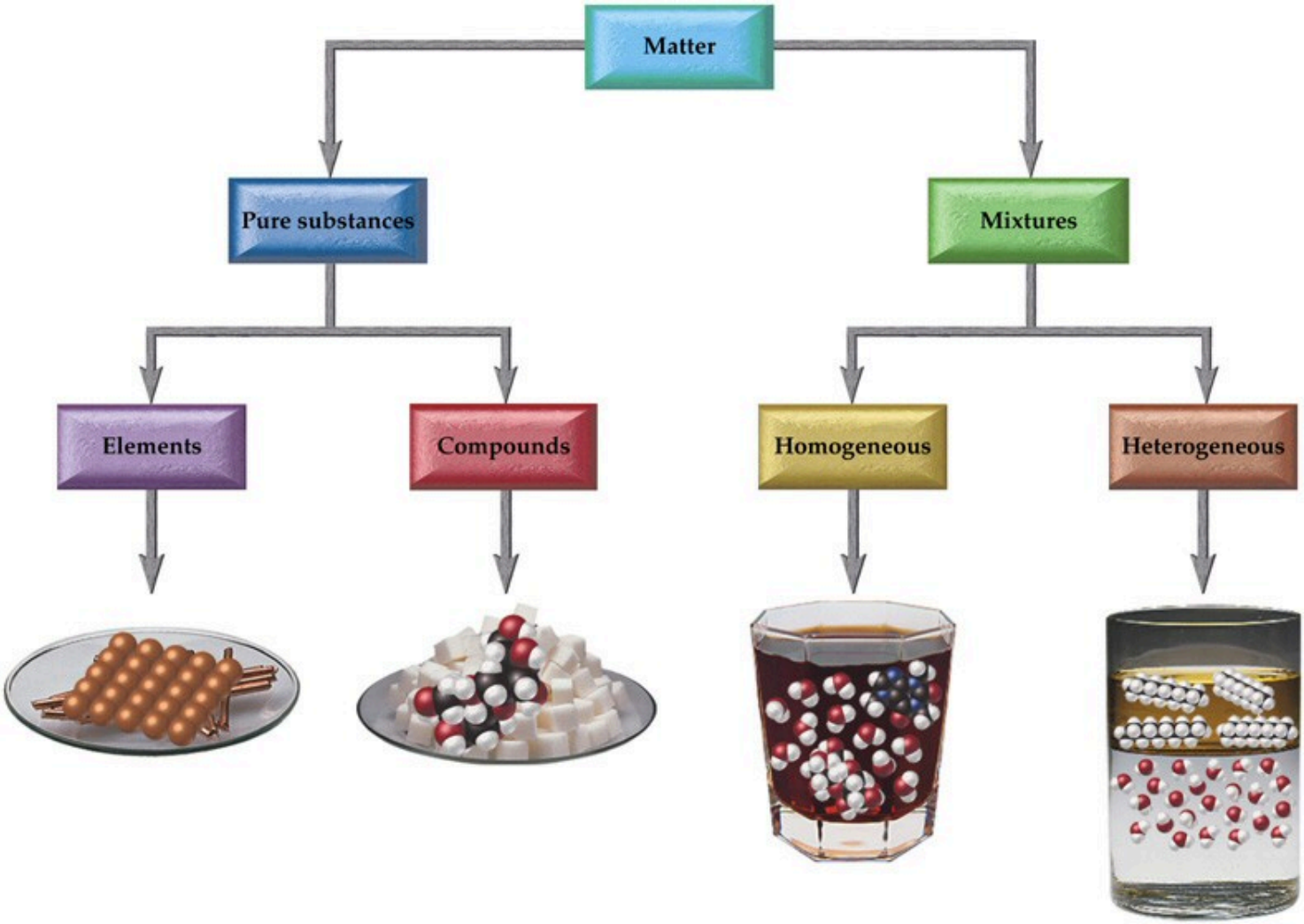
Gases

- In the gas state, the particles have complete freedom from each other
- The particles are constantly flying around, bumping into each other and the container
- In the gas state, there is a lot of empty space between the particles
  - on average
- Because there is a lot of empty space, the particles can be squeezed closer together – therefore gases are compressible



- Because the particles are not held in close contact and are moving freely, gases expand to fill and take the shape of their container, and will flow

Pure Substances vs. Mixtures



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- In a **Pure Substance** the entire sample is made of the same atoms or molecules.
    - All samples have the same properties
- In **Mixtures** different samples may have components present in different percentages

- Samples with varying composition will exhibit different properties.

Elements and Compounds

- Substances which can not be broken down into simpler substances by chemical reactions are called **elements**.
- Most substances are chemical combinations of elements. These are called **compounds**.
  - Compounds can be broken down into elements
  - Properties of the compound not related to the properties of the elements that compose it

Elements

- Smallest piece of an **element** is called an **atom**
  - There are subatomic particles, but these are no longer the element
- Every sample of an element is made up of lots of identical atoms
- 118 known, of which about 91 are found in nature
  - The others are man made. Usually inside a particle accelerator.

- There is a natural distribution of elements known as there **abundance**
  - The abundance and form of an element varies in different parts of the environment
  - Oxygen most abundant element (by mass) on earth and in the human body

Compounds

- Smallest piece of a **compound** is called a **molecule**
  - Molecules are made of atoms
  - All molecules of a compound are identical
  - Each molecule has the same number and type of atoms
- Composed of elements in fixed percentages
- water is 89 mass% O & 11 mass% H
- Billions of known compounds
- Same elements can form more than one different compound
  - Water and hydrogen peroxide contain just hydrogen and oxygen
  - Carbohydrates all contain just C, H and O

Mixtures

- Mixtures come in two forms **homogeneous** and **heterogeneous**.
- Homogeneous Mixtures are uniform throughout
  - Appears to be one thing
  - Every piece of a sample has identical properties
    - Another sample with the same components may have different properties
  - Homogeneous mixtures are sometimes called solutions
- Heterogeneous Mixtures are non-uniform
- They contain different regions with different properties

Pure Substances	Mixtures
All samples have the same physical and chemical properties	Different samples may show different properties
Constant composition; all samples have the same components in the same percentages.	Variable composition; samples made with the same pure substances may have different percentages

Pure Substances	Mixtures
Homogeneous	Homogeneous or Heterogeneous
Separate components of a **compound** based on chemical properties	Separate into components based on physical properties
Temperature usually stays constant while melting or boiling	Temperature changes while melting or boiling because composition changes

Properties of Matter

- **Physical Properties** are the characteristics of matter that can be changed without changing its composition
- **Chemical Properties** are the characteristics that determine how the composition of matter changes as a result of contact with other matter or the influence of energy

Some Physical Properties

- mass
- volume
- density

- phase
- magnetic susceptability
- specific heat
- melting point
- boiling point
- volatility
- taste
- solubility
- electrical condictivity
- thermal conductivity
- malleability
- ductility

Some Chemical Properties

- acidity
- basicity
- corrosiveness
- reactivity
- explosiveness

- flammability
- combustibility
- reduction potential

Properties of a Mixture

Type	Color	% Cu	% Zn	Density g/cm³	MP °C	Tensile Strength psi	Uses
Gilding	reddish	95	5	8.86	1066	50K	pre-83 pennies, munitions, plaques
Commercial	bronze	90	10	8.80	1043	61K	door knobs, grillwork
Jewelry	bronze	87.5	12.5	8.78	1035	66K	costume jewelry
Red	golden	85	15	8.75	1027	70K	electrical sockets, fasteners & eyelets
Low	deep yellow	80	20	8.67	999	74K	musical instruments, clock dials
Cartridge	yellow	70	30	8.47	954	76K	car radiator cores
Common	yellow	67	33	8.42	940	70K	lamp fixtures, bead chain
Muntz metal	yellow	60	40	8.39	904	70K	nuts & bolts,

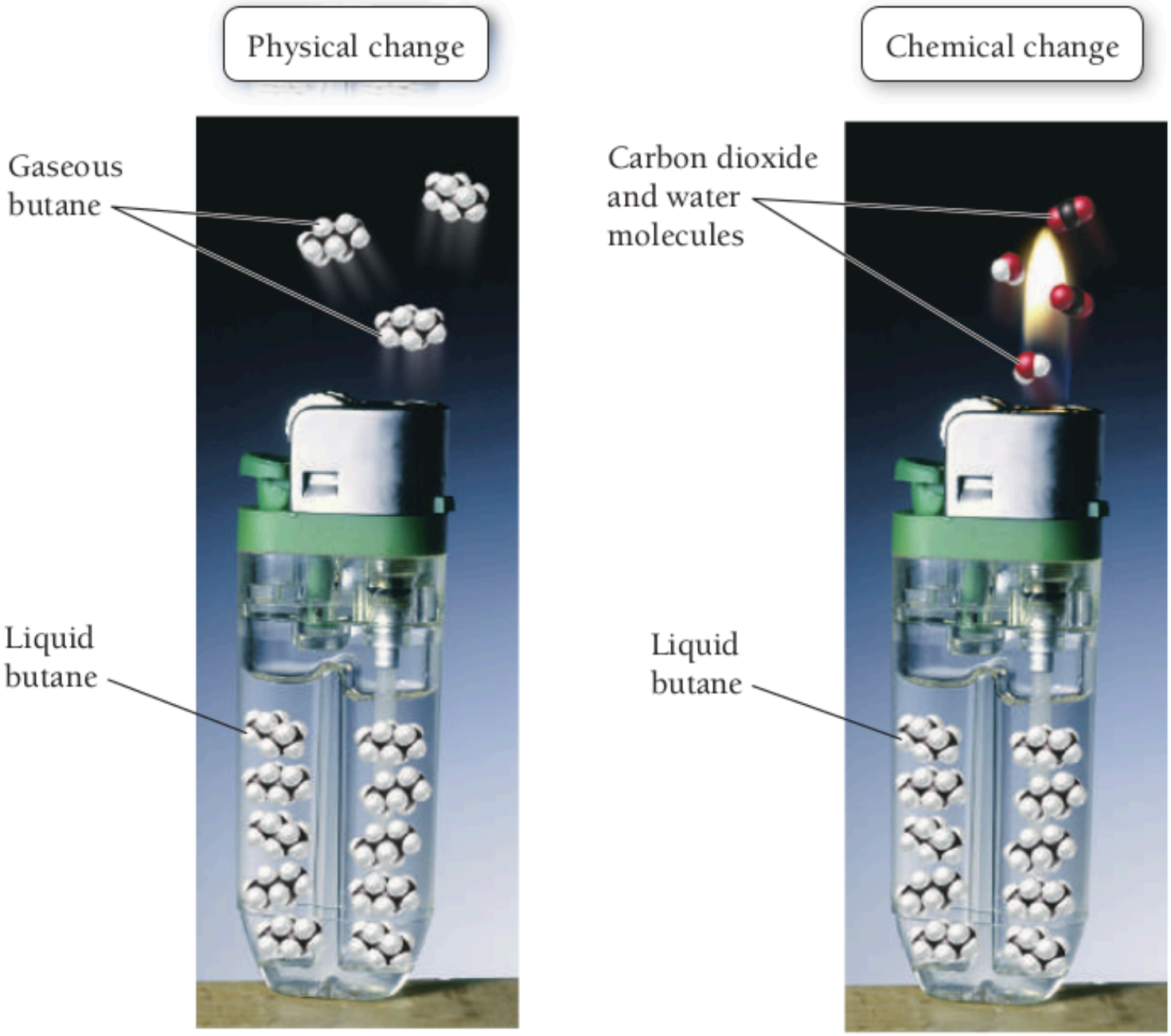
Physical Changes

- Physical Changes are changes in the properties of matter that do not effect its composition

- Heating water raises its temperature. But it is still water.
- Evaporating butane from a lighter
- Dissolving sugar in water

Chemical Changes

- Chemical Changes involve a change in a sample's composition. A Chemical Reaction.
  - Rusting is iron combining with oxygen to make iron(III) oxide
  - Burning butane from a lighter changes it into carbon dioxide and water
  - Silver combines with sulfur in the air to make tarnish



Phase Changes are Physical Changes

- Boiling = liquid to gas

- Melting = solid to liquid
- Subliming = solid to gas
- Condensing = gas to liquid
- Freezing = liquid to solid
- Deposition = gas to solid
- State changes require heating or cooling the substance
    - Evaporation is not a simple phase change, it is a solution process

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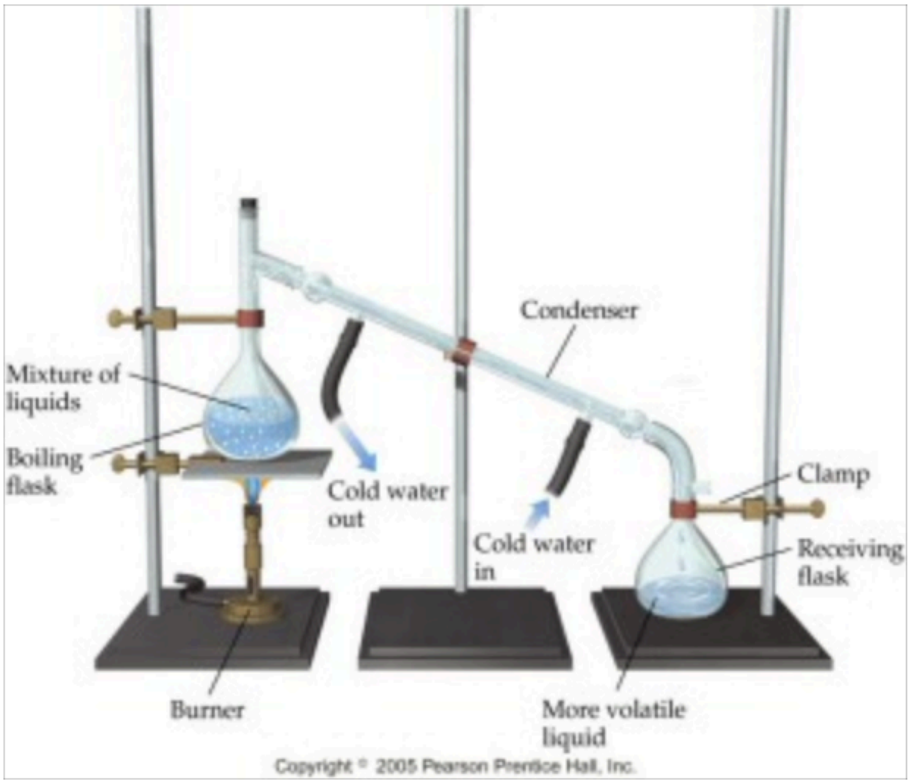
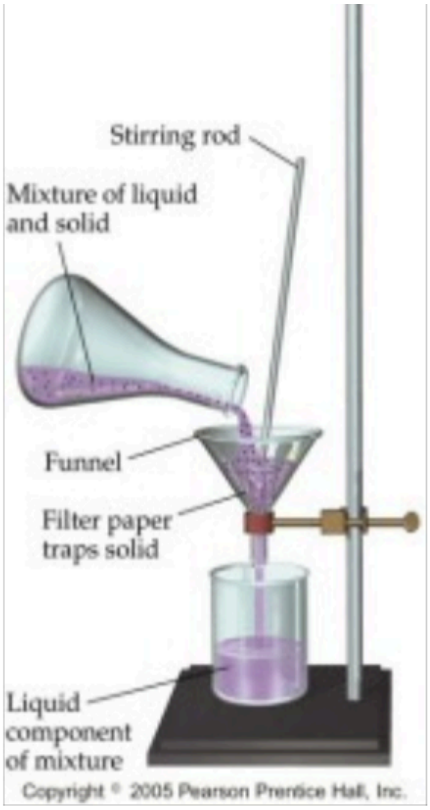
Separation of Mixtures

- Mixtures are separated based on different physical properties of the components

Physical Property	Separation Technique
Boiling Point	Distillation
Phase	Filtration
Surface Adhesion	Chromatography
Volatility	Evaporation
Density	Centrifugation

Distillation and Filtration





Law of Conservation of Mass

- This Law is attributed to Antoine Lavoisier

> *"Matter is neither created nor destroyed in a chemical reaction"*

- The total amount of matter present before a chemical reaction is always the same as the total amount after
- The total mass of all the reactants is equal to the total mass of all the products
- Total amount of matter remains constant in a chemical reaction

58 grams of butane burns in 208 grams of oxygen to form 176 grams of carbon dioxide and 90 grams of water

$$\begin{gathered}\text{butane} + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water} \\ 58\,\text{g} + 208\,\text{g} \longrightarrow 176\,\text{g} + 90\,\text{g} \\ 266\,\text{g} \longrightarrow 266\,\text{g}\end{gathered}$$

Energy

- We have observed something that has neither mass or volume, Energy.
- Energy is anything that has the capacity to do work

- Even though Chemistry is the study of matter, matter is effected by energy
    - it can cause physical and/or chemical changes in matter

Law of Conservation of Energy

> *"Energy can neither be created nor destroyed"*

- The total amount of energy in the universe is constant – there is no process that can increase or decrease that amount
- However we can transfer energy from one place in the universe to another, and we can change its form
    - When a piece of matter possesses energy, it can give some it to another object
- All chemical and physical changes result in matter releasing or absorbing energy

Kinds of Energy

- **Kinetic Energy** is energy of motion, or energy that is being transferred from one object to another
- **Potential Energy** is energy that is stored

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- **Electrical Energy** is kinetic energy associated with the flow of electrical charge

- **Thermal Energy** is kinetic energy associated with molecular motion
- **Light** or **Radiant Energy** is kinetic energy associated with energy of subatomic particles called photons
- **Nuclear Energy** is potential energy in the nucleus of atoms
- **Chemical Energy** is potential energy in the attachment of atoms or because of their position

- We use energy to accomplish all kinds of processes, but according to the Law of Conservation of Energy we don't really use it up!
  - When we use energy we are changing it from one form to another
  - For example, converting the chemical energy in gasoline into mechanical energy to make your car move
- In practice no process is 100% efficient. Some energy will be loss usually in the form of heat.

Units of Energy

- **calorie** (cal) is the amount of energy needed to raise one gram of water by 1°C

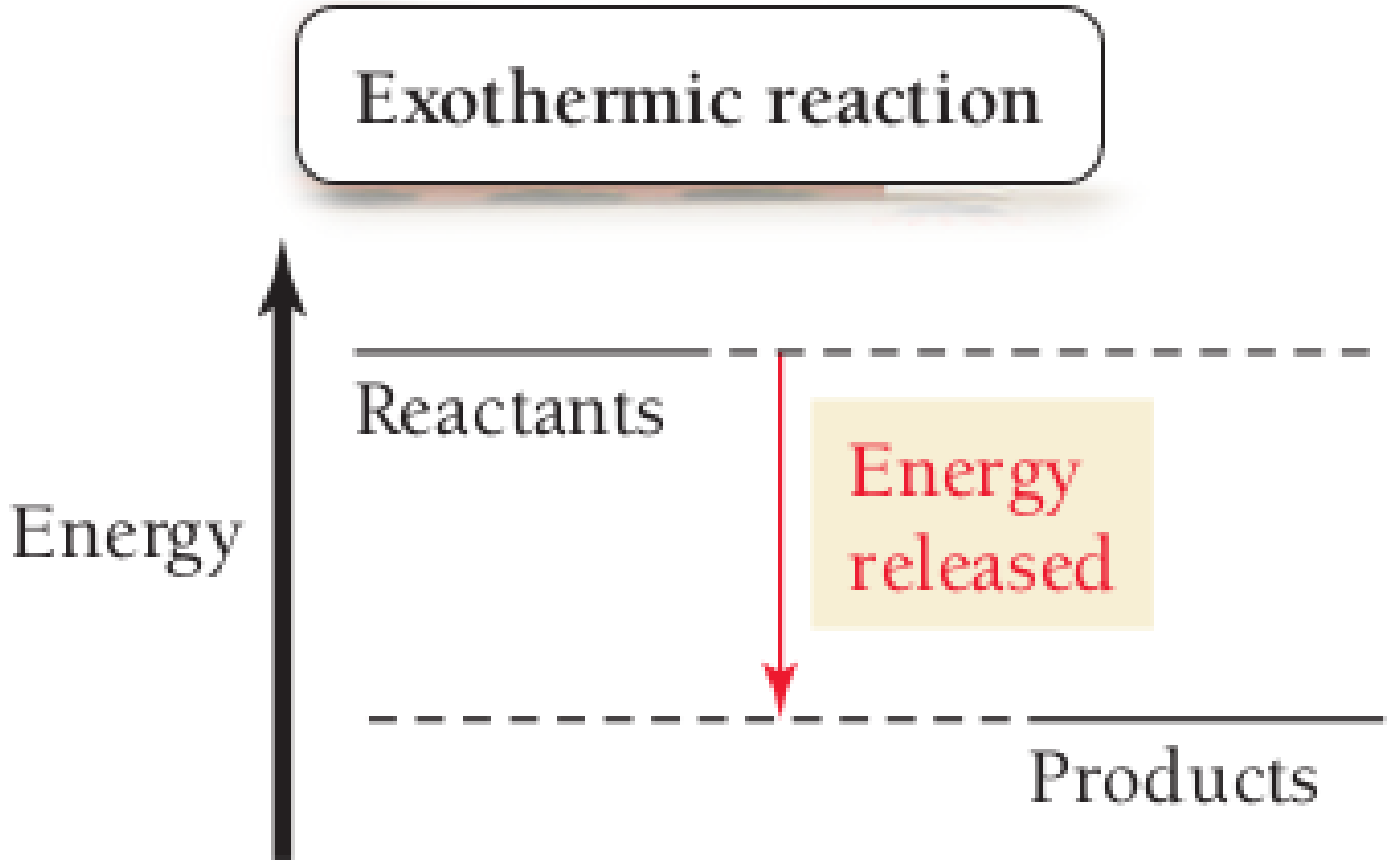
- 1 *food calorie* or **Calorie** (cal)[Note the capital "C"} is 1,000 *calories* (cal) [Lower case "c"]l
    - 1 Cal = 1000 cal = 1 kcal
- Joule (J) is equal to the amount of work done when a force of 1 newton displaces a mass through a distance of 1 meter in the direction of the force applied.
    - It is the standard SI unit for energy
- Kilowatt-hour (kWh) is the energy delivered by 1000 Watts of power over one hour.
    - Typically used when dealing with large amounts of energy

Unit	Energy Required to Raise Temperature of 1 g of Water by 1 °C	Energy Required to Light 100-W Bulb for 1 Hour	Total Energy Used by Average U.S. Citizen in 1 Day
joule (J)	4.18	3.6×10^5	9.0×10^8
calorie (cal)	1.00	8.60×10^4	2.2×10^8
Calorie (Cal)	0.00100	86.0	2.2×10^5
kilowatt-hour (kWh)	1.16×10^{-6}	0.100	2.50×10^2

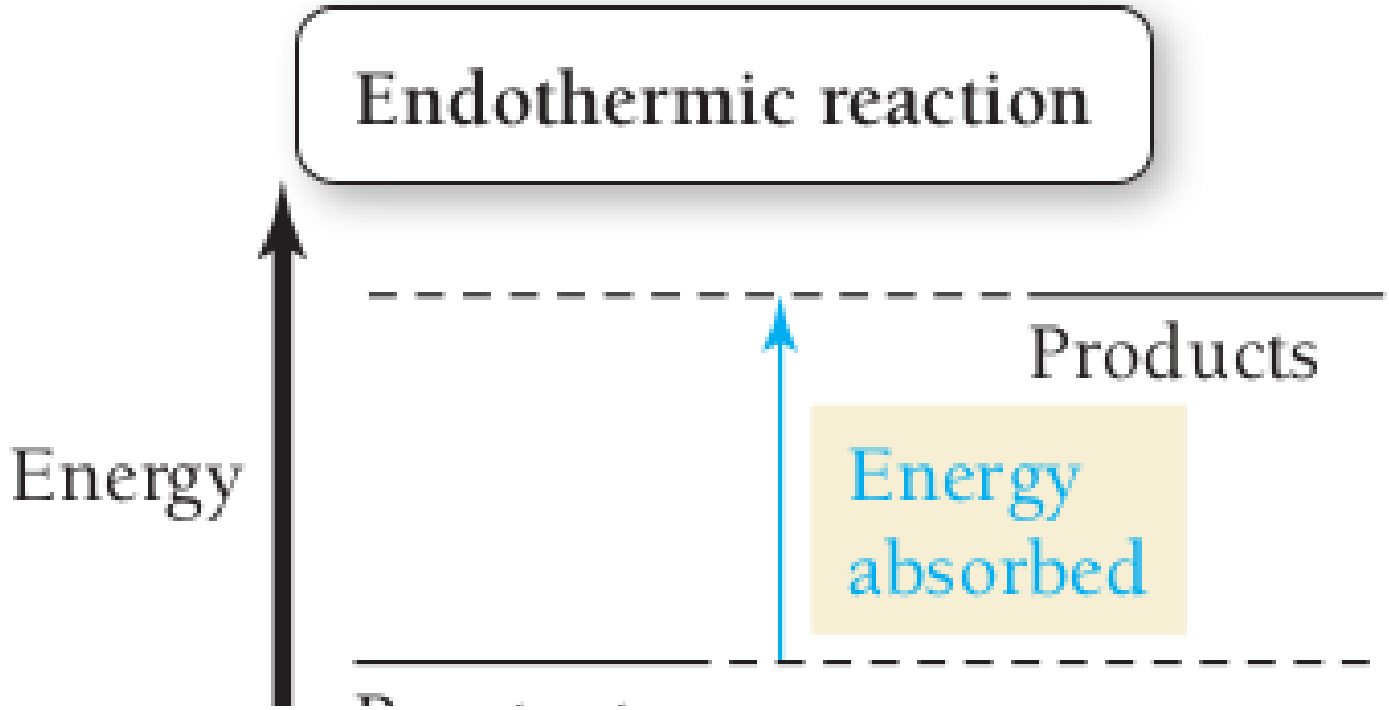
A candy bar contains 225 Cal of nutritional energy. How many joules does it contain?

Exothermic vs. Endothermic

- A chemical change (reaction) can either release or absorb energy
- Chemical reactions where energy is released are called **exothermic**
- Chemical reactions where energy is absorbed are called **endothermic**
- Energy is usually transferred in the form of heat


Exothermic reaction
Energy
Reactants
Energy released
Products

(a)


Endothermic reaction
Energy
Products
Energy absorbed

Classify each process as exothermic or endothermic.

a. gasoline burning in a car
b. isopropyl alcohol evaporating from skin
c. water condensing as dew during the night

Heat

- Heat is the exchange of thermal energy between samples of matter
- Heat flows from the matter that has high thermal energy to matter that has low thermal energy
  - Until they reach the same temperature
- Heat is exchanged through molecular collisions between two samples

The Meaning of Temperature

- Temperature is a measure of the average kinetic energy of the molecules in a sample
- Not all molecules have in a sample the same amount of kinetic energy
- A higher temperature means a larger *average* kinetic energy

Fahrenheit

- The Fahrenheit Temperature Scale used as its two reference points the freezing point of concentrated saltwater (0 °F) and average body temperature (100 °F)
    - more accurate measure now set average body temperature at 98.6 °F
- Room temperature is about 75 °F

Celsius

- The Celsius Temperature Scale used as its two reference points the freezing point of distilled water (0 °C) and boiling point of distilled water (100 °C)
    - more reproducible standards

    - most commonly used in science
- Room temperature is about 25 °C

Fahrenheit vs. Celsius

- A Celsius degree is 1.8 times larger than a Fahrenheit degree
- 0 °C is 32 °F
- Because the scales are offset from one another, we have a more complicated formula for converting between them.

$$T\circ_C = \frac{T\circ_F - 32}{1.8} \qquad T\circ_F = 1.8T\circ_C + 32$$

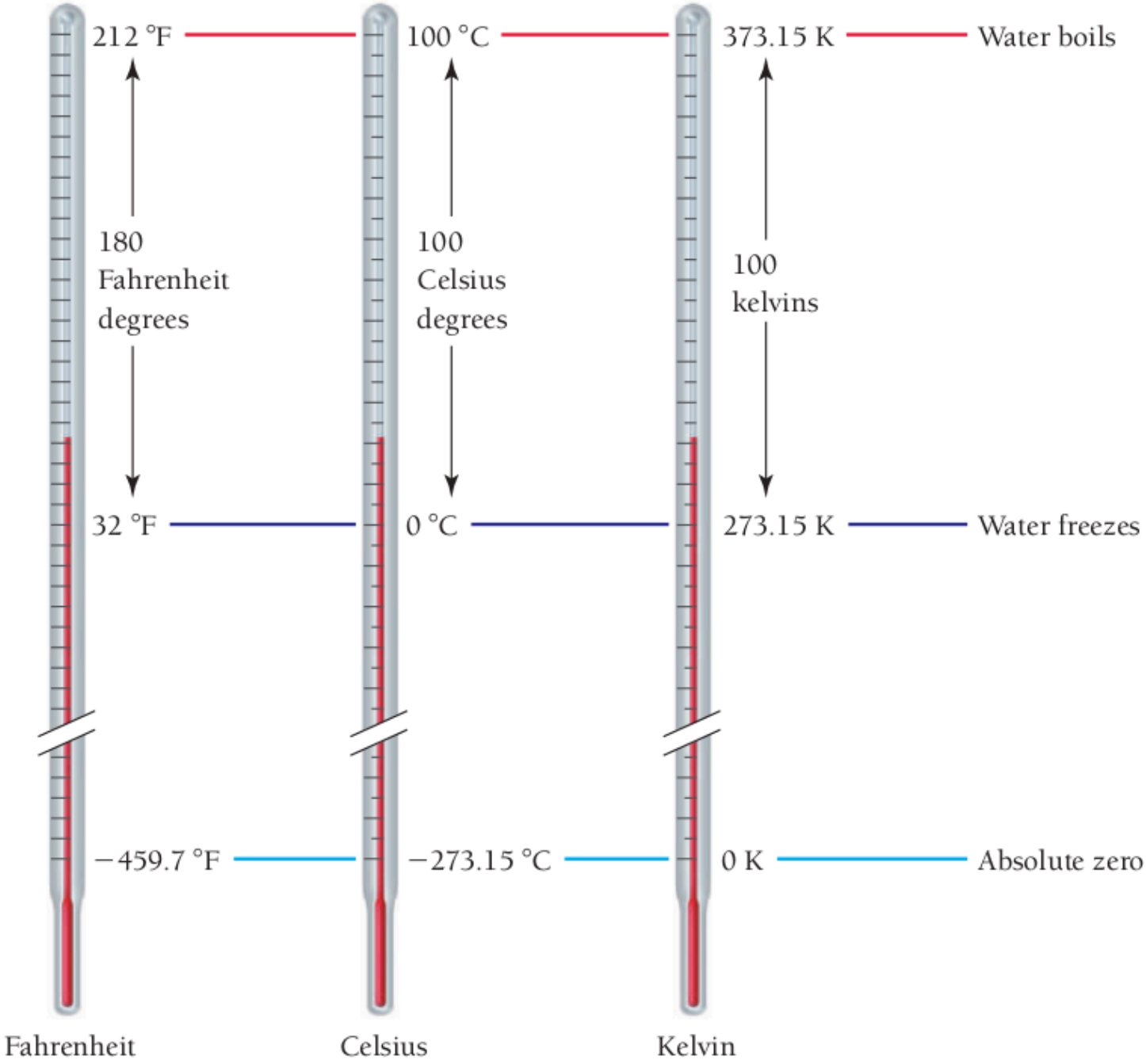
The Kelvin Temperature Scale

- Both the Celsius and Fahrenheit scales have negative numbers
- The Kelvin scale is an absolute scale, meaning it does not allow for negative values.
- 0 K is called **Absolute Zero**. The lowest possible temperature.
    - All molecular motion would stop at 0 K
    - Absolute Zero is a theoretical value and has not yet been achieved in lab.

Kelvin vs. Celsius

- The size of a “degree” on the Kelvin scale is the same as on the Celsius scale
  - that makes 1 K 1.8 times larger than 1°F
- The 0 standard on the Kelvin scale is a much lower temperature than on the Celsius scale

$$T_K = T\circ_C + 273$$


212 °F
100 °C
373.15 K
Water boils
180
Fahrenheit
degrees
100
Celsius
degrees
100
kelvins
32 °F
0 °C
273.15 K
Water freezes
−459.7 °F
−273.15 °C
0 K
Absolute zero
Fahrenheit
Celsius
Kelvin

Convert 310 K to Fahrenheit

Energy and the Temperature of Matter

- The amount the temperature of an object increases depends on the amount of heat energy added (q).
  - If you double the added heat energy the temperature will increase twice as much.
- The amount the temperature of an object increases depends on its mass
  - If you double the mass it will take twice as much heat energy to raise the temperature the same amount.

Heat Capacity

- Heat capacity is the amount of heat *an object* must absorb to raise its temperature 1°C
  - cal/°C or J/°C
  - Metals have low heat capacities
  - Thermal insulators high
  - Extensive quantity

- Specific heat = heat capacity of 1 gram of the substance
  - cal/g°C or J/g°C
  - Waters specific heat = 4.184 J/g°C for liquid
    - 1.000 cal/g°C
- The larger a material's specific heat is, the more energy it takes to raise its temperature.
  - Water's high specific heat is the reason it is such a good cooling agent
- like density, specific heat is a property of the type of matter
  - It can be used to identify the type of matter
- it doesn't matter how much material you have
  - Intensive quantity

Substance	Specific Heat ( )
Aluminum	0.895
Calcium	0.656
Carbon (diamond)	0.508
Carbon (graphite)	0.708

Substance	Specific Heat ($\frac{J}{g \cdot C}$)
Copper	0.377
Gold	0.129
Iron	0.448
Lead	0.129
Silver	0.712
Water (*l*)	4.184
Water (*s*)	2.03
Water (*g*)	2.02

Heat Gain or Loss by an Object

- The amount of heat energy gained or lost by an object depends on 3 factors
  - The mass of the substance (*m*)
  - The sunstances **Specific Heat Capacity** (*c*)
  - The temperature changed, T_F - T_i, or ΔT

$$q = mc\Delta T$$

Gallium is a solid metal at room temperature, but melts at 29.9°C. If you hold gallium in your hand, it melts from body heat. How much heat must 2.5 g of gallium absorb from your hand to raise its temperature from 25.0°C to 29.9°C? The heat capacity of gallium is 0.372 J/g°C

A backpacker wants to carry enough fuel to heat 2.5 kg ofwater from 25 °C to 100.0 °C. If the fuel she carries produces 36 kJ of heat per gram when it burns, how muchfuel should she carry? (For the sake of simplicity, assume that the transfer of heat is 100% efficient.)

If 89 J of heat are added to a pure gold coin with a mass of 12 g, what is its temperature change?

An iron nail with a mass of 12 g absorbs 15 J of heat. If the nail was initially at 28 °C, what is its final temperature?

Chapter 3 Quiz

1. An amorphous solid has long range, repeating order. (1 Point)

   True False

2. Liquids have definite volume and shape. (1 Point)

   True False

3. Solids have indefinite shape but a defined volume. (1 Point)

   True False

4. Saltwater is a mixture. (1 Point)

   True False

5. Calcium is a pure substance. (1 Point)

   True False

6. Flammability of gasoline is a chemical property. (1 Point)

   True False

7. The boiling of water is a physical change. (1 Point)

True False

8. In physical changes, the atoms or molecules that compose the matter do not change their identity, even though the matter may change its appearance. (1 Point)

True False

9. All chemical changes are endothermic. (1 Point)

True False

10. Like mass, energy can neither be created nor destroyed. (1 Point)

True False

11. An increase of temperature implies the atoms/molecules of a substance are moving faster. (1 Point)

True False

12. The coldest temperature possible is 0 °C. (1 Point)

True False

13. Which of the following statements about matter is FALSE? (2 Point)

A) Matter occupies space and has mass.

B) Matter exists in either a solid, liquid or gas state.

C) Matter is ultimately composed of atoms.

D) Matter is always smooth and continuous.

E) none of the above

14. A pure substance is: (2 Point)

A) composed of two or more different types of atoms or molecules combined in variable proportions.

B) composed of only one type of atom or molecule.

C) composed of two or more regions with different compositions.

D) composed of two or more different types of atoms or molecules with a constant composition.

E) none of the above

15. When ammonium nitrate (NH_4NO_3) is heated, it produces nitrous oxide (N_2O) and water (H_2O). If 50.0 g of ammonium nitrate is heated yielding 32.0 g of water, what mass of nitrous oxide was produced? (2 Point)

A) 8.0 g

B) 18.0 g

C) 82.0 g

D) 32.0 g

E) none of the above

16. How many joules are there in a 255 Calorie snack bar? (2 Point)

A) 2.55×10^{-6} J

B) 1.07×10^{6} J

C) 1.07×10^{3} J

D) 6.09×10^{4} J

E) none of the above

17. If a particular process is exothermic, then energy was (2 Point)

A) absorbed

B) not transferred

C) released

D) destroyed

E) none of the above

18. What is the value of 47°C on the Kelvin temperature scale? (2 Point)

A) 273 K

B) 320 K

C) 300 K

D) 226 K

E) none of the above

19. What is the value of 27°C on the Fahrenheit temperature scale? (2 Point)

A) -6.8

B) 106

C) 300

D) 81

E) none of the above

20. A lead ball is heated from 25.0 °C to 103 °C by 40.5 joules of heat. Given the specific heat of lead is 0.128 J/g·°C, what is the mass of the ball? (2 Point)

A) 3.07 g

B) 12.7 g

C) 41.0 g

D) 4.06 g

E) none of the above

21. How much heat (kJ) is needed to raise the temperature of 100.0 grams of water from 25.0°C to 50.0°C? (2 Point)

A) 10450 kJ

B) 0.598 kJ

C) 1.05 kJ

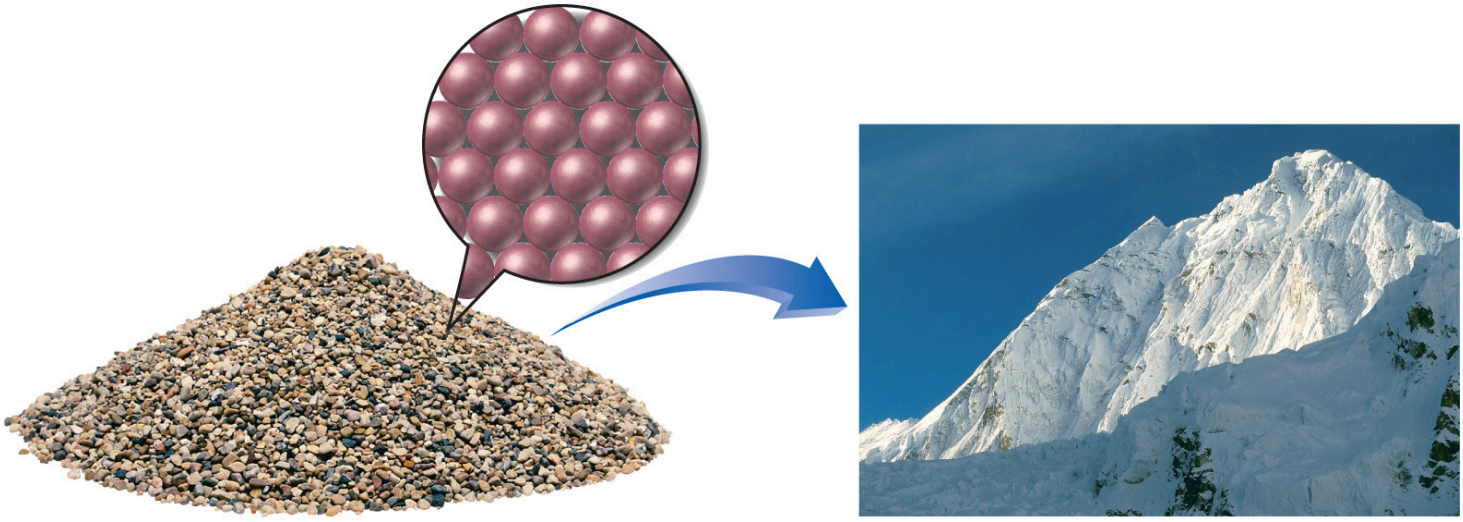
D) 10.5 kJ

E) none of the above

Chapter 4

Atoms and Elements

- Atoms are incredibly small, yet they compose everything.
- Atoms are the pieces of elements.
  - Each has its own, unique kind of atom.
  - They have different structures. Therefore they have different properties.
- Properties of the atoms determine the properties of the elements.



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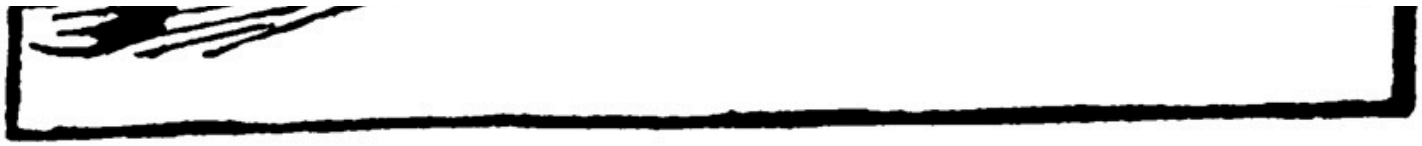
The Divisibility of Matter

- Infinitely divisible
  - For any two points, there is always a point between.
- Ultimate particle
  - Upon division, eventually a particle is reached which can no longer be divided.

“Nothing exists except atoms and empty space; everything else is opinion.” - Democritus 460–370 B.C.

Diogenes ~
Democritus ·



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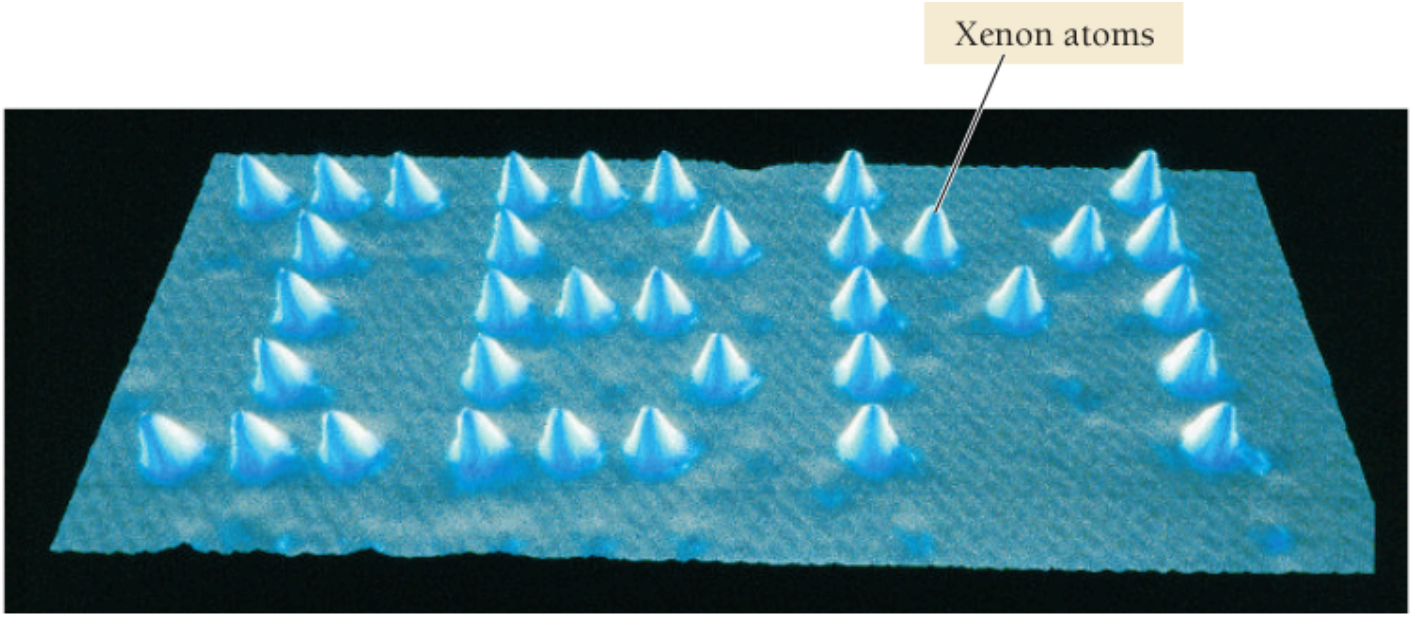
Dalton’s Atomic Theory

1. Each Element is composed of tiny, indestructible particles called atoms.
2. All atoms of an element are identical.

- They have the same mass, volume, and other physical and chemical properties.

3. Atoms combine in simple, whole-number ratios to form molecules of compounds.

- Because atoms are unbreakable, they must combine as whole atoms.
- The nature of the atom determines the ratios in which it combines.
- Each molecule of a compound contains the exact same types and numbers of atoms.
    - Law of Constant Composition
    - Chemical formulas

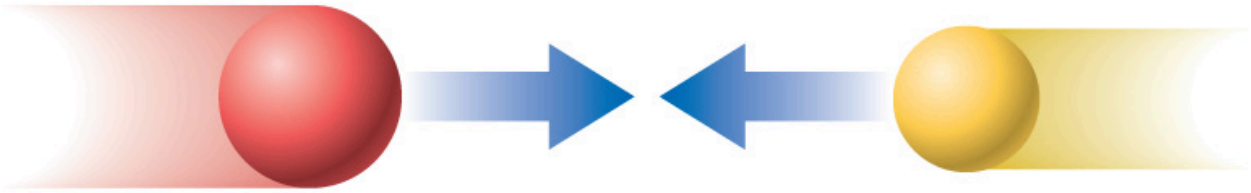


- Using compositions of compounds and assumed formulas, Dalton was able to determine the relative masses of all the atoms.
  - Dalton based his scale on H = 1 amu.
  - We now base it on C-12 = 12 amu exactly.
  - amu = atomic mass unit.
- Absolute sizes of atoms:
  Mass of H atom= 1.67×10^{-24} g.
  Volume of H atom = 2.1×10^{-25} cm^3.

Charges

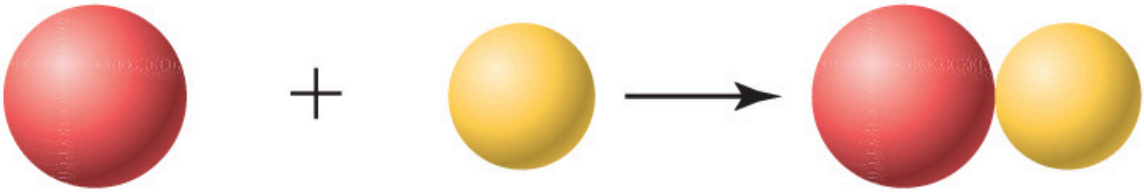
- There are two kinds of charges, called positive (+) and negative (-).

- Opposite charges attract.
- Like charges repel.
- Neutral objects **either** have no charge **or** equal amounts of opposite charges.



Positive (red) and negative (yellow) charges attract.

Positive–positive and negative–negative charges repel.



1+ + 1− = 0

Positive and negative charges cancel

Positive and negative charges cancel.

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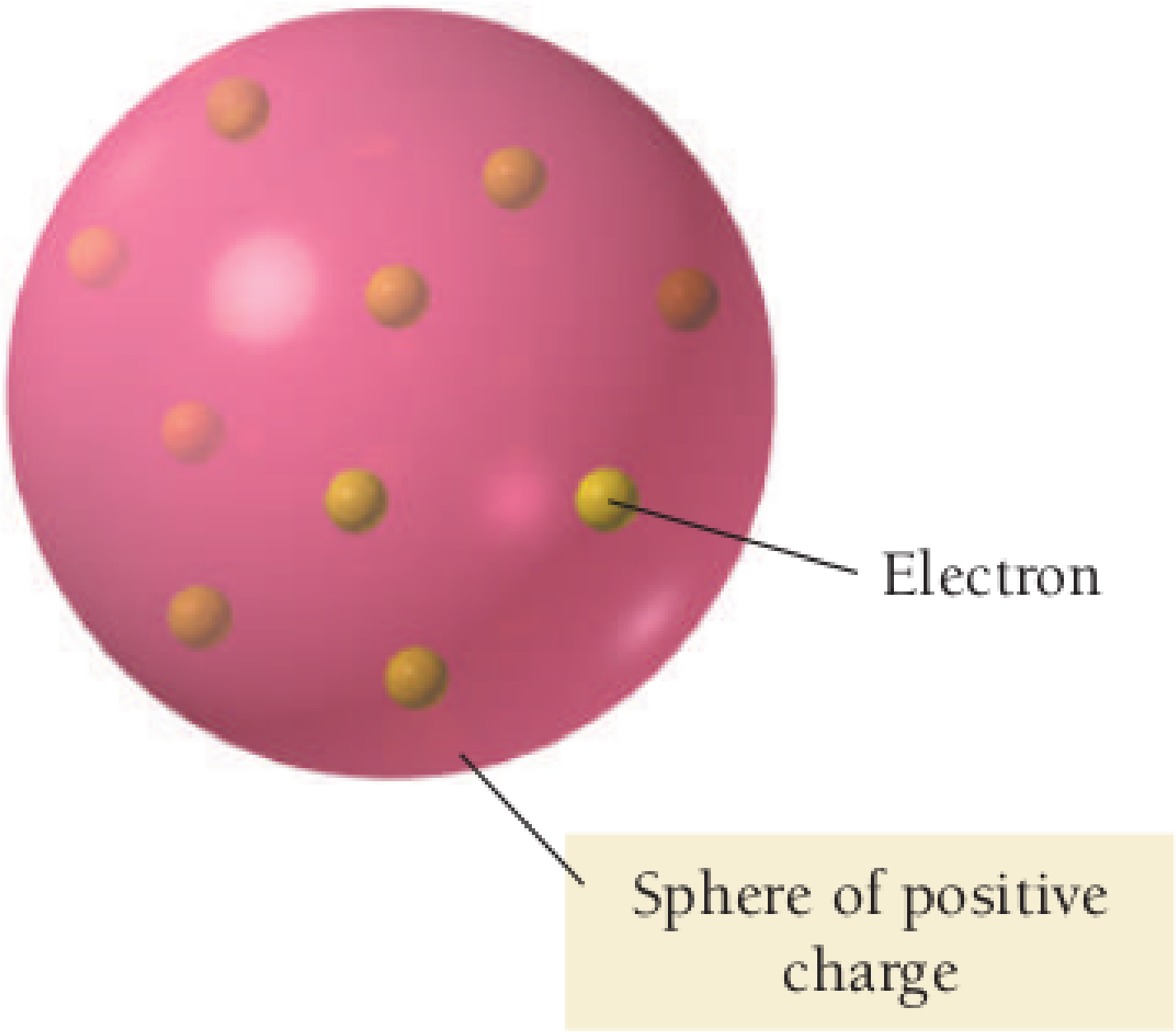
The Electron

- Work done by J. J. Thomson and others proved that the atom had pieces called **electrons**.
- Thomson found that electrons are much smaller than atoms and carry a negative charge.
  - The mass of the electron is 1/1836th the mass of a hydrogen atom.
  - The charge on the electron is the fundamental unit of charge that we call –1 charge unit.

The Plum Pudding Model



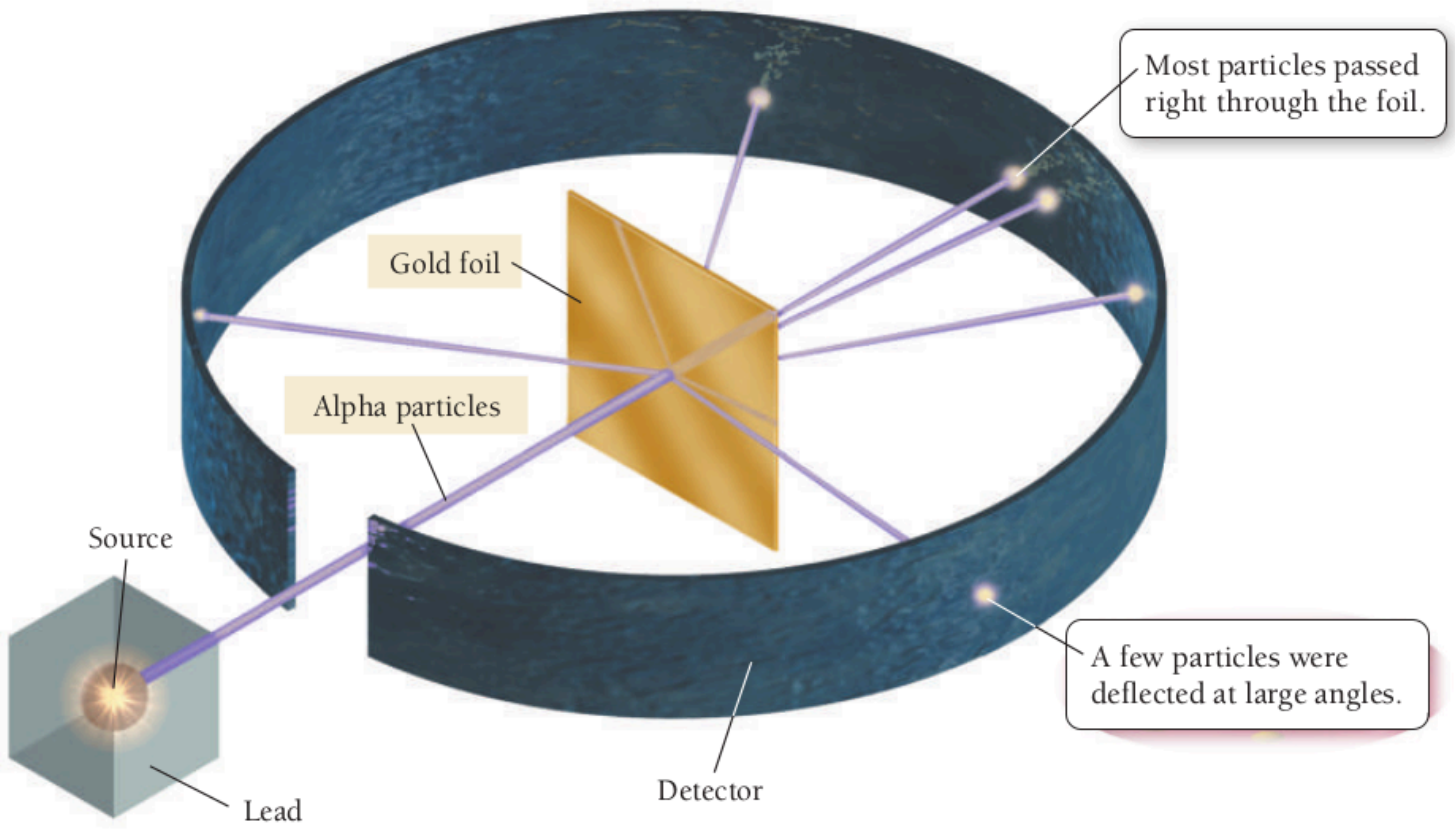
1. The mass of the atom is due to the mass of the electrons.
2. There must be a lot of empty space in the atom.

- Since the electrons are negative, it is assumed you must keep them apart so they will not repel each other.

Rutherford’s Experiment

How can you prove something is empty? Put something through it.

- Use large target atoms.
  - Use very thin sheets of target so they do not absorb the “bullet”.
- Use very small particles as “bullet” with very high energy.
  - But not so small that electrons will effect it.
- Rutherford used Gold Foil and alpha (α) particle radiation
  - Alpha particles have a mass of 4 amu & charge of +2 c.u.
  - Gold has a mass of 197 amu and is very malleable.



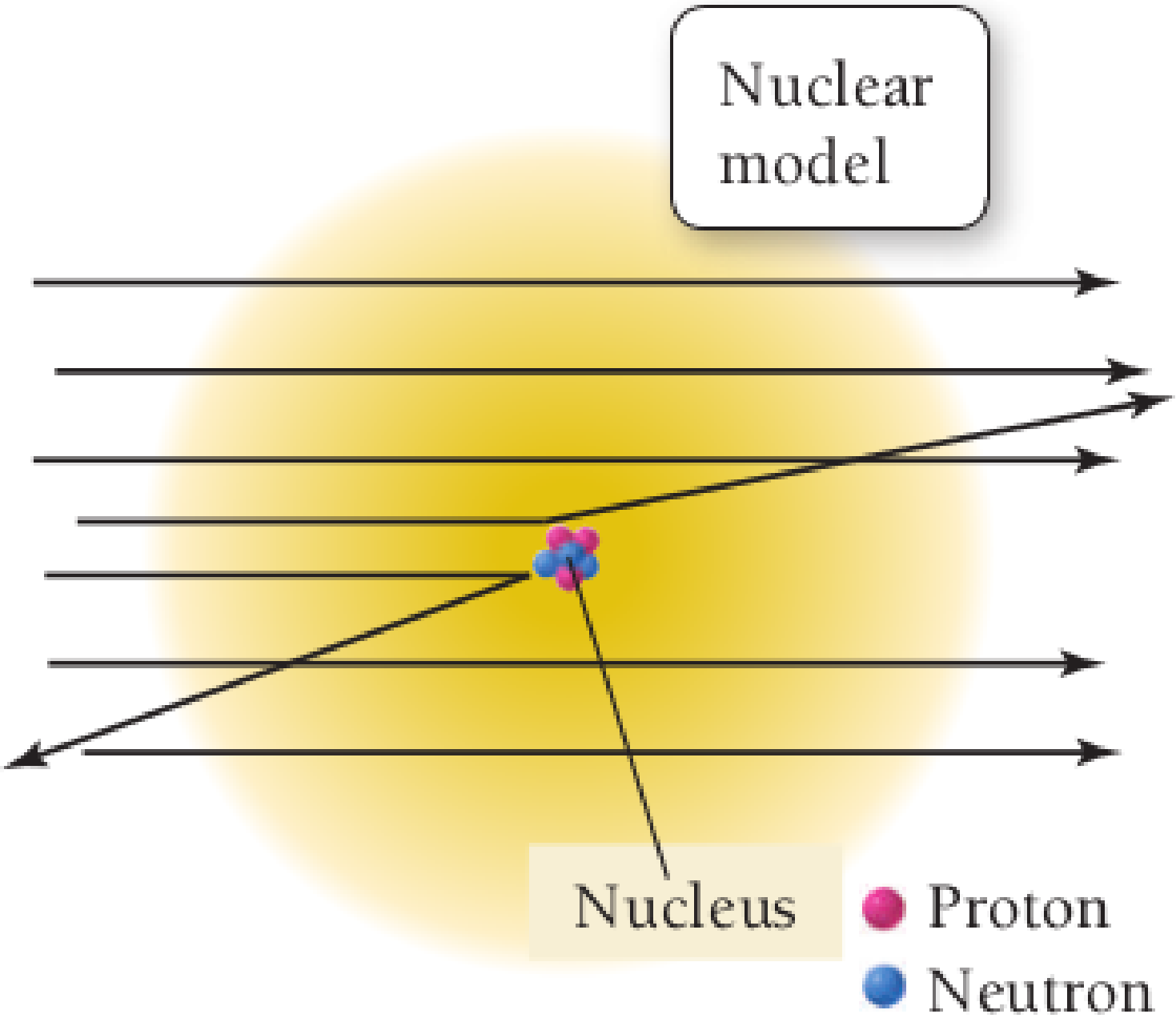
- Over 98% of the alpha particles went straight through.
- About 2% of the alpha particles went through, but were deflected by large angles.
- About 0.01% of the alpha particles bounced off the gold foil.

> *""...As if you fired a 15"-canon shell at a piece of tissue paper and it came back and hit you.""*

Rutherford's Conclusions

- Because almost all the particles went straight through, atoms are mostly empty space.

- Because of the few particles that bounced back, atoms contain a dense particle that was small in volume, compared to the atom, but large in mass.
- Because of the large deflections of some of the particles, he concluded that the dense particle was positively charged.
    - It would have to be to repel the positively charged alpha particles.



The Nuclear Model

1. The atom contains a tiny dense center called the nucleus.

- The amount of space taken by the nucleus is only about 1 trillionth the volume of the atom.

2. The nucleus has essentially the entire mass of the atom.
   The electrons weigh so little they contribute practically no mass to the atom.
3. The nucleus is positively charged.

- The amount of positive charge balances the negative charge of the electrons.

4. The electrons are dispersed in the empty space of the atom surrounding the nucleus.

- Like water droplets in a cloud.

The Proton

- Rutherford proposed that the nucleus had a particle that had the same amount of charge as an electron but opposite sign.
- He called these particles are called protons.
- Protons have a charge of +1 c.u. and a mass of 1 amu or 1.67262×10^{-27} kg.
- Since protons and electrons have the same amount of charge, for the atom to be neutral, there must be equal numbers of protons and electrons.

The Neutron

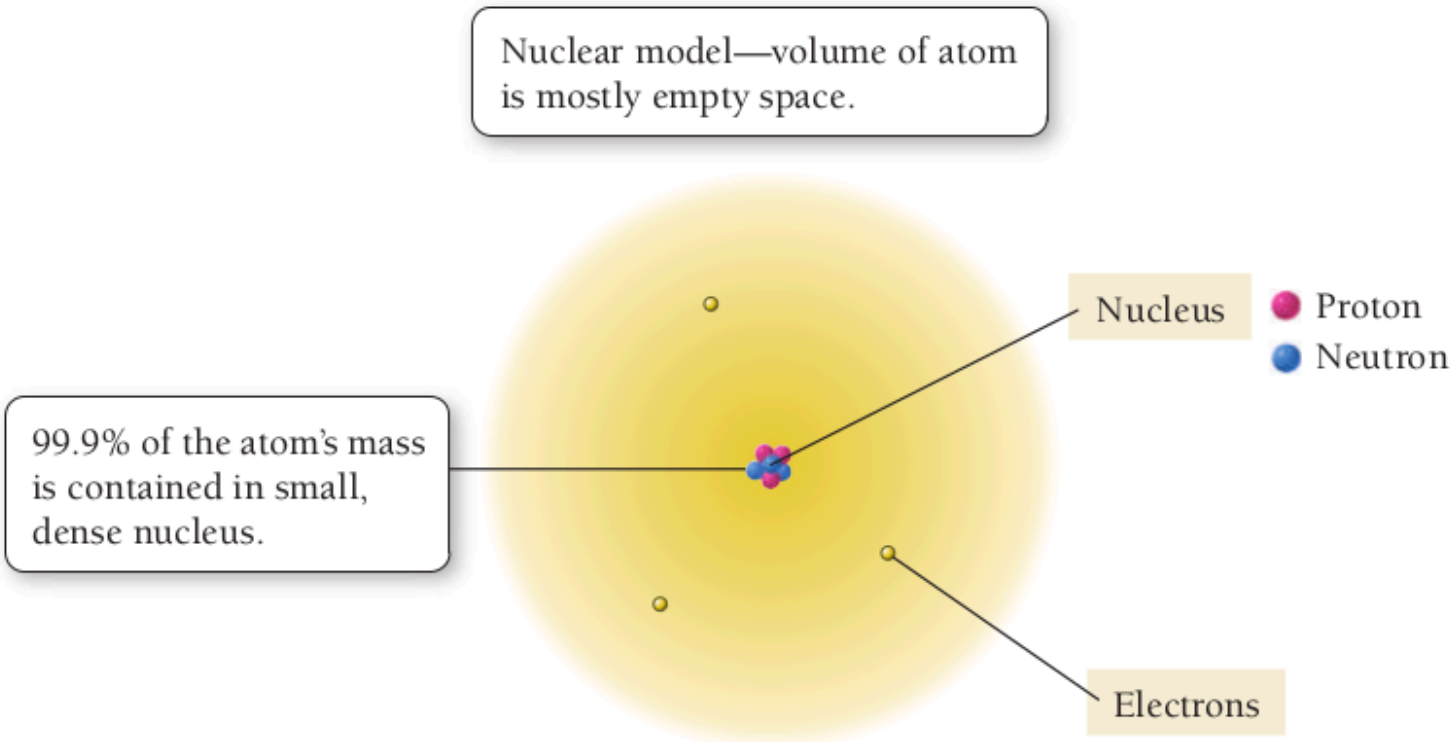
How could beryllium have 4 protons stuck together in the nucleus? Shouldn't they repel each other?

- If a beryllium atom has 4 protons, then it should weigh 4 amu, but it actually weighs 9.01 amu!
  - Where is the extra mass coming from?
- To answer these questions, Rutherford proposed that there was another particle in the nucleus.
  - Since this particle could not carry a charge he called it the neutron.
  - Neutrons have a mass of 1 amu or 1.67262×10^{-27} kg.

The Modern Atom

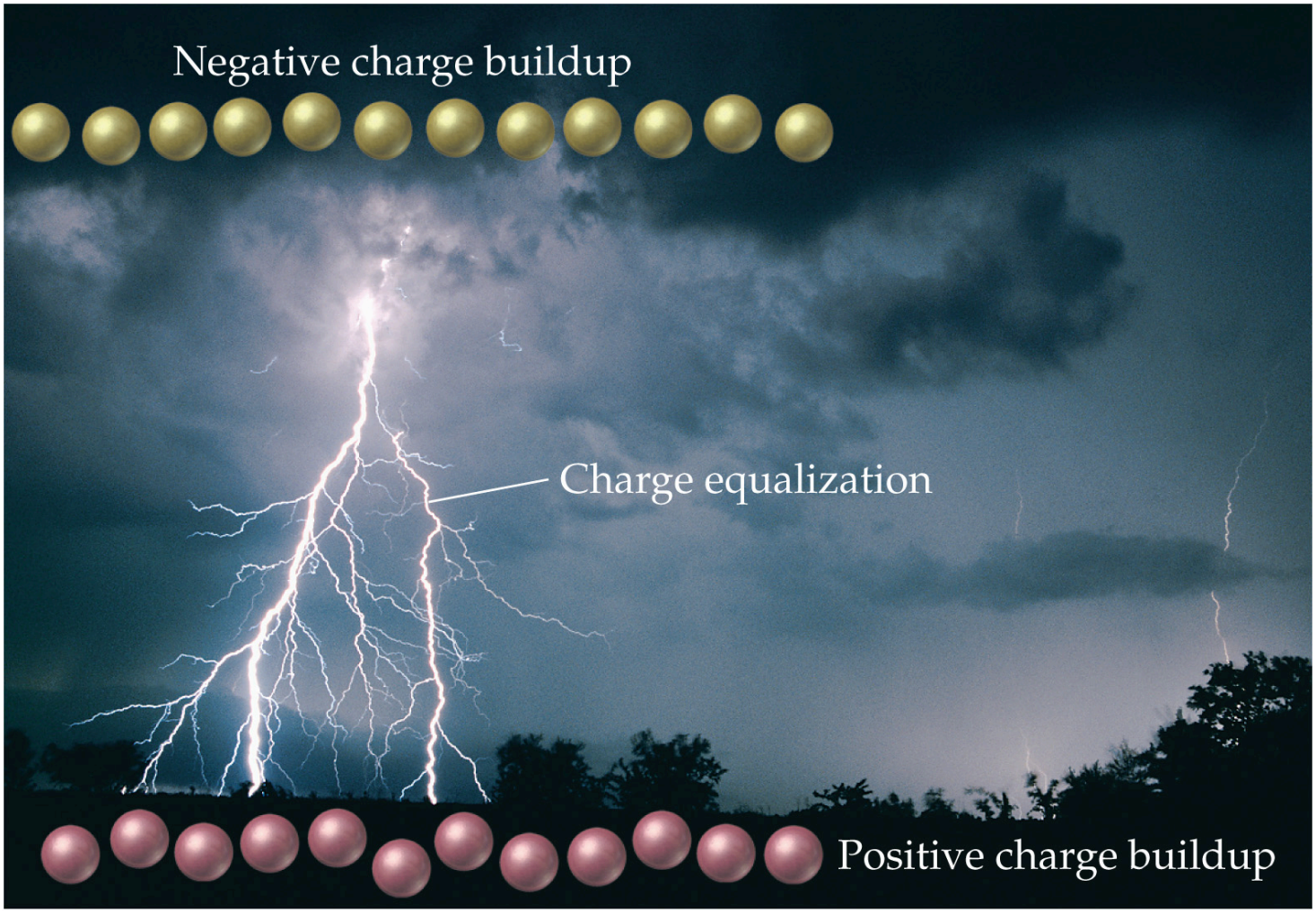
- We now know atoms are composed of three main pieces
  - protons
  - neutrons
  - electrons.
- The nucleus contains protons and neutrons.

- The radius of the atom is about 105 times larger than the radius of the nucleus.



The Nature of Electrical Charge

- Electrical charge is a fundamental property of protons and electrons.
- When a proton and electron are paired, the result is a neutral charge.
    - Because they have equal amounts of charge.



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Subatomic Particle	Mass (g)	Mass (amu)	Charge (c.u.)	Location	Symbol
Proton	1.67262×10^{-24}	1.0073	1+	Nucleus	p^+, H^+
Electron	9.1×10^{-28}	0.00055	1-	Orbital	e^-
Neutron	1.67493×10^{-24}	1.0087	0	Nucleus	n, n^0

An Atom Has 20 Protons. Determine if Each of the Following Statements Is True or False?

A. If it is a neutral atom, it will have 20 electrons.
B. If it also has 20 neutrons, its mass will be approximately 40 amu.
C. If it has 18 electrons, it will have a net 2+ charge.

The Periodic Table

Mendeleev

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- Ordered elements by atomic mass.
- Saw a repeating pattern of properties.
- Periodic law
  - When the elements are arranged in order of increasing relative mass, certain sets of properties recur periodically?
  - Used pattern to predict properties of undiscovered elements.

- Where atomic mass order did not fit other properties, he reordered by other properties.
    - Te & I

The properties (colors) of these elements form a repeating pattern.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
H	He	Li	Be	B	C	N	O	F	Ne	Na	Mg	Al	Si	P	S	Cl	Ar	K	Ca

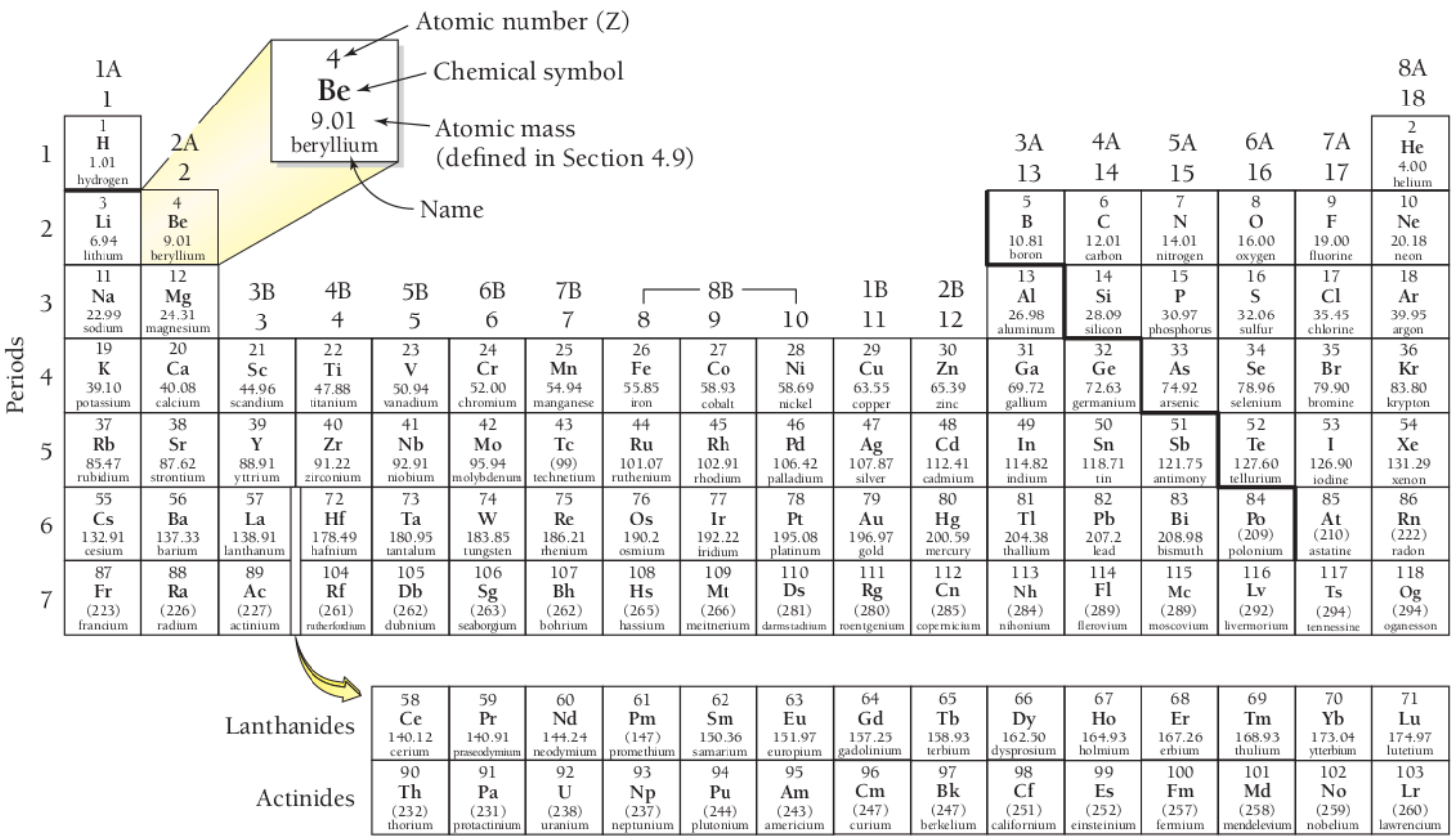
Elements with similar properties align in vertical columns.

1 H							2 He
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca						

Modern Periodic Table

- Each element has a unique number of protons in its nucleus.
  - All carbon atoms have 6 protons in their nuclei.
  - The number of protons in the nucleus of an atom is called the atomic number.
  - Z is the short-hand designation for the atomic number.
- Because each element's atoms have a unique number of protons, each element can be identified by its atomic number.
- The elements are arranged on the Periodic Table in order of their atomic numbers.
- Each element has a unique name and symbol.
  - The symbol is either one or two letters
- Elements with similar chemical and physical properties are in the same column.
  - Columns are called **Groups** or **Families**.
  - Designated by a number

- Rows are called **Periods**.
  - Each period shows the pattern of properties repeated in the next period.



What is the atomic number of boron, B?

What is the atomic mass of silicon, Si?

How many protons does a chlorine atom have?

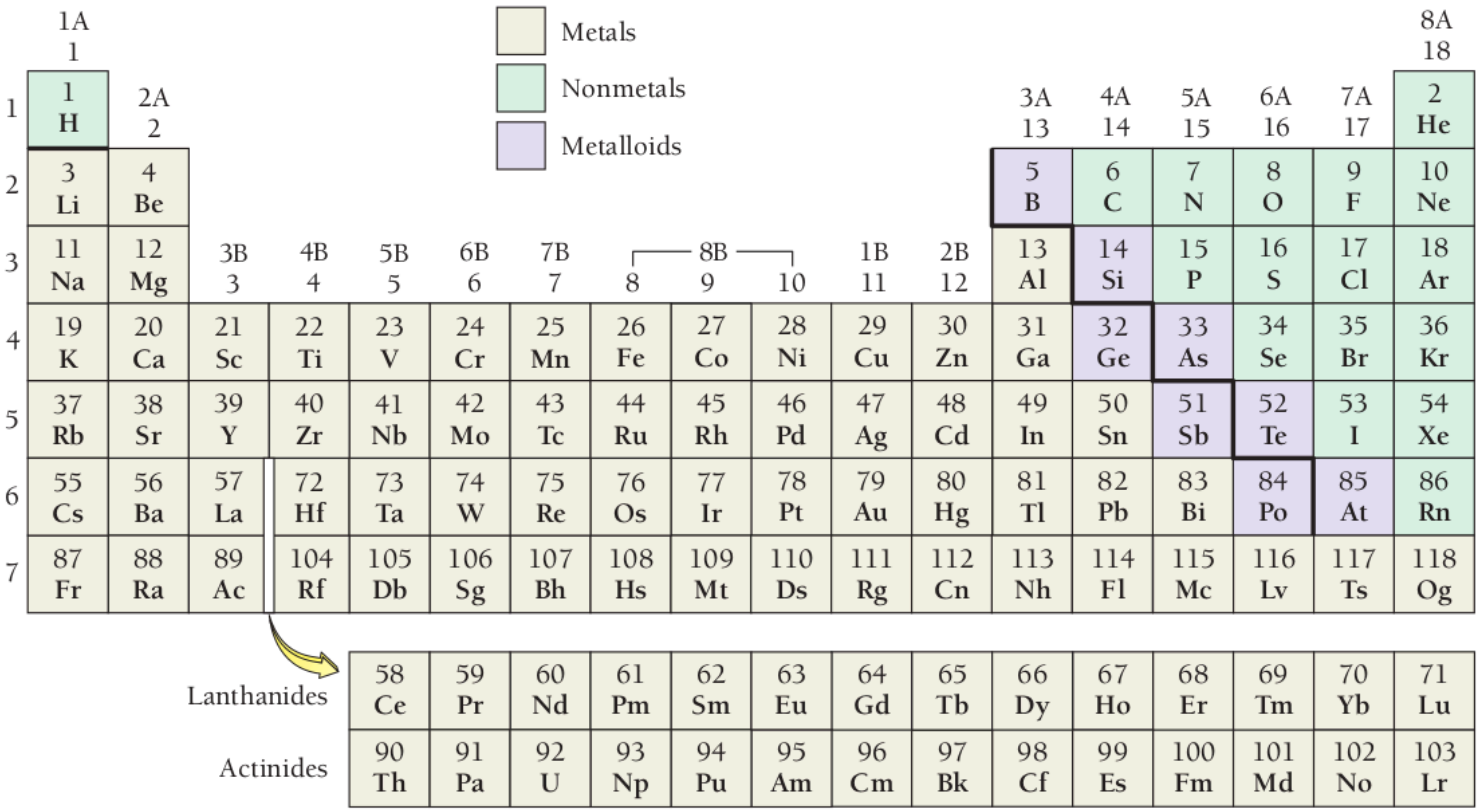
How many electrons does a neutral neon atom have?

Will an atom with 6 protons, 6 neutrons, and 6 electrons be electrically neutral?

Will an atom with 27 protons, 32 neutrons, and 27 electrons be electrically neutral?

Will an Na atom with 10 electrons be electrically neutral?

Periodicity



Metals

- Solids at room temperature, except Hg.
- Reflective surface.
- Conduct heat.
- Conduct electricity.
- Malleable..
- Ductile.
- Lose electrons and form cations in reactions.

- About 75% of the elements are metals.
- Lower left on the table.

Nonmetals

- Found in all 3 states at standard temperature and pressure.
- Poor conductors of heat.
- Poor conductors of electricity.
- Solids are brittle.
- Gain electrons in reactions to become anions.
- Upper right on the table.
    - Except H.

Metalloids

- Show some properties of metals and some of nonmetals.
- Also known as semiconductors.

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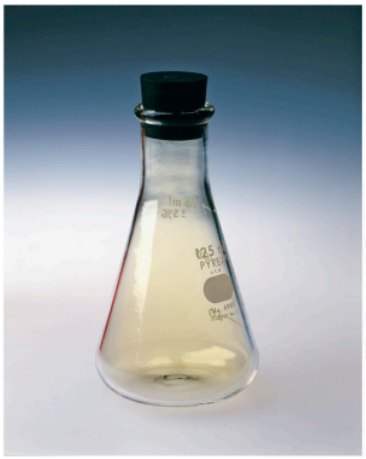

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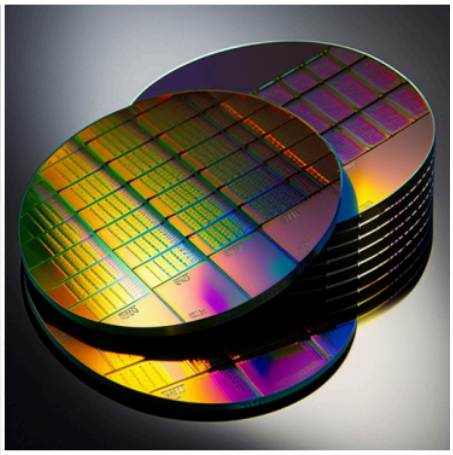
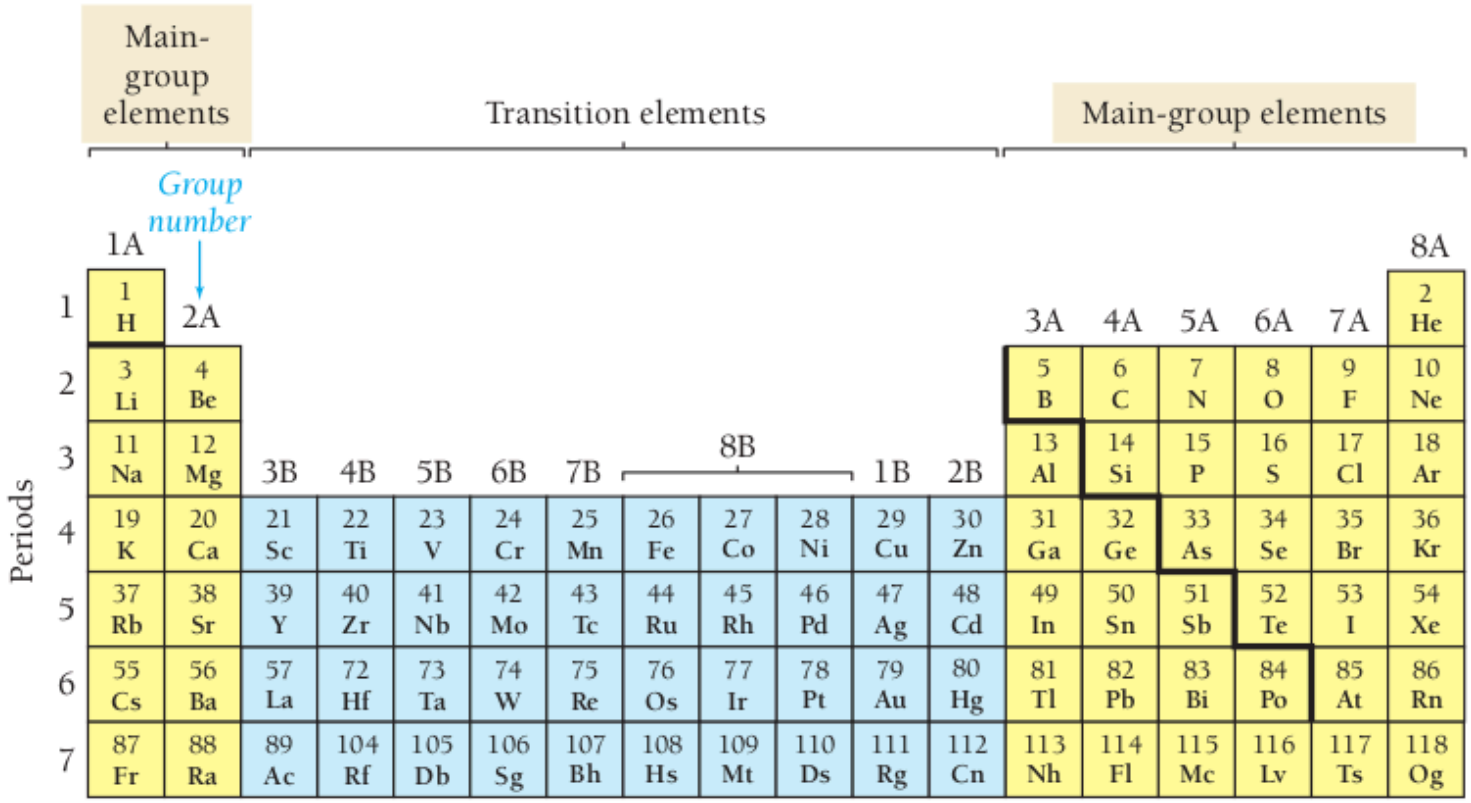
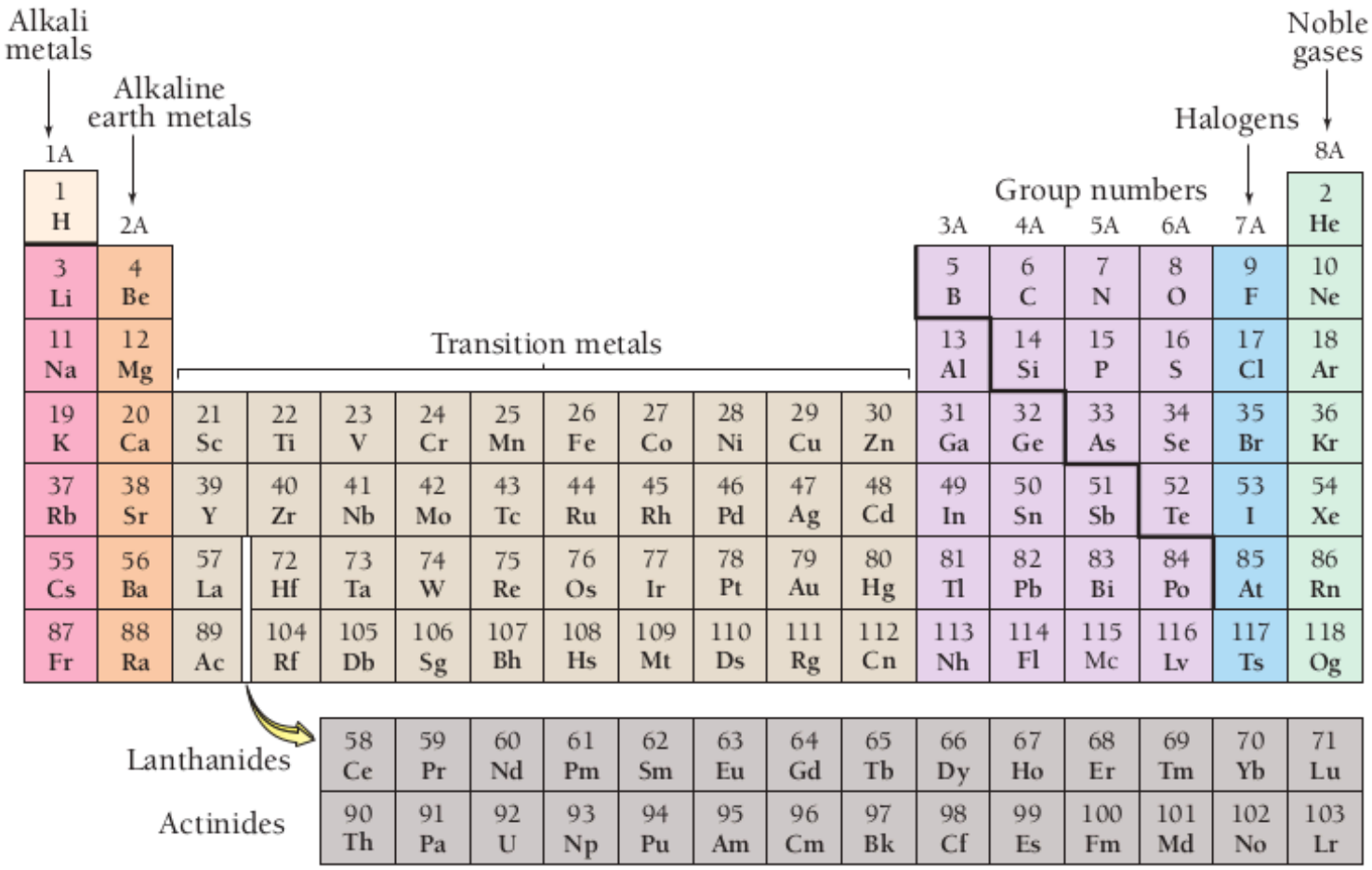


Classify Each Element as Metal, Nonmetal, or Metalloid.

- Xenon, Xe
- Tungsten, W
- Bromine, Br
- Arsenic, As
- Cerium, Ce

Groups





Charge and Ions

- In a chemical change, the number of protons in the nucleus of the atom doesn't change.
    - Radioactive and nuclear changes are an exception
- Atoms in a compound are often electrically charged, these are called ions.
- Atoms acquire a charge by gaining or losing electrons.

- Never protons!

$$\text{Ionic Charge} = \#\text{ protons} - \#\text{ electrons}$$

$$q = n_{p^+} - n_{e^-}$$

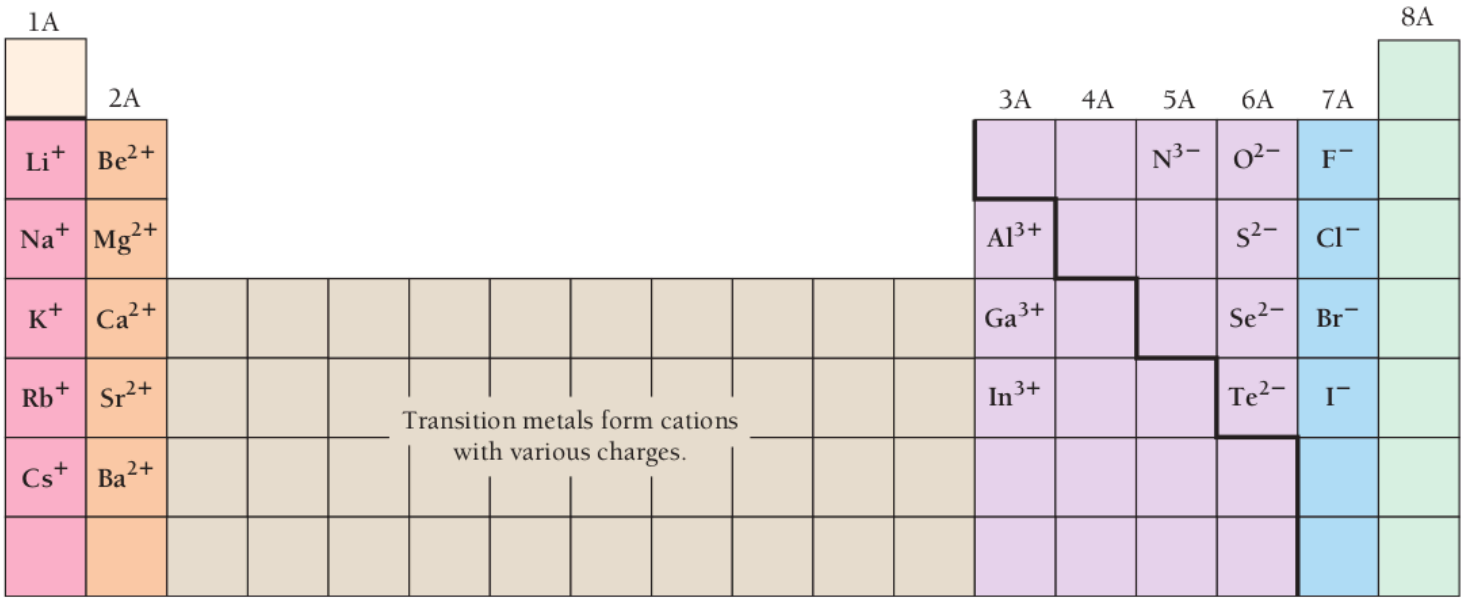
- Ions with a positive charge are called **cations**.
  - Metals
  - More protons than electrons.
  - Form by losing electrons.
- Ions with a negative charge are called **anions**.
  - Nonmetals
  - More electrons than protons.
  - Form by gaining electrons.
- Chemically, ions are much different than the neutral atoms.
- Anions are named by changing the ending of the name to –ide.

$$\begin{array}{rcl} \text{Fluorine} & \mathrm{F} + 1\,\mathrm{e}^- \longrightarrow \mathrm{F}^- & \text{Fluoride Ion} \\ \text{Oxygen} & \mathrm{O} + 2\,\mathrm{e}^- \longrightarrow \mathrm{O}^{2-} & \text{Oxide Ion} \end{array}$$

- Cations are named the same as the metal.

$$\begin{array}{rcl} \text{Sodium} & \mathrm{Na} \longrightarrow \mathrm{Na}^+ + 1\,\mathrm{e}^- & \text{Sodium Ion} \\ \text{Calcium} & \mathrm{Ca} \longrightarrow \mathrm{Ca}^{2+} + 2\,\mathrm{e}^- & \text{Calcium Ion} \end{array}$$

- The charge on a ion can often be determined from the group number on the periodic table.



Fill in the Table

Ion	p^+	e^-
Cl^{1-}		
K^{1+}		
S^{2-}		
Sr^{2+}		
Ca^{2+}		

Valence Electrons and Ion Charge

- The highest energy electrons in an atom are called the valence electrons.
- Metals form cations by losing their valence electrons to get the same number of electrons as the previous noble gas.
  - Main group metals.
  - Li^{+} has the same number of electrons as He
  - Al^{3+} has the same number of electrons as Ne
- Nonmetals form anions by gaining electrons to have the same number of electrons as the next noble gas.
  - Cl^{-} has the same number of electrons as Ar
  - Se^{2-} has the same number of electrons as Kr

Isotopes

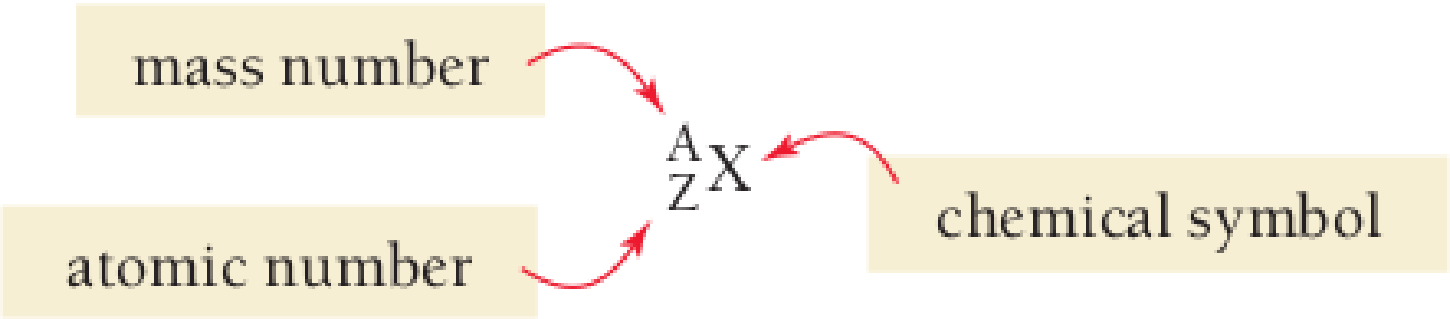
- Soddy discovered that the same element could have atoms with different masses, which he called isotopes.
  - There are two isotopes of chlorine found in nature, one that has a mass of about 35 amu and another that weighs about 37 amu.

- The **atomic mass** is a *weighted average* of the weights of all the naturally occurring atoms.
    - The atomic mass of chlorine is 35.45 amu.
- The **exact mass** is the mass of a specific isotope
- All isotopes of an element are chemically identical.
- All isotopes of an element have the same number of protons and a different number of neutrons.
    - Isotopes of an element have different masses.
- Isotopes are identified by their **mass numbers**.

$$\text{Mass Number} = \#\ \text{Protons} + \#\ \text{Neutrons}$$

- Unlike the atomic mass or the exact mass, *mass number* is always a whole number

Isotopes



$^{20}_{10}Ne$ $^{21}_{10}Ne$ $^{22}_{10}Ne$

X - A

chemical symbol or name

mass number

Ne-20 neon-20

Ne-21 neon-21

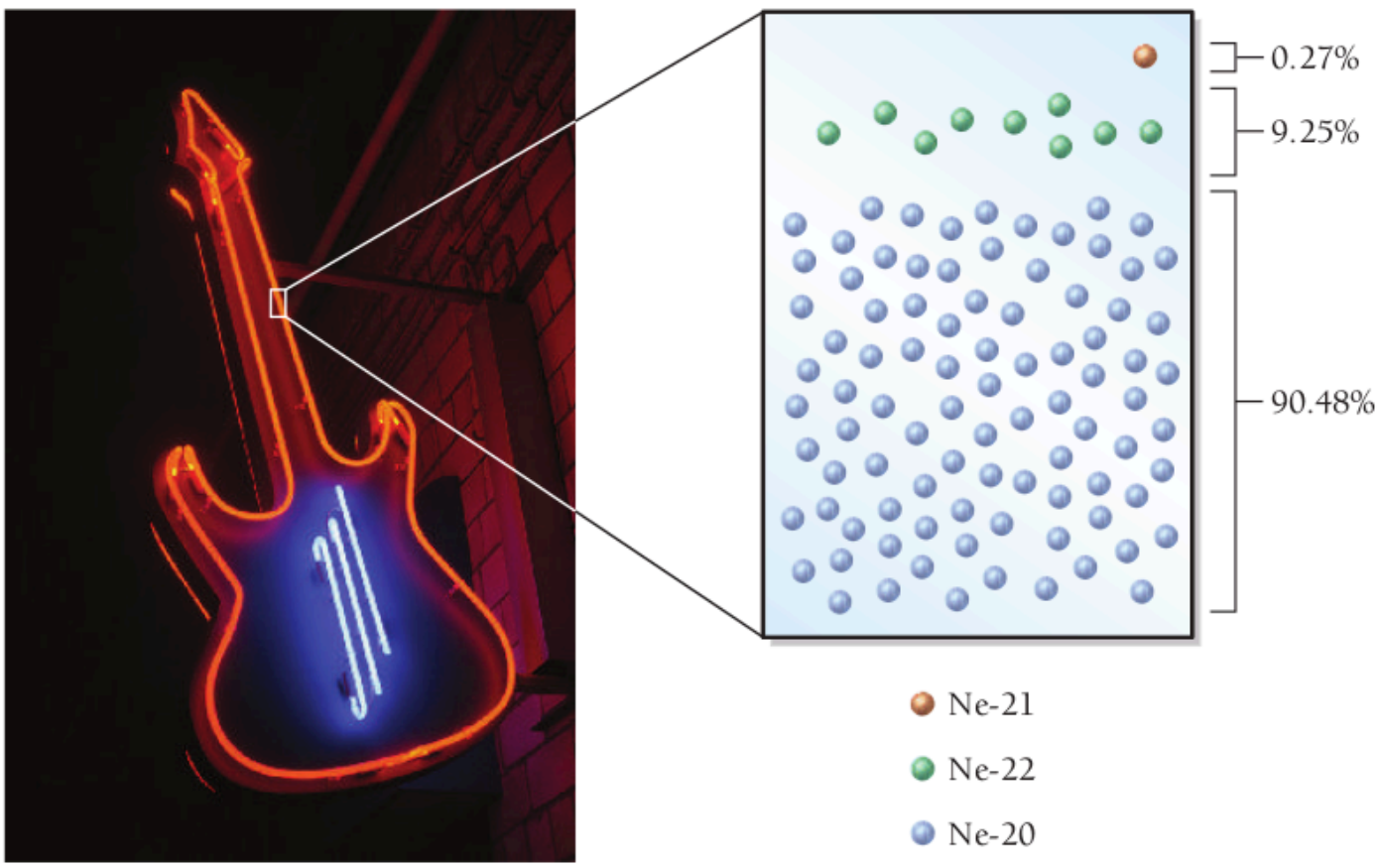
Ne-22 neon-22

- Each isotope has a **natural abundance** based on the relative amount of the isotope found in nature
  - Natural abundance is the probability of finding a particular isotope in a sample of an element

- Cl-35 makes up about 75% of chlorine atoms in nature, and Cl-37 makes up the remaining 25%.

TABLE 4.2 Neon Isotopes

Symbol	Number of Protons	Number of Neutrons	A (Mass Number)	Percent Natural Abundance
Ne-20 or $^{20}_{10}Ne$	10	10	20	90.48%
Ne-21 or $^{21}_{10}Ne$	10	11	21	0.27%
Ne-22 or $^{22}_{10}Ne$	10	12	22	9.25%



What is the atomic mass of Neon?

Isotope	Atomic Mass	Natural Abundance (%)
^{20}Ne	19.992	90.48
^{21}Ne	20.994	0.27
^{22}Ne	21.991	9.25

How Many Protons and Neutrons Are in an Atom of $^{52}_{24}\mathrm{Cr}$?

Gallium has two naturally occurring isotopes. Ga-69 with Mass 68.9256 Amu and Abundance of 60.11% and Ga-71 with Mass 70.9247 Amu and Abundance of 39.89%. Calculate the Atomic Mass of Gallium.

If Copper Is 69.17% Cu-63 with a Mass of 62.9396 Amu and the Rest Cu-65 with a Mass of 64.9278 Amu, Find Copper’s Atomic Mass.

Chapter 4 Quiz

1. The atom is the fundamental building block of matter. (1 Point)

   True False

2. A molecule is the smallest identifiable unit of a compound. (1 Point)

   True False

3. You can continually divide matter into smaller and smaller pieces without ever coming to an end. (1 Point)

   True False

4. The gold foil experiment proved that atoms contain a dense nucleus. (1 Point)

   True False

5. Protons and neutrons each have a mass of approximately 1 amu. (1 Point)

   True False

6. A positive charge attracts negative charges and repels other positive charges. (1 Point)

   True False

7. If two atoms each contain different numbers of neutrons, the atoms must be from different elements. (1 Point)

   True False

8. In the modern periodic table, elements are listed in order of increasing atomic number rather than increasing relative mass. (1 Point)

True False

9. Metals are located on the upper-right side of the periodic table. (1 Point)

True False

10. A cation forms when an atom losses an electron. (1 Point)

True False

11. The atomic mass of individual atoms of an element may vary. (1 Point)

True False

12. Which statement below accurately describes the contributions of Thomson? (2 Points)

A) discovered the existence of protons
B) created the modern periodic table
C) proposed the atom contained a dense nucleus
D) discovered the existence of electrons
E) none of the above

14. Which statement reflects the results of Rutherford's gold foil experiments? (2 Points)

A) Almost all of the alpha particles were deflected back in the direction from which they came.
B) Almost all of the alpha particles sputtered gold atoms off of the surface of the foil.
C) Almost all of the alpha particles were deflected while passing through the foil.
D) Almost all of the alpha particles passed directly through the foil.
E) none of the above

15. An atom containing 12 protons, 10 neutrons, and 12 electrons (2 Points)

A) is charge-neutral.
B) is an ion.
C) is an oxygen atom.
D) cannot exist.
E) none of the above

16. Which of the following elements has an atomic number of 6? (2 Points)

A) H
B) C
C) He
D) Be
E) none of the above

17. Which of the following elements has 16 protons? (2 Points)

A) C

B) Zn

C) Mg

D) O

E) none of the above

18. What is the atomic symbol for goldr? (2 Points)

A) S

B) Ag

C) Au

D) Si

E) none of the above

19. What is the atomic symbol for silicon? (2 Points)

A) Sn

B) Ti

C) Tn

D) Si

E) none of the above

20. Ne is a member of which family? (2 Points)

A) noble gases

B) halogens

C) alkaline earth metals

D) alkali metals

E) none of the above

21. Identify the element that is a nonmetal, a gas, and has an elemental symbol that starts with the letter "A." (2 Points)

A) Ac

B) Ar

C) Au

D) Al

E) none of the above

22. Ions are formed when atoms (2 Points)

A) gain or lose protons.

B) gain or lose electrons.

C) gain or lose neutrons.

D) Each of these results in ion formation.

E) None of these results in ion formation.

23. What is the correct formula for sodium ion with 10 electrons? (2 Points)

A) S^+

B) K+

C) Na^+

D) So^-

E) none of the above

24. How many electrons are in Br^- ? (2 Points)

A) 4

B) 7

C) 34

D) 36

E) none of the above

25. How many electrons does O^{2-} have? (2 Points)

A) 2

B) 16

C) 10

D) 8

E) none of the above

26. How many protons and neutrons are in Cl-37? (2 Points)

A) 20 protons, 17 neutrons

B) 17 protons, 37 neutrons

C) 17 protons, 20 neutrons

D) 37 protons, 17 neutrons

E) none of the above

27. A specific isotope of an element is known to have 35 protons and 45 neutrons. Which symbol would properly represent this isotope? (2 Points)

A) Br-80

B) Hg-45

C) Hg-35

D) Br-45

E) none of the above

Chapter 5

Molecules and Compounds

- Compounds have chemical and physical properties distinct from their component elements.
  - Salt
    - Sodium—shiny, reactive, poisonous.
    - Chlorine—pale yellow gas, reactive, poisonous.
    - Sodium chloride—table salt.
  - Sugar
    - Carbon—pencil or diamonds.
    - Hydrogen—flammable gas.
    - Oxygen—a gas in air.

Law of Constant Composition

- All samples of a pure substance contain the same elements in the same percentages (ratios).
  - The smallest piece of a compound is called a molecule.

    - Every molecule of a compound has the same number and type of atoms.
    - Since all the molecules of a compound are identical, every sample will have the same
      ratio of the elements.
    - Since all molecules of a compound are identical, every sample of the compound will have the same properties.
- Mixtures have variable composition.

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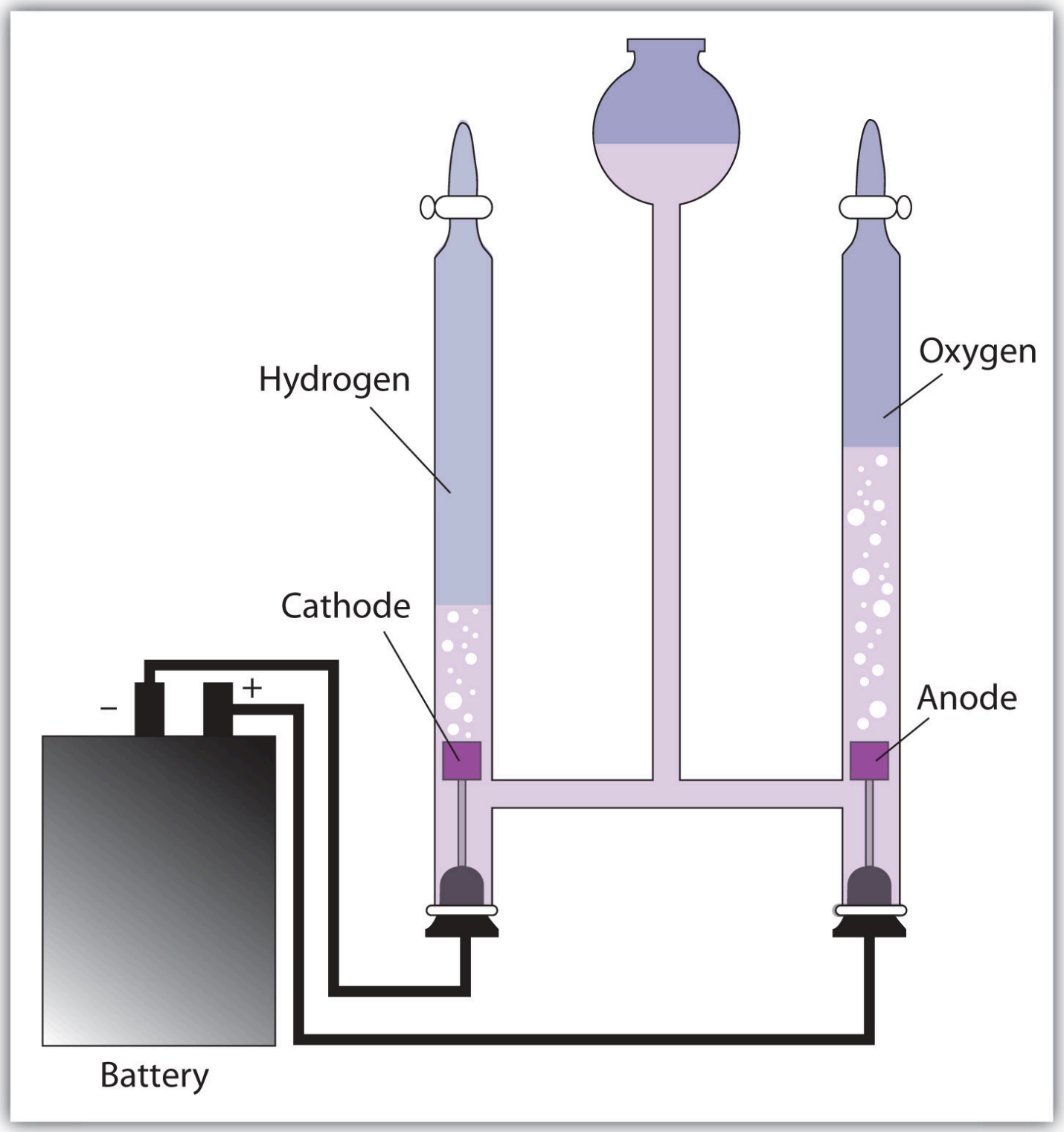
- If we decompose water by electrolysis, we get 16.0 grams of oxygen to every 2.00 grams of hydrogen.



- Water has a constant mass ratio of oxygen to hydrogen of 8.0

$$\text{Mass Ratio} = \frac{m_O}{m_H} = \frac{16.0\ \text{g}}{2.0\ \text{g}} = 8.0$$

Show that Two Samples of Carbon Dioxide Are Consistent with the Law of Constant Composition.

Sample	Carbon (g)	Oxygen (g)
1	1.8	4.8
2	6.4	17.1

Show that Hematite Has Constant Composition if a 10.0 g Sample Has 7.2 g Fe and the Rest Is Oxygen; and a Second Sample Has 18.1 g Fe and 6.91 g O.

Polyatomics

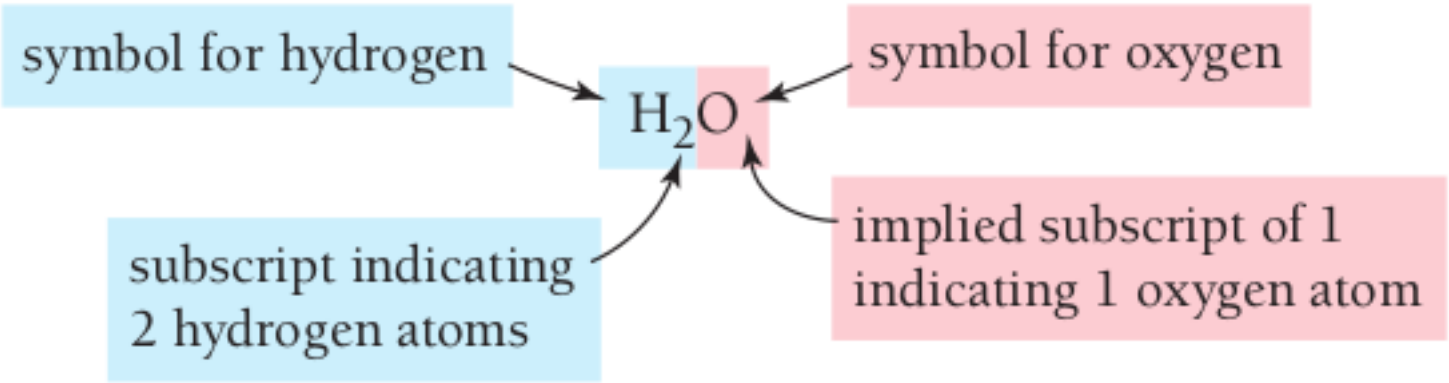
Certain groups of atoms are bonded together to form what is called a polyatomic ion that acts as a single unit

TABLE 5.3 Some Common Polyatomic Ions

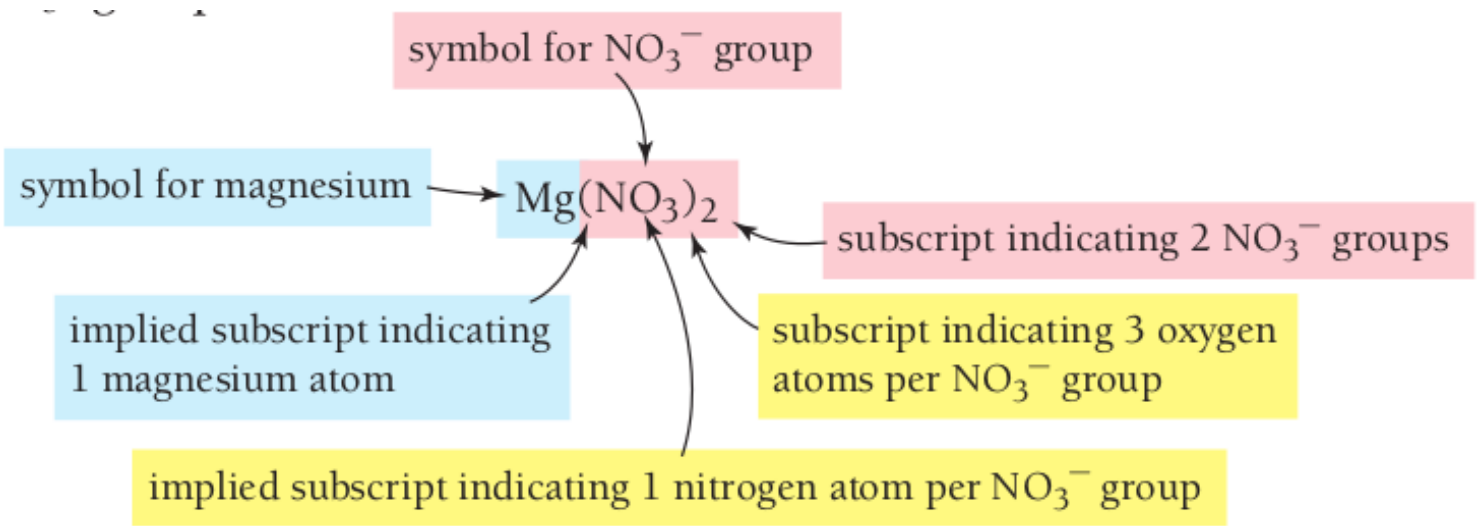
Name	Formula	Name	Formula
acetate	$C_2H_3O_2^-$	hypochlorite	ClO^-
carbonate	CO_3^{2-}	chlorite	ClO_2^-
hydrogen carbonate (or bicarbonate)	HCO_3^-	chlorate	ClO_3^-
hydroxide	OH^-	perchlorate	ClO_4^-
nitrate	NO_3^-	permanganate	MnO_4^-
nitrite	NO_2^-	sulfate	SO_4^{2-}
chromate	CrO_4^{2-}	sulfite	SO_3^{2-}
dichromate	$Cr_2O_7^{2-}$	hydrogen sulfite (or bisulfite)	HSO_3^-
phosphate	PO_4^{3-}	hydrogen sulfate (or bisulfate)	HSO_4^-
hydrogen phosphate	HPO_4^{2-}	peroxide	O_2^{2-}
ammonium	NH_4^+	cyanide	CN^-

Molecular Formulas Describe Compounds

- We describe the compound by describing the number and type of each atom in the simplest unit of the compound.
  - Molecules or ions.
- Each element is represented by its letter symbol.
- The number of atoms of each element is written to the right of the element as a subscript.
  - If there is only one atom, the 1 subscript is not written.



- **Polyatomic groups** are placed in parentheses.
  - If more than one.



- Order of Elements in a Formula
  - Metals are written first.
  - Nonmetals are written in order
- There are occasional exceptions for historical or informational reasons.
  - H_2O, but NaOH .

TABLE 5.1 Order of Listing Nonmetal Elements in a Chemical Formula									
C	P	N	H	S	I	Br	Cl	O	F

Elements on the left are generally listed before elements on the right.

Hematite is composed of four oxide ions for every three iron ions. What is the chemical formula for hematite?

Acetone molecules contain six hydrogen atoms, three carbon atoms, and one oxygen atom. What is it's chemical formula?

Determine the Total Number of Atoms or Ions in One Formula Unit of Each of the Following.

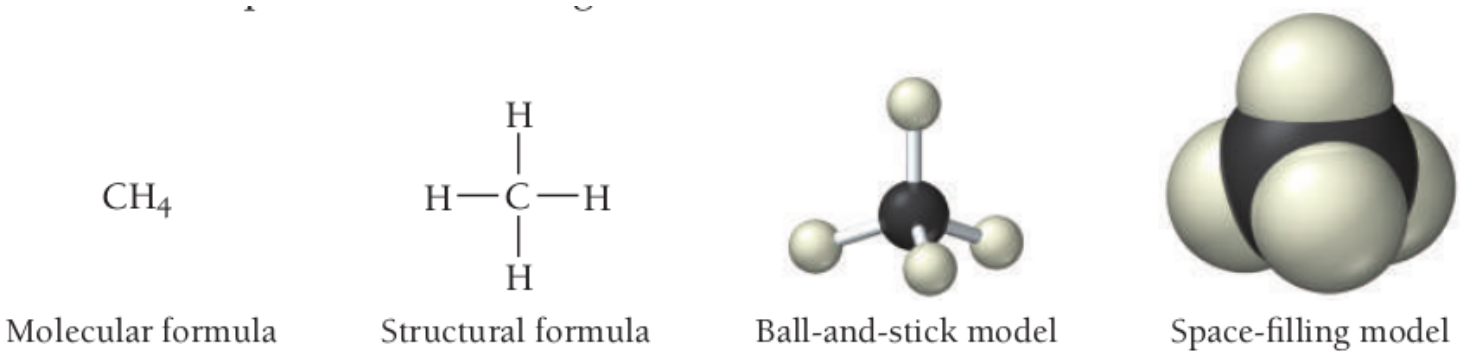
$Mg(C_2H_3O_2)_2$

$Hg_6(PO_4)_2$

$(NH_4)_3PO_4$

Structural Formulas

- **Structural formulas** use lines to represent chemical bonds
- Unlike **molecular formulas**, structural formulas demonstrate how the atoms in a molecular are connected.

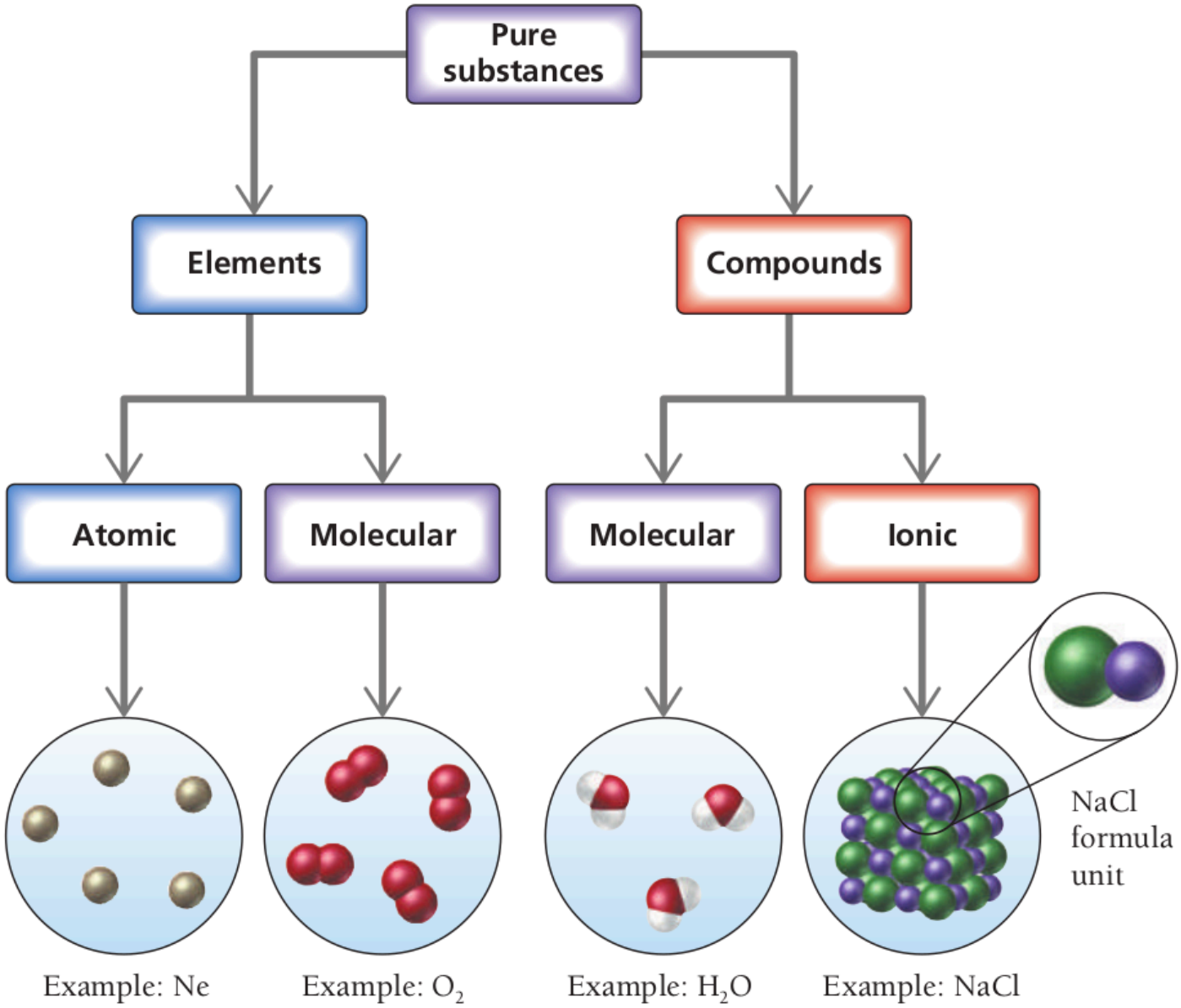


Empirical Formulas

- An **empirical formula** is the simplest whole-number ratio of atoms of each element in a compound.

Molecular Formula	Empirical Formula
C_6H_6	CH
$C_2O_4H_2$	CO_2H
$Al(NO_3)_3$	$Al(NO_3)_3$

Classifying Materials



Atomic Elements

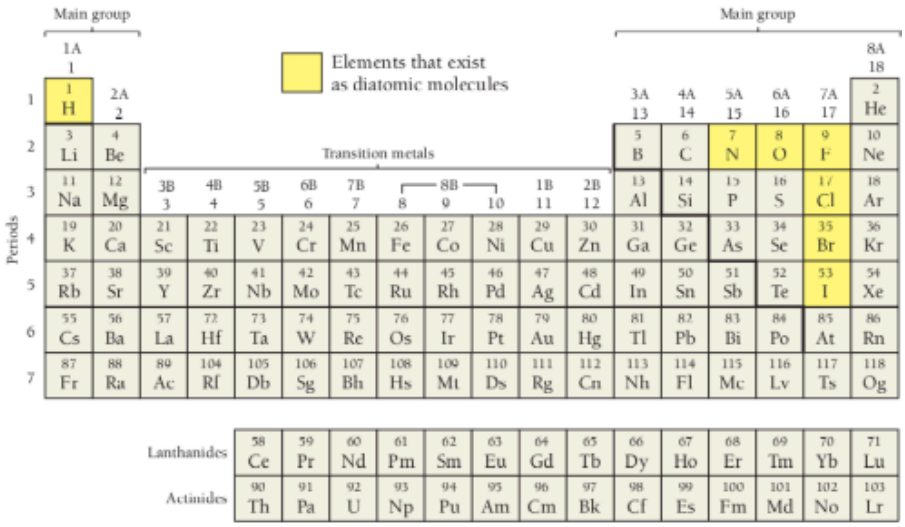
- Atomic elements have single atoms as their basic units.
    - Most elements fall into this category

Molecular Elements

- Molecular elements do not normally exist in nature with single atoms as their basic units.
  - Smallest unit is a molecule.
  - Two or more nonmetals.
  - These elements usually exist as diatomic molecules.

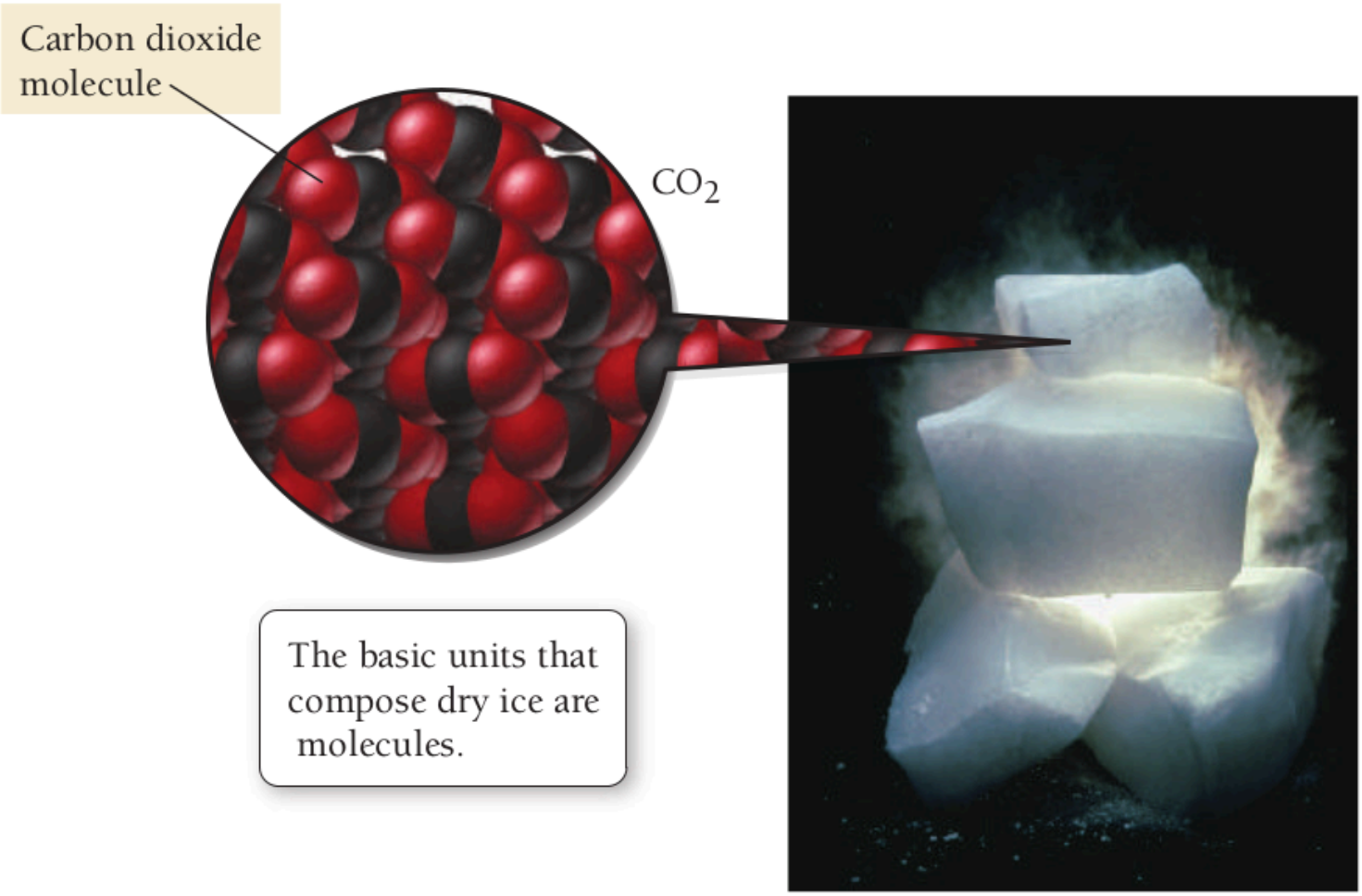
TABLE 5.2 Elements That Occur as Diatomic Molecules

Name of Element	Formula of Basic Unit
hydrogen	H_2
nitrogen	N_2
oxygen	O_2
fluorine	F_2
chlorine	Cl_2
bromine	Br_2
iodine	I_2



Molecular Compounds

- **Molecular compounds** are composed of two or more nonmetals.
  - The basic units are molecules.
  - For instance, H_2O, CO_2, C_3H_6O

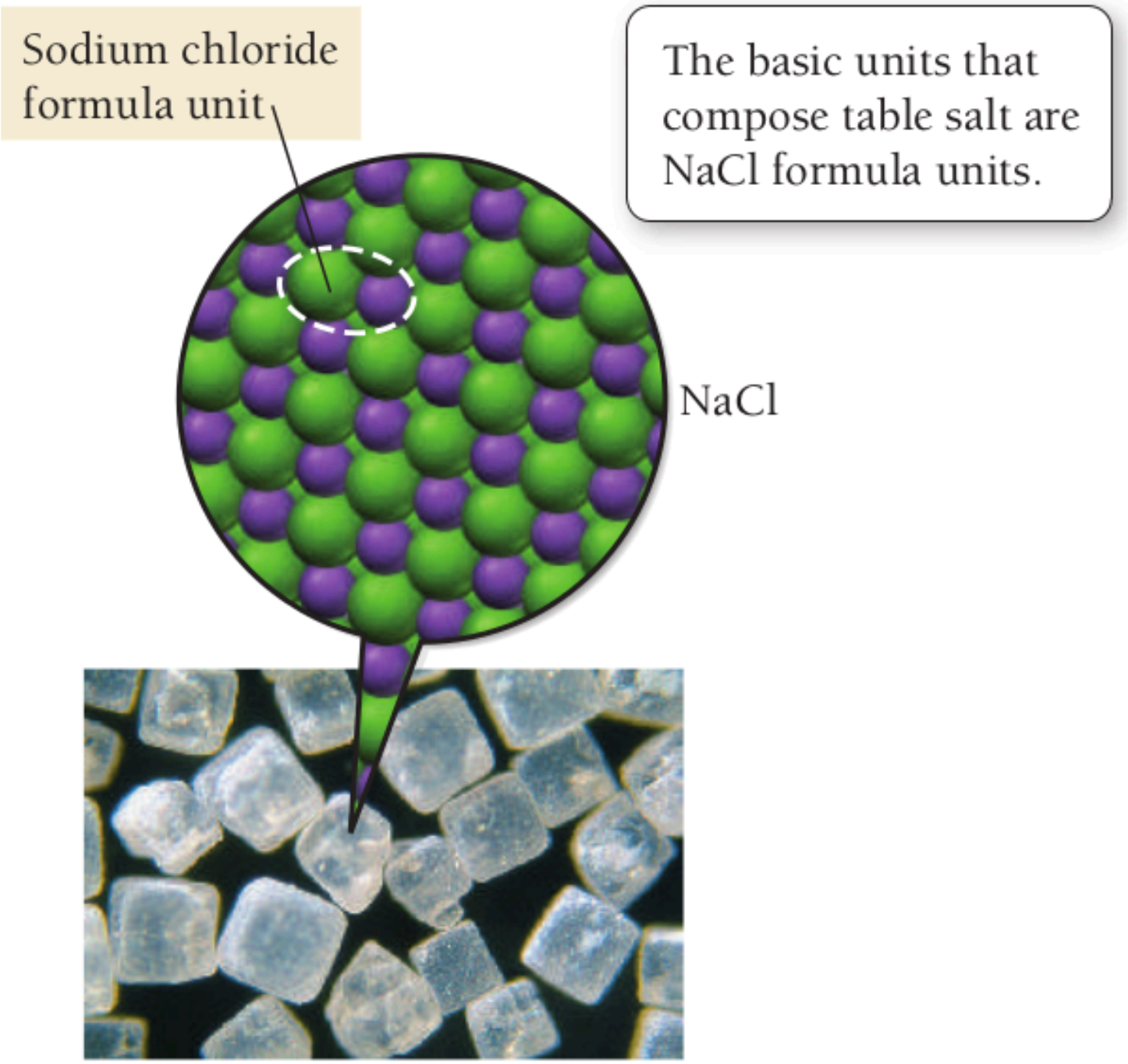


Ionic Compounds

- **Ionic compounds** are composed of one or more cations (+) paired with one or more
  anions (-)
- Usually metals + nonmetals.
- The basic unit of ionic compounds is the **formula unit**.
- Smallest electrically neutral collection of ions

-

- No real individual units, instead have a 3-dimensional array of cations and anions.



Classify Each of the Following as Either an Atomic Element, Molecular Element, Molecular Compound, or Ionic Compound.

Aluminum, Al.

Aluminum chloride, $AlCl_3$.

Chlorine, Cl_2.

Acetone, C_3H_6O.

Carbon monoxide, CO.

Cobalt, Co.

Writing Ionic Formulas

Ionic compounds are electrically neutral therefore there must be an equal number of positive and negative charges

$$\text{Total Positive Charge} + \text{Total Negative Charge} = 0$$

We know sodium (Na) tends to form a cation with a 1+ charge

$$\mathrm{Na} \longrightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$$

We also know that sulfur tends to form an anion with a 2- charge

$$\mathrm{S} + 2\,\mathrm{e}^- \longrightarrow \mathrm{S}^{2-}$$

To achieve an electrically neutral ionic compound, we will need two sodium ions for each sulfide ion.

$$2\,\mathrm{Na}^+ + \mathrm{S}^{2-} \longrightarrow \mathrm{Na_2S}$$

Rules

1. Write the symbol for the metal cation and its charge.
2. Write the symbol for the nonmetal anion and its charge.
3. Charge (without sign) becomes subscript for the other ion.
4. Reduce subscripts to smallest whole-number ratio.
5. Check that the sum of the charges of the cation cancels the sum of the anions.

Al^{+3} (Group 13)

O^{2-} (Group 16)

$Al^{+3}\ O^{2-}$

Al_2O_3

$Al = (2)\cdot(+3) = +6$

$O = (3)\cdot(-2) = -6$

What Are the Formulas for Compounds Made from the Following Ions?

- Potassium ion with a nitride ion
- Calcium ion with a bromide ion
- Aluminum ion with a sulfide ion
- Magnesium ion with sulfite ion
- Copper ion with a chloride ion
- Ammonium ion with nitrate ion

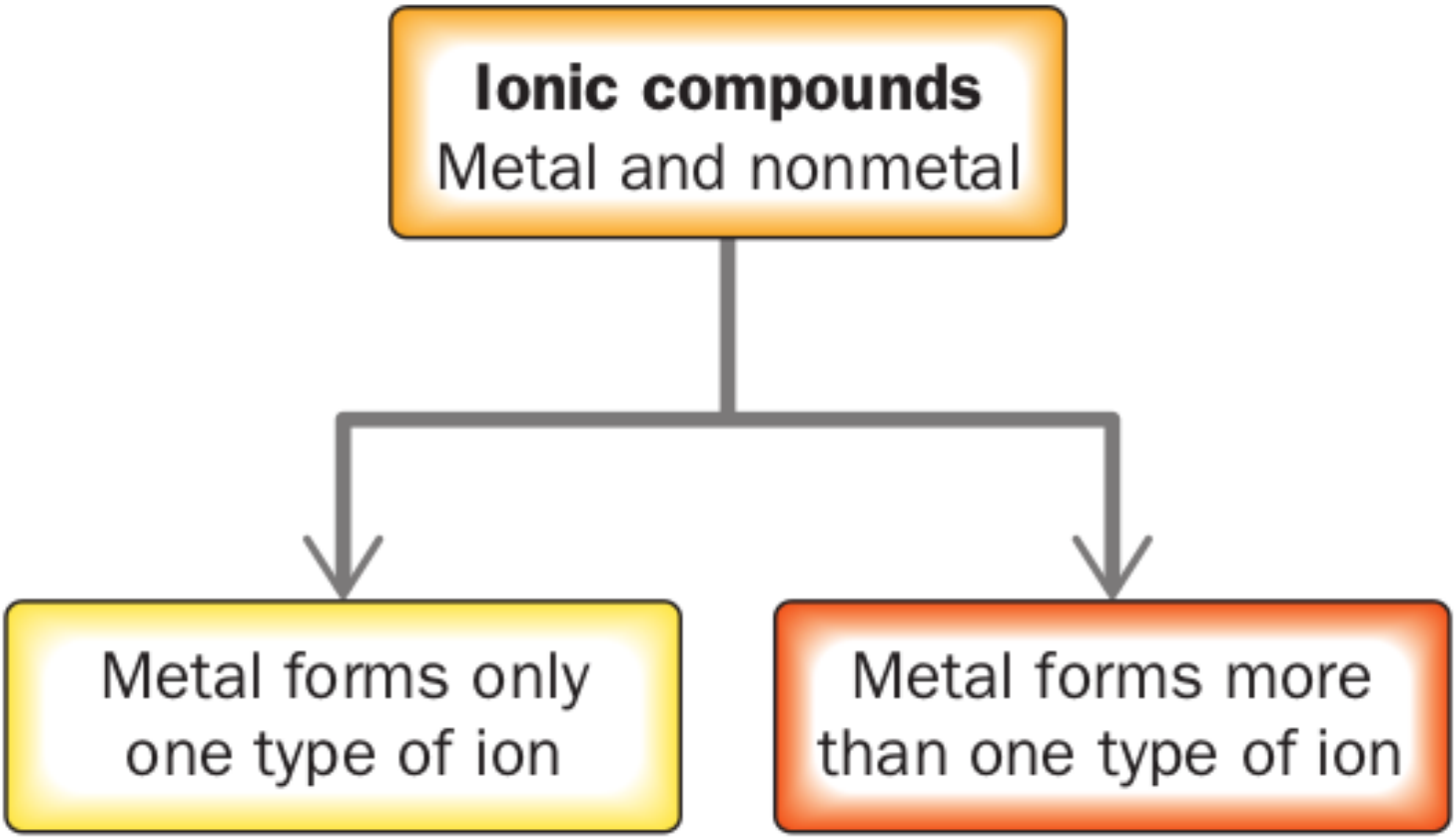
Naming

Common Names

Some compounds have common names mostly due to historic significance

- H_2O = Water, steam, ice.
- NH_3 = Ammonia.
- CH_4 = Methane.
- NaCl = Table salt.
- $C_{12}H_{22}O_{11}$ = Table sugar.

Ionic compounds.



Metals Whose Charge Is Invariant from One Compound to Another

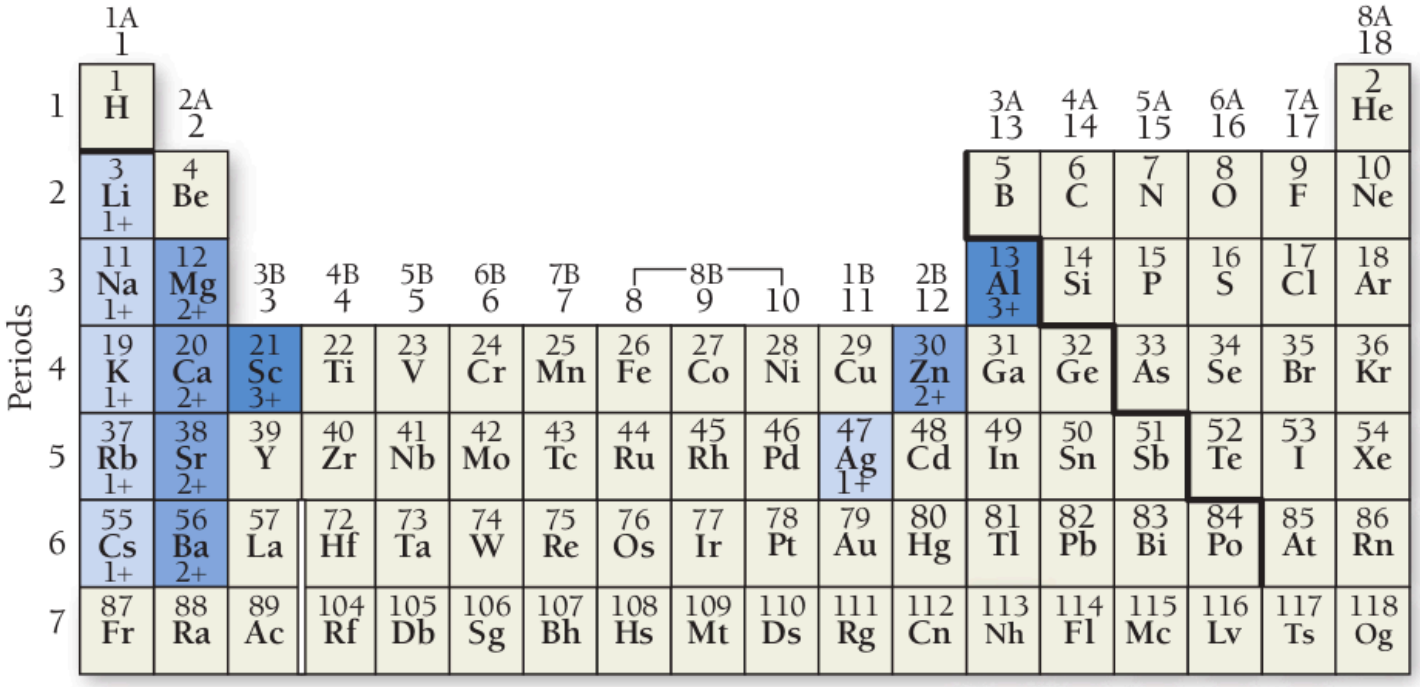


TABLE 5.4 Some Metals That Form More Than One Type of Ion and Their Common Charges (*This list is not exhaustive but meant to show examples.*)

Metal	Symbol Ion	Name	Older Name*
chromium	Cr^{2+} Cr^{3+}	chromium(II) chromium(III)	chromous chromic
iron	Fe^{2+} Fe^{3+}	iron(II) iron(III)	ferrous ferric
cobalt	Co^{2+} Co^{3+}	cobalt(II) cobalt(III)	cobaltous cobaltic
copper	Cu^{+} Cu^{2+}	copper(I) copper(II)	cuprous cupric
tin	Sn^{2+} Sn^{4+}	tin(II) tin(IV)	stannous stannic
mercury	Hg_2^{2+} Hg^{2+}	mercury(I) mercury(II)	mercurous mercuric
lead	Pb^{2+} Pb^{4+}	lead(II) lead(IV)	plumbous plumbic

TABLE 5.5 Some Common Anions

Nonmetal	Symbol for Ion	Base Name	Anion Name
fluorine	F^-	fluor-	fluoride
chlorine	Cl^-	chlor-	chloride
bromine	Br^-	brom-	bromide
iodine	I^-	iod-	iodide
oxygen	O^{2-}	ox-	oxide
sulfur	S^{2-}	sulf-	sulfide
nitrogen	N^{3-}	nitr-	nitride

Binary Ionic

Type 1

name of cation (metal)

base name of anion (nonmetal) + *-ide*

Type 2

name of cation (metal) (charge of cation (metal) in roman numerals in parentheses) base name of anion (nonmetal) + *-ide*

When the anion is a polyatomic, the suffix is not changed

Write the name for the following ionic compounds

- KCl
- Na_2O
- $CaBr_2$
- CoF_2
- $CuCl$
- $Mg(NO_2)_2$
- Li_2SO_4
- $(NH_4)_3PO_4$
- $Al_2(SO_3)_3$

Write the formula for the following ionic compounds

- Copper(II) Bromide
- Iron(III) fluoride
- Calcium Sulfate
- Lithium Phosphate
- Sodium Oxide

Molecular compounds.

- 2 or more nonmetals

mono- 1	*hexa-* 6
di- 2	*hepta-* 7
tri- 3	*octa-* 8
tetra- 4	*nona-* 9
penta- 5	*deca-* 10

prefix	name of 1st element	prefix	base name of 2nd element + *-ide*

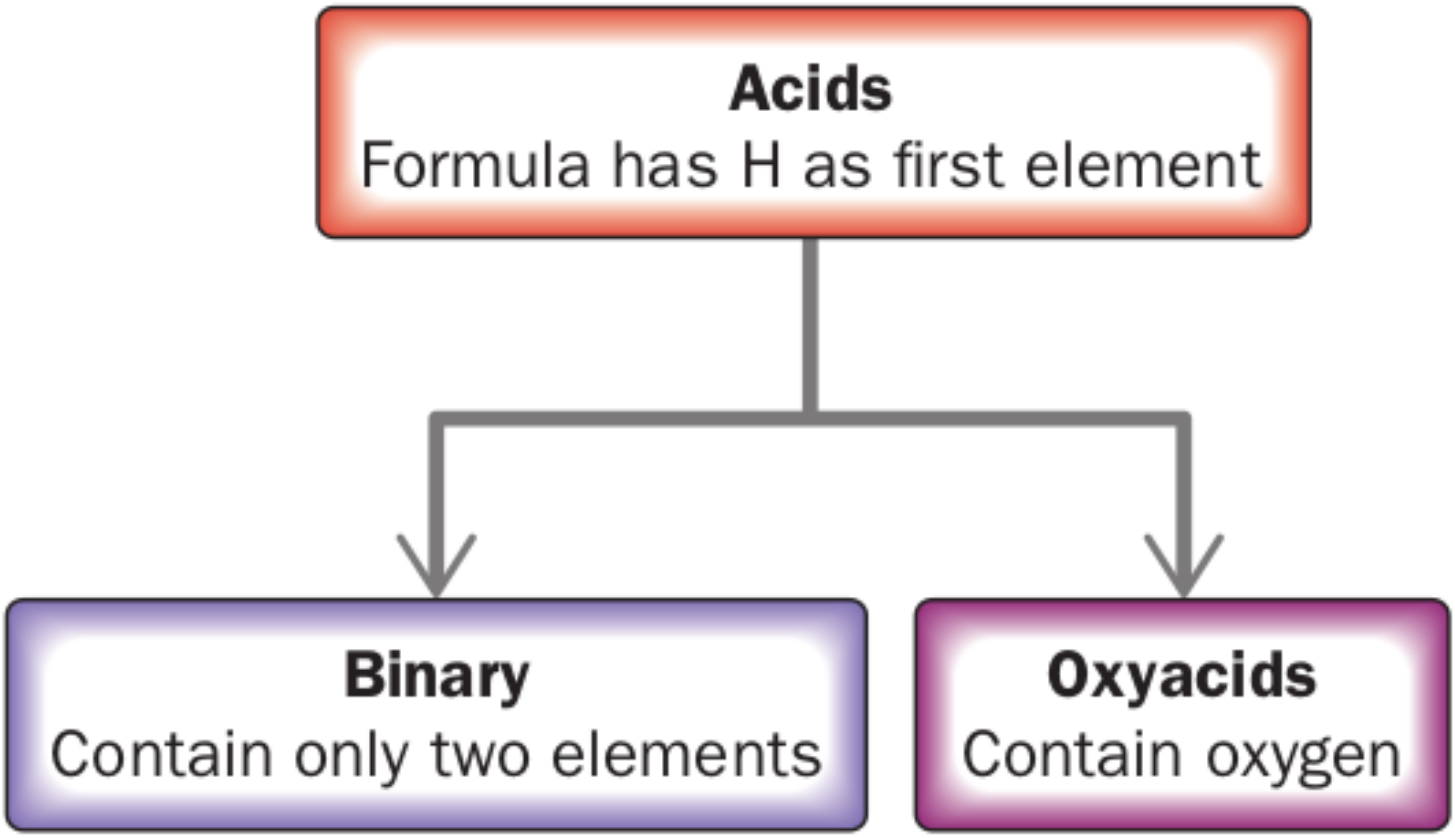
Acids

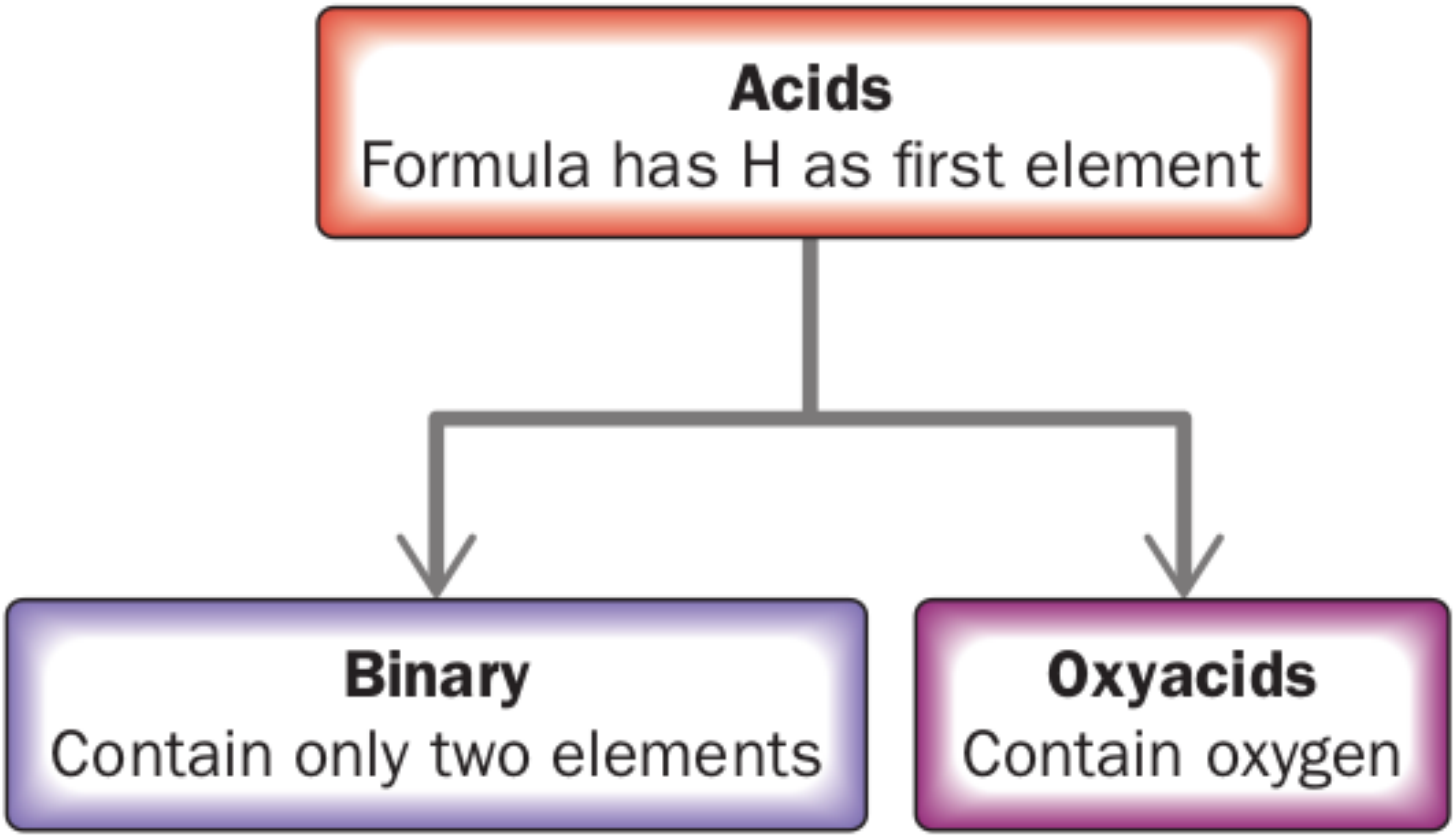
- Formula starts with hydrogen (H).
  - H_2SO_4, HBr
- Sour taste

- Though acids are molecular, they behave as ionic when dissolved in water.

$$HA \longrightarrow H^+ + A^-$$

- May be binary or oxyacid.



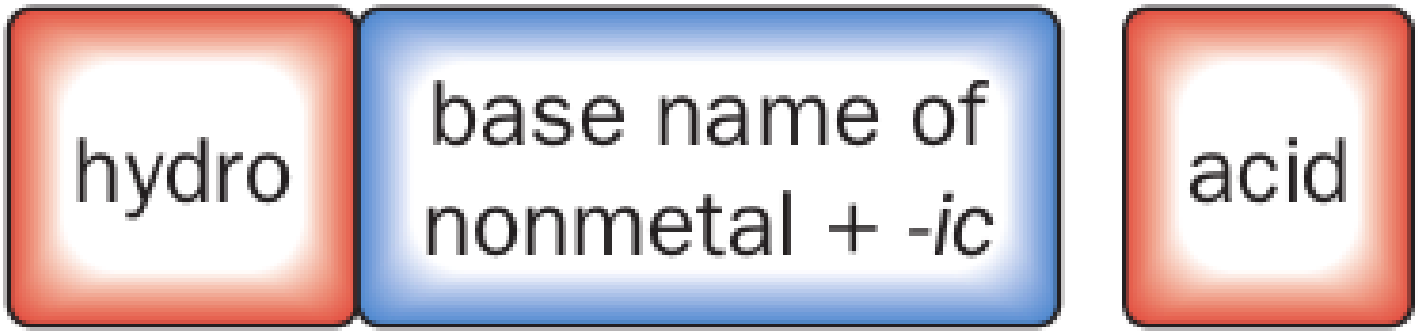


$$HCl \longrightarrow H^+ + Cl^-$$

$$HNO_3 \longrightarrow H^+ + NO_3^-$$

Binary Acids

- Binary acids have H+1 cation and nonmetal anion

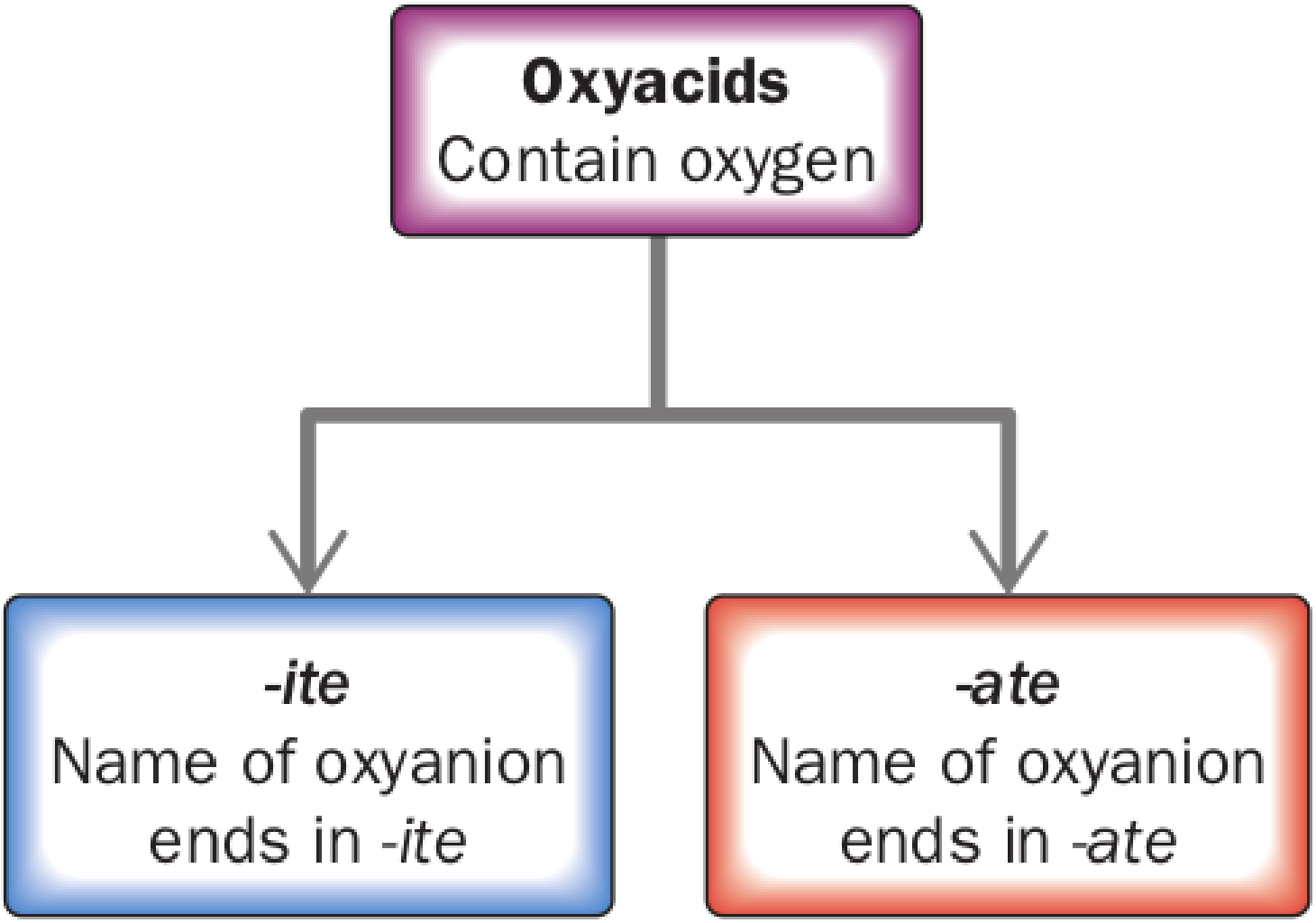


Name the following binary acids

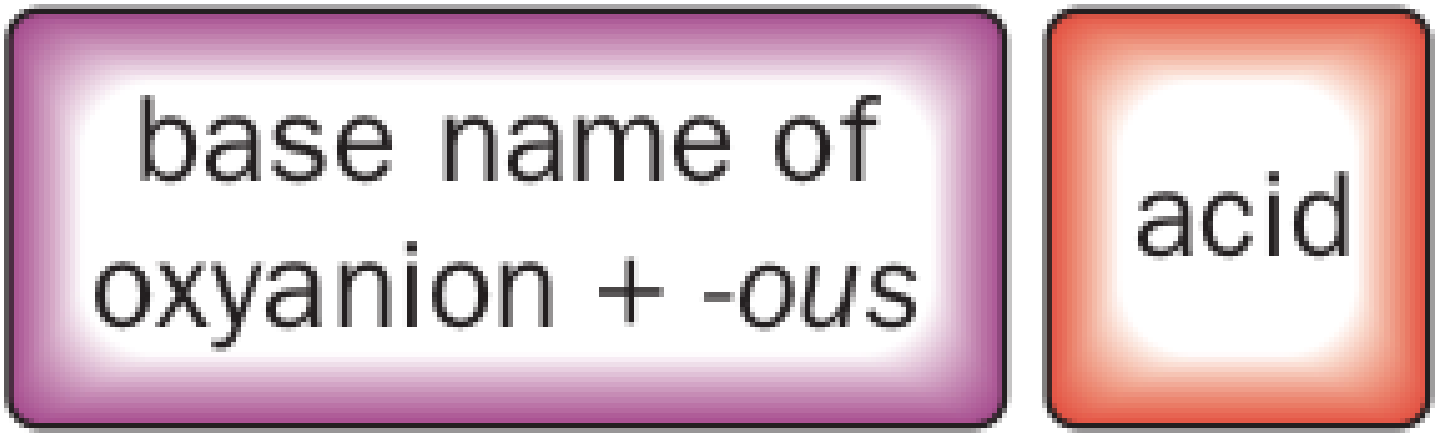
- HF
- HBr
- HI
- H_2S

Oxyacids

- Oxyacids have H^{+1} cation and polyatomic anion.



- The names of acids containing oxyanions ending with *-ite*



- The names of acids containing oxyanions ending with *-ate*

Name the following oxyacids

- H_2SO_4
- HNO_3
- HNO_2

- H_3PO_4

Write the chemical equation for the following acid

- Sulfurous Acid
- Hydrochloric Acid
- Nitrous Acid
- Chromic Acid

Formula Mass

- The mass of an individual molecule or formula unit.
  - Also known as molecular mass or molecular weight.
  - Sum of the masses of the atoms in a single molecule or formula unit.

$$\mathrm{H_2O}$$

$$2\ \mathrm{H}\ \ 2(1.008\ \mathrm{amu}) = 2.016\ \mathrm{amu}$$

$$\mathrm{O}\ \ 15.99\ \mathrm{amu}$$

$$2.016\ \mathrm{amu} + 15.99\ \mathrm{amu} = 18.01\ \mathrm{amu}$$

Calculate the Formula Mass for the following compounds

$Al_2(SO_4)_3$

$C_6H_8O_2$

$Ba(NO_3)_2$

Chapter 6

Chemical Composition

Why Is Knowledge of Composition Important?

- Everything in nature is either chemically or physically combined with other substances.
    - To know the amount of a specific element in a sample, you need to know what fraction of the sample it is.
- Some Applications:
    - The amount of sodium in sodium chloride for diet.
    - The amount of iron in iron ore for steel production.
    - The amount of hydrogen in water for hydrogen fuel.
    - The amount of chlorine in freon to estimate ozone depletion.

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The Mole

Counting Nails by the Pound

- I want to buy a certain number of nails for a project, but the hardware store sells nails by the pound.
    - How do I know how many nails I am buying when I buy a pound of nails?

A hardware store customer buys 2.60 pounds of nails. A dozen nails has a mass of 0.150 pounds. How many nails did the customer buy?

1 dozen nails = 0.150 lbs.
12 nails = 1 dozen nails

A marble company produces three kinds of marbles. What is the average mass of the marbles? The company sells the marbles in bags of sixteen. What is the average mass of a bag of marbles in pounds?

Color	Mass (oz)	Daily Production
Red	2.1	1500
Blue	2.4	1300
Orange	1.9	1400

- If we know the average mass of a particular number of atoms, we can use this information to convert the mass of an element sample to the number of atoms in the sample.
- We can choose a clever quantity to make the units work out convienently
  - The quantity of atoms we will use is 6.022×10^{23} and we call this a **mole**.
  - 1 mole = 6.022×10^{23} things.
  - Like 1 dozen = 12 things or 1 bag = 16 things
  - Avogadro's number.

- The mole is based on careful measurements made on the carbon-12 isotope
    - Mole = Number of things equal to the number of atoms in 12 g of C-12.
    - 1 atom of C-12 weighs exactly 12 amu.
    - 1 mole of C-12 weighs exactly 12 g.

A Silver Ring Contains 1.1×10^{22} Silver Atoms. How Many Moles of Silver Are in the Ring?

Calculate the Number of Atoms in 2.45 Mol of Copper.

Moles and Mass

- The mass of one mole of atoms is called the molar mass.
  - The molar mass ($\frac{\mathrm{g}}{\mathrm{mol}}$) of an element is numerically equal to the element's atomic mass ($\frac{\mathrm{amu}}{\mathrm{atom}}$).
  - The molar mass ($\frac{\mathrm{g}}{\mathrm{mol}}$) of a compound is numerically equal to the compounds formula mass ($\frac{\mathrm{amu}}{\mathrm{molecule}}$).

Calculate the Moles of Sulfur in 57.8 G of Sulfur.

Calculate the Mass of Carbon 2.21×10^{-3} moles of Pencil Lead.

Calculate the Mass of Carbon 2.21×10^{-3} moles of Pencil Lead.

Mass and Atoms

- Using the number of moles allows us to convert between the mass of a sample (a measurable quantity) and the number of atoms or molecules.

How Many Aluminum Atoms Are in a Can Weighing 16.2 g?

What is the mass of 2.94×10^{22} atoms of Cu?

What is the mass of 2.94×10^{22} atoms of Cu?

Calculate the Mass of 1.75 Mol of H_2O.

How Many Moles Are in 50.0 g of PbO_2?

How Many Formula Units Are in 50.0 g of PbO_2?

What Is the Mass of 4.78×10^{24} NO_2 Molecules?

Chemical Formulas as Conversion Factors

1 spider ≡ 8 legs.

1 chair ≡ 4 legs.

1 H_2O molecule ≡ 2 H atoms ≡ 1 O atom.

8 legs ≡ 1 spider 4 legs ≡ 1 chair 2 H atoms ≡ 1 H_2O molecule

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- If we know how many parts are in the whole unit, by counting the number of whole units, we can effectively count the parts.
- Since we count atoms and molecules in mole units, we can find the number of moles of a constituent element if we know the number of moles of the compound.

$$16\,\text{mole H}_2\text{O}\left(\frac{2\,\text{mole H}}{1\,\text{mole H}_2\text{O}}\right) = 32\,\text{mole H}$$

Calculate the Moles of Oxygen in 1.7 Moles of $CaCO_3$.

Find the Mass of Carbon in 55.4 g $C_{10}H_{14}O$.

Find the Mass of Sodium in 6.2 g of NaCl

Percent Composition

- To determine the mass of a component from the mass of a compound you must go through the molar ratios.
- However, the molar ratios and molar masses of the components for a specific compound is always the same.
    - We can be clever to make our math a little simpler

Find the Mass of Sodium in 6.2 g of NaCl using the percent composition.

$$6.2\,\mathrm{g\ NaCl}\left(\frac{1\,\mathrm{mole\ NaCl}}{58.44\,\mathrm{g\ NaCl}}\right)\left(\frac{1\,\mathrm{mole\ Na}}{1\,\mathrm{mole\ NaCl}}\right)\left(\frac{22.99\,\mathrm{g\ Na}}{1\,\mathrm{mole\ Na}}\right)$$

$$= 6.2\,\mathrm{g\ NaCl}\left(\frac{22.9\,\mathrm{g\ Na}}{58.44\,\mathrm{g\ NaCl}}\right)$$

$$= 6.2\,\mathrm{g\ NaCl}\,(0.392) = 2.43\,\mathrm{g\ Na}$$

$$\text{Percent Composition Sodium in Sodium Chloride} = 0.392\ \times 100\,\% = 39.$$

- The percent composition tells you the mass of a constituent element in 100 g of the compound.

Find the Mass Percent of Cl in $C_2Cl_4F_2$.

The experimental mass analysis of the compound.

- The percent composition of a sample can be measured directly through experimentation
  - The percentages may not always total to 100% due to rounding.
- The percent composition data can be used to find the empirical formula of the compound
  - The simplest, whole-number ratio of atoms in a molecule.
  - The molecular formula is a multiple of the empirical formula.

Rules for finding an Empirical Formula from Percent Composition

1. Convert the percentages to grams.
   a. Skip if already grams.
2. Convert grams to moles.
   a. Use molar mass of each element.
3. Write a pseudoformula using moles as subscripts.
4. Divide all by smallest number of moles.
5. Multiply all mole ratios by number to make all whole numbers, if necessary.
   a. If ratio 0.5, multiply all by 2; if ratio 0.33 or 0.67, multiply all by 3, etc.
   b. Skip if already whole numbers after Step 4.

A laboratory analysis of aspirin determined the following mass percent composition. Find the empirical formula.

Element	Percent Composition
C	60.00 %
H	4.48 %
O	35.53 %

A 3.24-g sample of titanium reacts with oxygen to form 5.40 g of the metal oxide. What is the formula of the oxide?

Determine the Empirical Formula of Stannous Fluoride, which Contains 75.7% tin and the Rest Fluorine.

Determine the Empirical Formula of Hematite, which Contains 72.4% Fe and the Rest Oxygen.

Molecular Formulas From Empirical Formulas

- The molecular formula is a multiple of the empirical formula.

- To determine the molecular formula, you need to know the empirical formula and the molar mass of the compound.

Determine the Molecular Formula of Cadinene if it has a Molar Mass of 204 g/mol and an Empirical Formula of C_5H_8.

Benzopyrene has a Molar Mass of 252 g/mol and an Empirical Formula of C_5H_3. What is its Molecular Formula?

Determine the Molecular Formula of Nicotine, which has a Molar Mass of 162 g/mol and is 74.0% C, 8.7% H, and the Rest N.

Chapter 6 Quiz

1. Volume is used as a method of counting atoms. (1 Point)

   True False

2. The numerical value of the mole is defined as being equal to the number of atoms in exactly 1 grams of pure hydrogen-1. (1 Point)

   True False

3. Avogadro's Number is 6.022×10^{23} (1 Point)

   True False

4. The lighter the atom, the more mass in one mole of that atom. (1 Point)

   True False

5. One mole of helium has more atoms in it than one mole of neon. (1 Point)

   True False

6. One mole of zinc contains 65.39 zinc atoms. (1 Point)

   True False

7. One mole of CH_4 has more atoms in it than one mole of Na. (1 Point)

True False

8. A molecular formula gives the specific number of each type of atom in a molecule. (1 Point)

True False

9. An empirical formula gives the smallest whole number ratio of each type of atom in a molecule. (1 Point)

True False

10. How many atoms are in 6.72 moles of scandium? (2 Points)

A) 6.02×10^{23}

B) 1.12×10^{-23}

C) 4.05×10^{24}

D) 1.12×10^{23}

E) none of the above

11. How many moles of Co are in 2.38×10^{25} Co atoms? (2 Points)

A) 39.5

B) 1.43×10^{49}

C) 3.95×10^{24}

D) 79.0

E) none of the above

12. One mole of beryllium has a mass of ____ g. (2 Points)

A) 208.98

B) 10.811

C) 9.012

D) 4.00

E) none of the above

13. What is the mass of 0.605 moles of chlorine gas (Cl_2)? (2 Points)

A) 21.45 g

B) 42.90 g

C) 117.2 g

D) 58.6 g

E) none of the above

14. You have 10.0 g each of Na, C, Pb, Cu and Ne. Which contains the largest number of moles? (2 Points)

A) Na

B) C

C) Pb

D) Cu

E) Ne

15. How many moles of silver are contained in 0.926 kg of silver? (2 Points)

A) 1.54×10^{-21}

B) 8.58

C) 5.58×10^{26}

D) 8.58×10^{-3}

E) none of the above

16. How many atoms are in 15.6 grams of silicon? (2 Points)

A) 2.64×10^{26}

B) 3.34×10^{23}

C) 0.555

D) 438

E) none of the above

17. What is the mass of 3.09×10^{24} molecules of C_2H_2 in grams? (2 Points)

A) 5.07 g

B) 9.91 g

C) 133 g

D) 0.160 g

E) none of the above

18. If 3.011×10^{23} molecules have a mass of 20.04 grams, what is the molar mass of this substance? (2 Points)

A) 40.08 g/mol

B) 10.02 g/mol

C) 20.04 g/mol

D) 6.658×10^{-23} g/mol

E) none of the above

19. Calculate the molar mass of ammonium phosphate, $(NH_4)_3PO_4$. (2 Points)

A) 114.12 g/mol

B) 113.01 g/mol

C) 131.05 g/mol

D) 149.09 g/mol

E) none of the above

21. A 42.7 gram sample of potassium nitrate (KNO_3) contains how many grams of potassium? (2 Points)

A) 39.1 g

B) 16.5 g

C) 21.4 g

D) 8.54 g

E) none of the above

22. Bauxite is an ore that contains the element aluminum. If you obtained 589 grams of aluminum from an ore sample that initially weighed 2.03 kilograms, what is the mass percent of aluminum in this bauxite ore? (2 Points)

A) 71.0 %

B) 3.44 %

C) 29.0 %

D) 96.6 %

E) none of the above

23. What is the mass percent of chlorine in hydrochloric acid? (2 Points)

A) 2.8 %

B) 35.5 %

C) 97.2 %

D) 70.1 %

E) none of the above

Chapter 7

Chemical Reactions

- Reactions involve chemical changes in matter resulting in new substances.
- Chemical bonds are broken and formed to produce new molecules.
  - Molecules can combine to make bigger molecules.
  - Molecules can decompose into smaller molecules or atoms.
  - Atoms can be exchanged between molecules or transferred to another molecule.
  - Atoms can gain or lose electrons, turning them into ions.
    - Or changing the charge on ions that are already there.

Evidence of Chemical Reactions

- Look for evidence of a new substance.
- Permanent Visual clues.

    - Color change.
    - Precipitate formation.
        - Solid that forms when liquid solutions are mixed.
    - Gas bubbles.
    - Large energy changes.
        - Large Temperature Changes
        - Emission of light.
- Other clues.
    - New odor.
    - Whooshing sound from a tube.
    - Permanent new phase.
      Evidence is Not Proof!

In order to be absolutely sure that a chemical reaction has taken place, you need to go down to the molecular level and analyze the structures of the molecules at the beginning and end.

Decide Whether Each of the Following Involve a Chemical Reaction.

- Photosynthesis
- Heating sugar until it turns black

- Heating ice until it becomes a liquid
- Digestion of food
- Dissolving sugar in water
- Burning paper

Chemical Equations

- Short-hand way of describing a reaction.
- Provides lots of information about the reaction.
  - Formulas of reactants and products.
  - Phases of reactants and products.
  - Relative numbers of reactant and product molecules that are required.
  - Can be used to determine masses of reactants used and products that can be made.

Symbols used in Chemical Equations

Phase	Symbol
gas	(*g*)

Phase	Symbol
liquid	(*l*)
solid	(*s*)
aqueous	(*aq*)

Energy	Symbol
heat	Δ
light	$h\nu$
mechanical	$shock$
electrical	$elec$

Symbol	Meaning
$\longrightarrow$	Indicates the direction of the reaction. From the **reactants** on the left to the **products** on the right.
$\rightleftharpoons$	Indicates the reaction is capable of running in both directions (*reversible*)

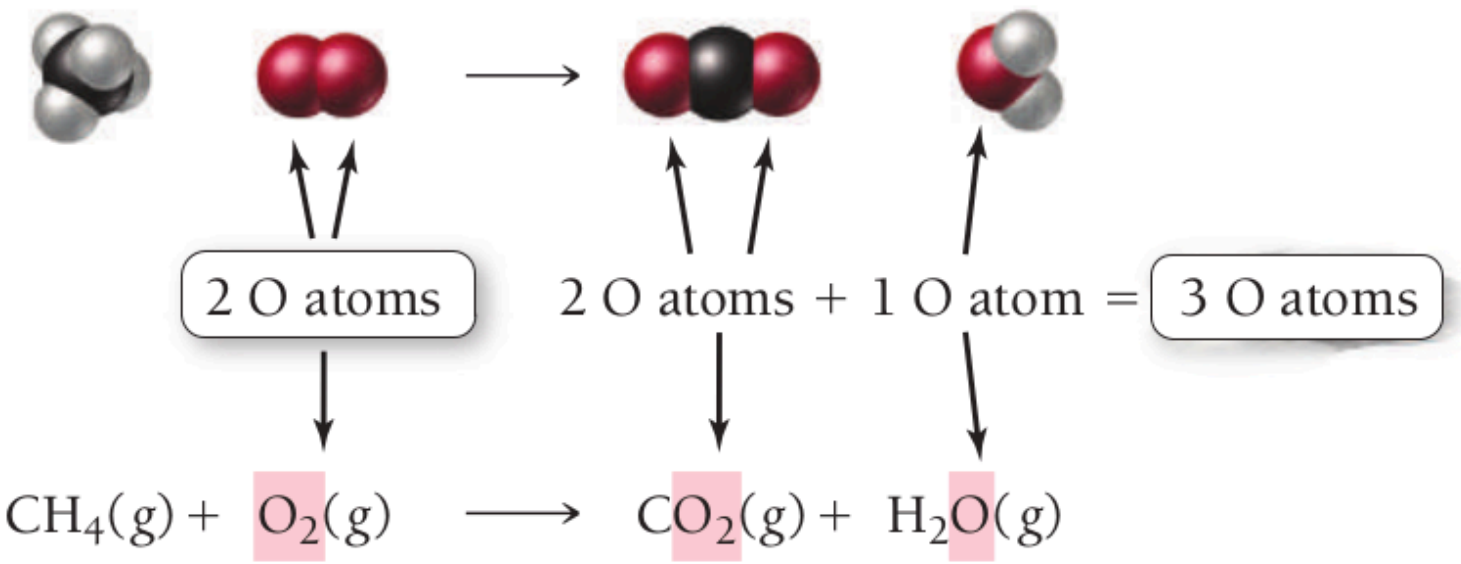
The Combustion of Methane

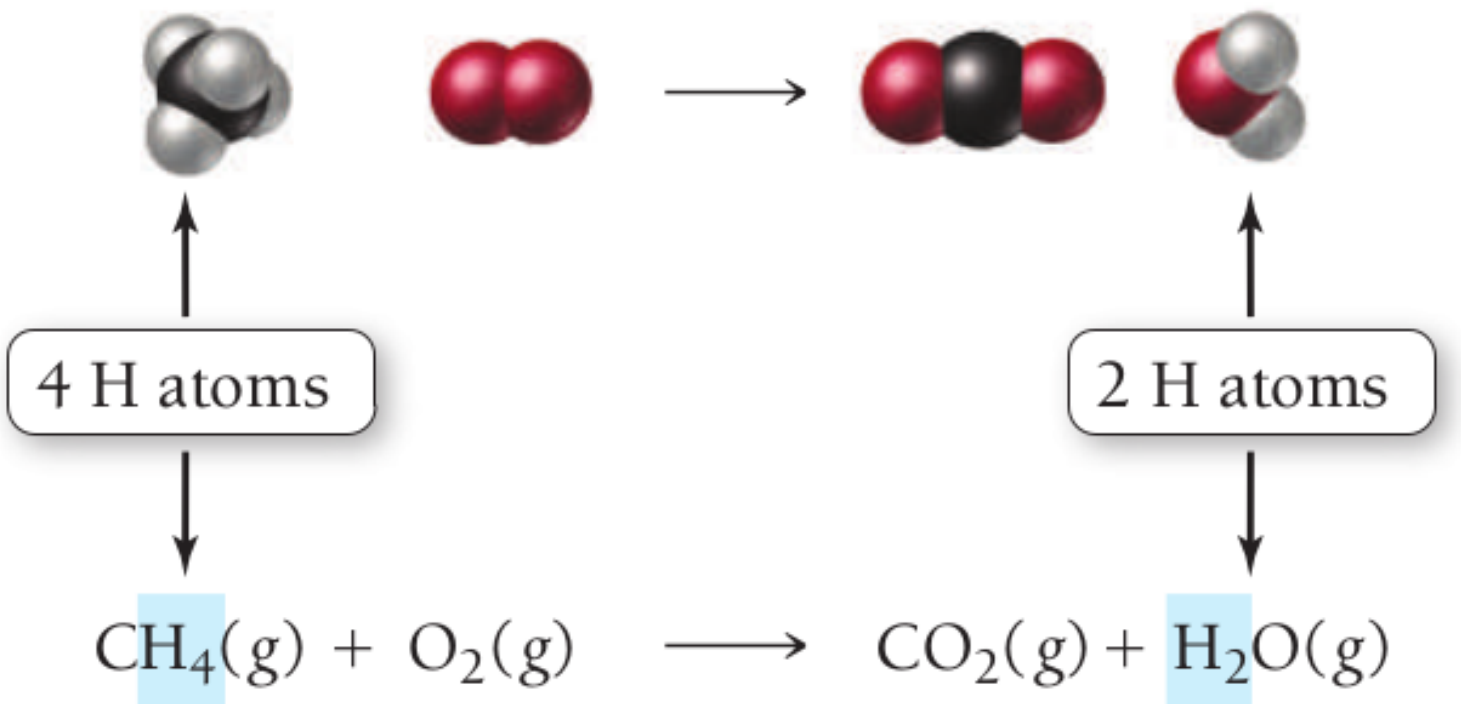
Methane gas burns to produce carbon dioxide gas and gaseous water.

$$\mathrm{CH_4}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) + \mathrm{H_2O}(g)$$

Balancing the Reaction

- As written there is not the same number of atoms on each side
  - This violates the *Law of Conservation of Mass*
  - O and H do not *balance*.





- To correct this we have to adjust the **stoichiometric coefficients**
  - These indicate the number of each molecule participate in the reaction

$$CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(g)$$

- This equation is balanced, meaning that there are equal numbers of atoms of each element on the reactant and product sides.
- To obtain the number of atoms of an element, multiply the subscript by the coefficient.

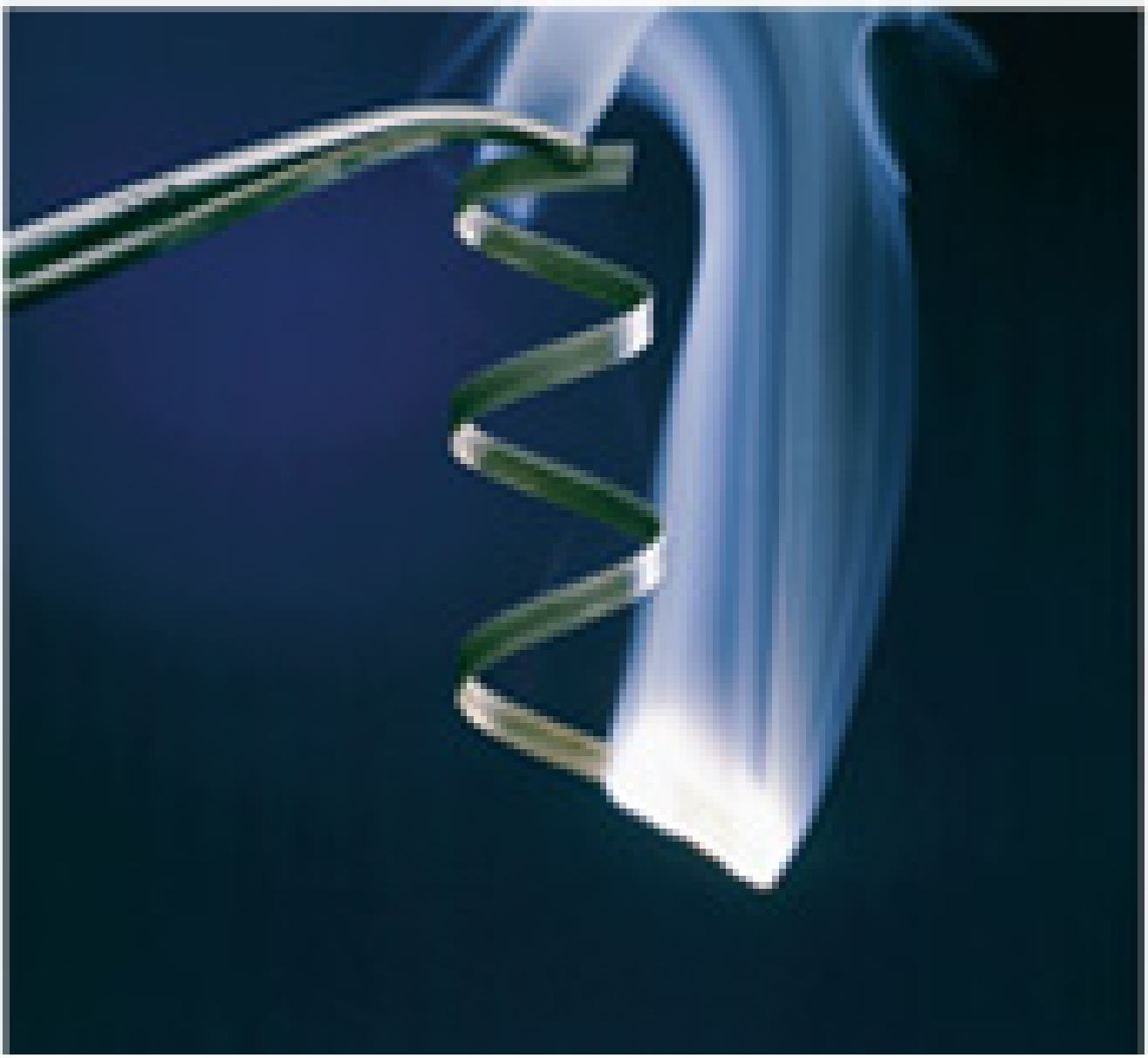
Reactants	Products
1 C atom ($1 \times \underline{C}H_4$)	1 C atom ($1 \times \underline{C}O_2$)
4 H atoms ($1 \times C\underline{H}_4$)	4 H atoms ($2 \times \underline{H}_2O$)
4 O atoms ($2 \times \underline{O}_2$)	4 O atoms ($1 \times C\underline{O}_2 + 2 \times H_2\underline{O}$)

Rules for Writing Balanced Chemical Equations

1. Write a skeletal equation by writing the formula of each reactant and product.

2. Count the number of atoms of each element on each side of the equation.
    - Polyatomic ions may often be counted as if they are one “element”.
3. Pick an element to balance.
    - If an element is found in only one compound on both sides, balance it first.
    - Metals before nonmetals.
    - Leave free elements until last.
4. Find the least common multiple (LCM) of the number of atoms on each side.
5. Multiply each count by a factor to make it equal to the LCM.
6. Use this factor as a coefficient in the equation.
    - If there is already a coefficient there, multiply it by the factor.
    - It must go in front of entire molecules, not between atoms within a molecule.
7. Recount and repeat until balanced.

When magnesium metal burns in air, it produces a white, powdery compound magnesium oxide.



1. Write a skeletal equation

$$\mathrm{Mg}(s) + \mathrm{O_2}(g) \longrightarrow \mathrm{MgO}(s)$$

2. Count the number of atoms on each side.

$$1 \longleftarrow \mathrm{Mg} \longrightarrow 1$$

$$2 \longleftarrow \mathrm{O} \longrightarrow 1$$

3. Pick an element to balance.

Magnesium is already balanced so oxygen is the obvious choice.

4. Find the LCM of both sides

The least common multiple of 2 and 1 is 2.

5. Multiply each side by factor so it equals LCM.
6. Use factors as coefficients in front of the compound containing the element.

$$\mathrm{Mg}(s) + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{MgO}(s)$$

$$1 \longleftarrow \mathrm{Mg} \longrightarrow 1 \times 2$$

$$2 \longleftarrow \mathrm{O} \longrightarrow 1 \times 2$$

7. Recount

$$1 \longleftarrow \mathrm{Mg} \longrightarrow 2$$

$$2 \longleftarrow \mathrm{O} \longrightarrow 2$$

8. and Repeat—attacking an unbalanced element.

$$2\,\mathrm{Mg}(s) + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{MgO}(s)$$

$$2 \ \times 1 \longleftarrow \mathrm{Mg} \longrightarrow 2$$

$$2 \longleftarrow \mathrm{O} \longrightarrow 2$$

9. Recount—Mg not balanced now

$$2 \longleftarrow \mathrm{Mg} \longrightarrow 2$$

$$2 \longleftarrow \mathrm{O} \longrightarrow 2$$

Under appropriate conditions at 1000°C, ammonia gas reacts with oxygen gas to produce gaseous nitrogen monoxide and steam. Write a balanced chemical equation for this reaction.

When aluminum metal reacts with oxygen in the air, it produces a white, powdery compound called aluminum oxide. Write a balanced chemical equation for this reaction.

Acetic acid reacts with the metal aluminum to make aqueous aluminum acetate and gaseous hydrogen. Write a balanced chemical equation for this reaction.

$$\mathrm{Al}(s) + \mathrm{HC_2H_3O_2}(aq) \longrightarrow \mathrm{Al(C_2H_3O_2)_3}(aq) + \mathrm{H_2}(g)$$

Write a balanced chemical for the combustion of ethanol (C_2H_5OH).

Aqueous Solutions

- Many times, the chemicals we are reacting together are dissolved in water.
    - Mixtures of a chemical dissolved in water are called **aqueous solutions**.
- Dissolving the chemicals in water helps them to react together faster.

    - The water separates the chemicals into individual molecules or ions.
    - The separate, free-floating particles come in contact more frequently so the reaction speeds up.

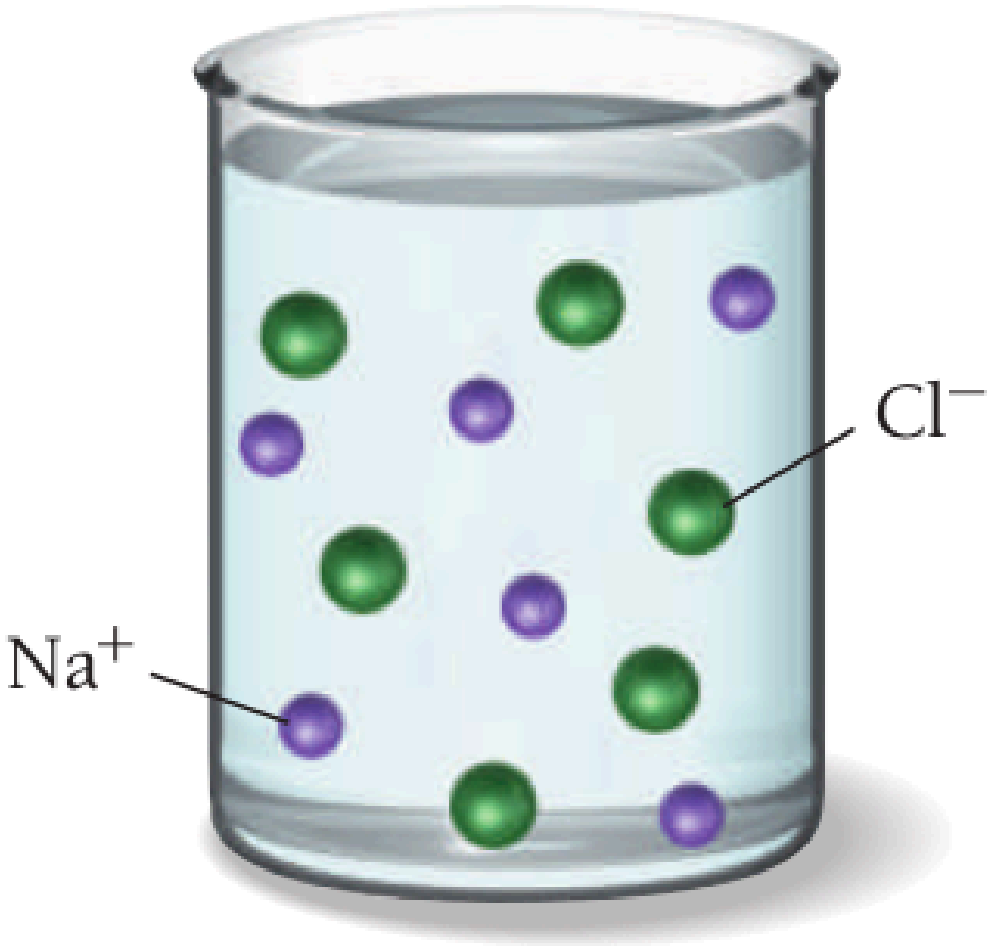
- We can predict whether or not a reaction will happen in aqueous media by considering various *driving forces*
    - “Forces” that drive a reaction:
        - Formation of a solid.
        - Formation of water.
        - Formation of a gas.
        - Transfer of electrons.

Dissociation

- When ionic compounds dissolve in water, the anions and cations are separated from each other.
    - This is called dissociation.
    - Not all ionic compounds will dissolve in water!
    - When compounds containing polyatomic ions dissociate, the polyatomic group stays together as one ion.

Sodium chloride dissociates in water to form sodium cations and chloride anions.

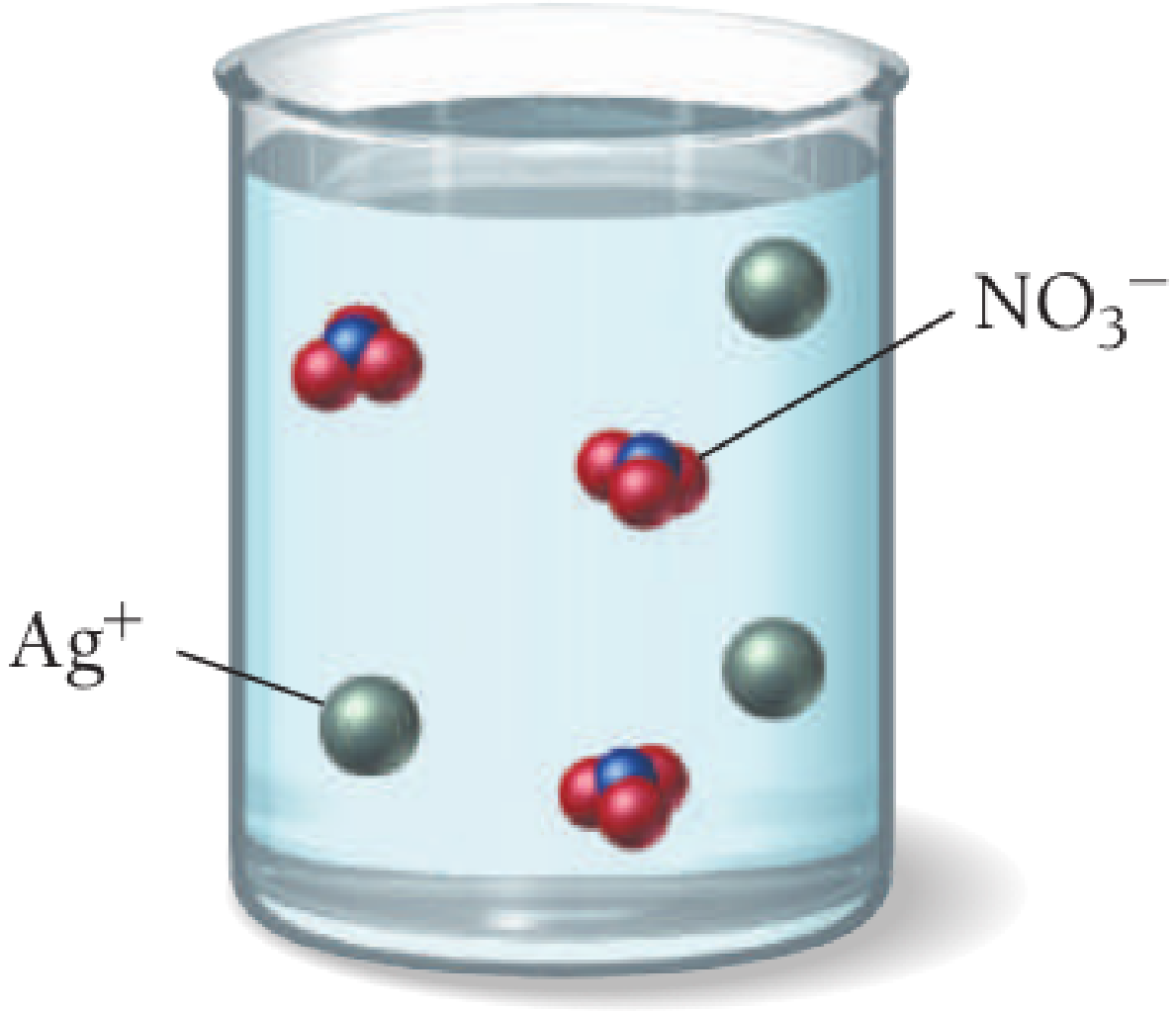
$$\mathrm{NaCl}(aq) \longrightarrow \mathrm{Na}^{+}(aq) + \mathrm{Cl}^{-}(aq)$$



A sodium chloride solution contains independent $\mathbf{Na}^{+}$ and $\mathbf{Cl}^{-}$ ions.

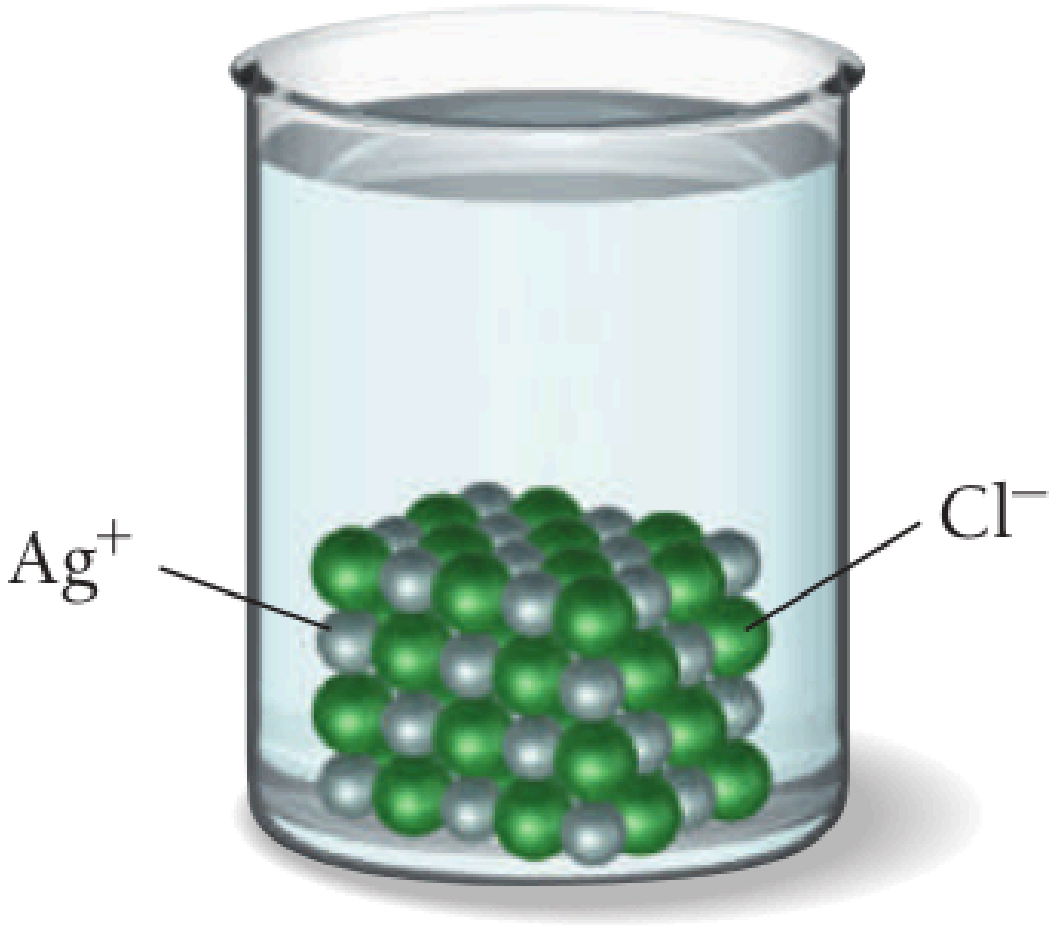
Silver Nitrate dissociates in water to form silver cations and nitrate anions

$$\mathrm{AgNO_3}(aq) \longrightarrow \mathrm{Ag^+}(aq) + \mathrm{NO_3}-(aq)$$



A silver nitrate solution contains independent Ag^+ and NO_3^- ions.

Silver chloride does not dissolve or dissociate. It is *insoluble* in water.



When silver chloride is added to water, it remains as solid AgCl—it does not dissolve into independent ions.

Electrolytes

- **Electrolytes** are substances whose water solution is a conductor of electricity.
  - All electrolytes have ions dissolved in water.
- **Strong electrolyte's** molecules or formula units dissociate completely into ions.
  - Salts, some acids and bases
- **Weak electrolyte's** molecules or formula units dissociate partially into ions.
  - Organic acids, alcohols
- **Nonelectrolyte's** molecules or formula units do not dissociate into ions.
  - Sugars

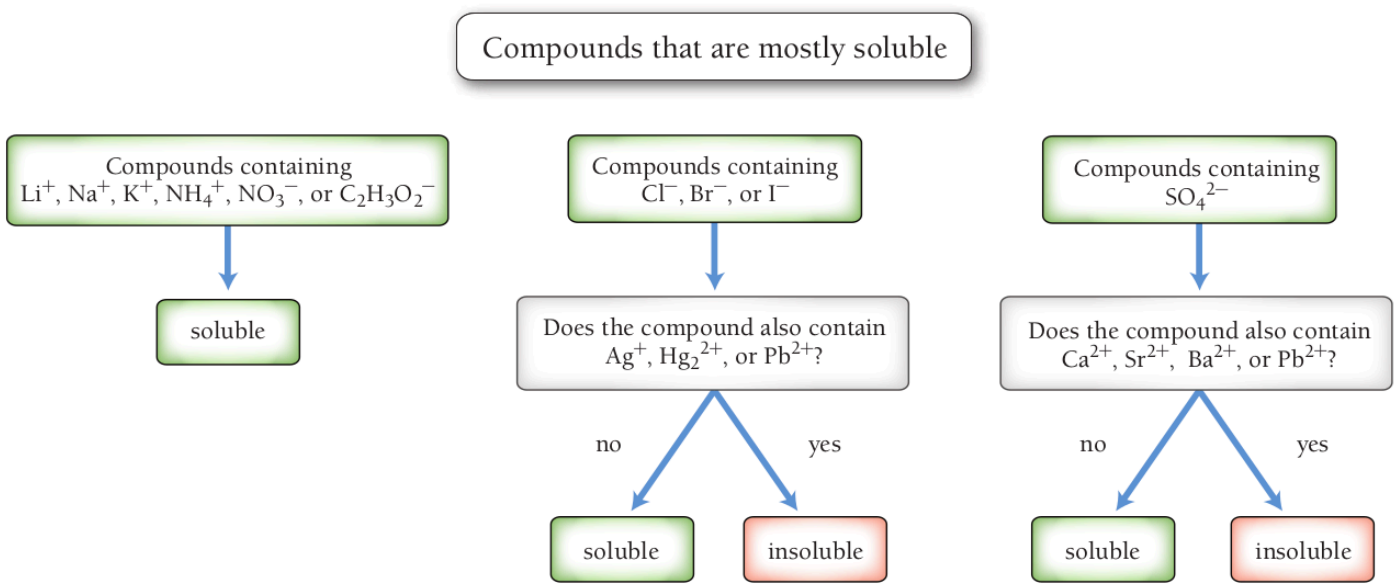
Solubility

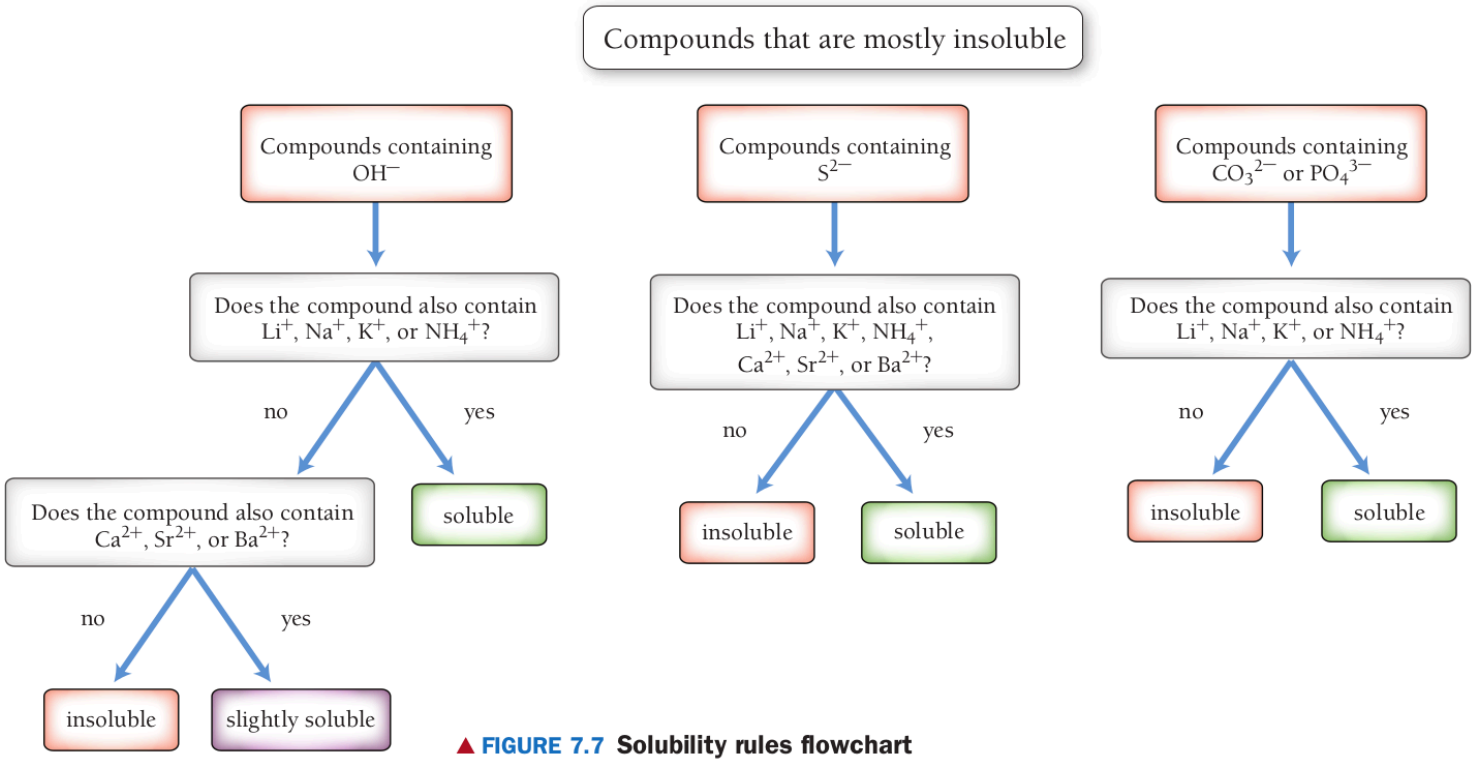
When Will a Salt Dissolve?

- A compound is **soluble** in a liquid if it dissolves in that liquid.
  - NaCl is soluble in water, but AgCl is not.

- A compound is **insoluble** if a significant amount does not dissolve in that liquid.
  - AgCl is insoluble in water.
- Predicting whether a compound will dissolve in water is not easy.
  - A convenient way to do it is to do some experiments to test whether a compound will dissolve in water, then develop some rules based on those experimental results.

Solubility Rules





▲ **FIGURE 7.7 Solubility rules flowchart**

Determine if Each of the Following Is Soluble in Water

- KOH
- AgBr
- $CaCl_2$
- $Pb(NO_3)_2$
- $PbSO_4$

Precipitation Reactions

- Many reactions are done by mixing aqueous solutions of electrolytes together.

- When this is done, often a reaction will take place from the cations and anions in the two solutions that are exchanging.
- If the ion exchange results in forming a compound that is insoluble in water, it will come out of solution as a precipitate.

Process for Predicting the Products of a Precipitation Reaction

1. Write the formula for the reactants
2. Determine what ions each aqueous reactant has.
3. Exchange ions.
    - cations from one reactant with anions from the other.
4. Balance charges of combined ions to get formula of each product.
5. Balance the equation.
6. Determine solubility of each product in water.
    - Use the solubility rules.

- If product is insoluble or slightly soluble, it will precipitate.
- If neither product will precipitate, no reaction.

When an Aqueous Solution of Sodium Carbonate Is Added to an Aqueous Solution of Copper(II) Chloride, a White Solid Forms. Write the formulas of the reactants and Determine the ions present when each reactant dissociates. Predict the Products and Balance the Equation

$$\mathrm{KCl}(aq) + \mathrm{AgNO_3}(aq) \longrightarrow$$

$$\mathrm{Na_2S}(aq) + \mathrm{CaCl_2}(aq) \longrightarrow$$

Ionic Equations

- Equations that describe the chemicals put into the water and the product molecules are called **molecular equations**.

$$2\,\mathrm{KOH}(aq) + \mathrm{Mg(NO_3)_2}(aq) \longrightarrow 2\,\mathrm{KNO_3}(aq) + \mathrm{Mg(OH)_2}(s)$$

- Equations that describe the actual dissolved species are called **complete ionic equations**.
  - Aqueous electrolytes are written as ions.
    - Soluble salts, strong acids, strong bases.
  - Insoluble substances and nonelectrolytes written in molecule form.

- Solids, liquids, and gases are not dissolved, therefore, molecule form.

$$2\,K^+(aq) + 2\,OH^-(aq) + Mg^{2+}(aq) + 2\,NO_3{}^-(aq) \longrightarrow 2\,K^+(aq) + 2\,NO_3{}^-$$

- Notice that both the reactant and the product sides contain $2K^+$ and $2NO_3^-$ ions.
  - These are called **spectator ions**
- Canceling the spectator ions results in the **net ionic equation**.

$$Mg^{2+}(aq) + 2\,OH^-(aq) \longrightarrow Mg(OH)_2(s)$$

Summary

- A molecular equation is a chemical equation showing the complete, neutral formulas for every compound in a reaction.
- A complete ionic equation is a chemical equation showing all of the species as they are actually present in solution.
- A net ionic equation is an equation showing only the species that actually participate in the reaction.

Write the Ionic and Net Ionic Equation.

$$K_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow 2\,KNO_3(aq) + BaSO_4(s)$$

$$\mathrm{Na_2CO_3}(aq) + 2\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{NaCl}(aq) + \mathrm{CO_2}(g) + \mathrm{H_2O}(l)$$

Acid/Base Reactions

Properties of Acids

- Sour taste.
- React with “active” metals, not noble metals.
  - I.e., Al, Zn, Fe, but not Cu, Ag or Au.

$$\mathrm{Zn}(\cdot\,\mathrm{s}\,\cdot) + 2\,\mathrm{HCl}(\cdot\,\mathrm{aq}\,\cdot) \longrightarrow \mathrm{ZnCl_2}(\cdot\,\mathrm{aq}\,\cdot) + \mathrm{H_2}(\cdot\,\mathrm{g}\,\cdot)$$

- Corrosive.

- React with carbonates, producing CO2.
    - Marble, baking soda, chalk, limestone.

$$\mathrm{CaCO_3(\cdot aq\cdot) + 2\,HCl(\cdot aq\cdot) \longrightarrow CaCl_2(\cdot aq\cdot) + CO_2(\cdot g\cdot) + H_2O(\cdot l\cdot}$$

- React with bases to form ionic salts and often water.

Properties of Bases

- Taste bitter
- Caustic
- Feel slippery
- React with acids to form ionic salts.

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{H_2O}(l) + \mathrm{NaCl}(aq)$$

TABLE 7.3 Some Common Acids and Bases

Acid	Formula	Base	Formula
hydrochloric acid	HCl	sodium hydroxide	NaOH
hydrobromic acid	HBr	lithium hydroxide	LiOH
nitric acid	HNO_3	potassium hydroxide	KOH
sulfuric acid	H_2SO_4	calcium hydroxide	$Ca(OH)_2$
perchloric acid	$HClO_4$	barium hydroxide	$Ba(OH)_2$
acetic acid	$HC_2H_3O_2$		

Neutralization Reactions

- The H^{+1} from the acid combines with the OH^{-1} from the base to make water.
  The cation from the base combines with the anion from the acid to make the salt.

$$\text{acid} + \text{base} \longrightarrow \text{salt} + \text{water}$$

$$2\,\mathrm{HNO_3}(aq) + \mathrm{Ca(OH)_2}(aq) \longrightarrow \mathrm{Ca(NO_3)_2}(aq) + 2\,\mathrm{H_2O}(l)$$

- The net ionic equation for an acid-base reaction is often

$$\mathrm{H^{+1}}(aq) + \mathrm{OH^{-1}}(aq) \longrightarrow \mathrm{H_2O}(l)$$

- As long as the salt that forms is soluble in water.

Process for Predicting the Products of an Acid–Base Reaction

1. Determine what ions each aqueous reactant has.
2. Exchange ions.
   - cation from one reactant with anion from the other.
   - H^+ combines with OH^- to make water.
3. Balance charges of combined ions to get formula of the salt.
4. Balance the equation.
5. Determine solubility of the salt.

- Use the solubility rules.
- If the salt is insoluble or slightly soluble, it will precipitate.

Write the Molecular, Ionic, and Net-Ionic Equation for the Reaction of Aqueous Nitric Acid with Aqueous Calcium Hydroxide.

Complete and Balance These Acid–Base Reactions.

$$\mathrm{NH_4OH}(aq) + \mathrm{H_2SO_4}(aq) \longrightarrow$$

$$\mathrm{Al(OH)_3}(aq) + \mathrm{H_2SO_3}(aq) \longrightarrow$$

$$\mathrm{Ba(OH)_2}(aq) + \mathrm{HF}(aq) \longrightarrow$$

Gas Evolution Reactions

- Reactions in which the driving force is the production of a material that escapes as a gas are called gas evolution reactions.
- Some reactions form a gas directly from the ion exchange.

$$\mathrm{K_2S}(aq) + \mathrm{H_2SO_4}(aq) \longrightarrow \mathrm{K_2SO_4}(aq) + \mathrm{H_2S}(g)$$

- Other reactions form a gas by the decomposition of one of the ion exchange products into a gas and water.

$$\mathrm{K_2SO_3}(aq) + \mathrm{H_2SO_4}(aq) \longrightarrow \mathrm{K_2SO_4}(aq) + \mathrm{H_2SO_3}(aq) \longrightarrow \mathrm{K_2SO_4}(aq) +$$

Compounds that Undergo Gas Evolving Reactions

- Metal Sulfides, M_nS or MHS

$$\mathrm{K_2S}(aq) + 2\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{KCl}(aq) + \mathrm{H_2S}(g)$$

- Carbonates, M_nCO_3 or $MHCO_3$

$$\mathrm{K_2CO_3}(aq) + 2\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{KCl}(aq) + \mathrm{CO_2}(g) + \mathrm{H_2O}(l)$$

- Sulfites, M_nSO_3 or $MHSO_3$

$$\mathrm{K_2SO_3}(aq) + 2\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{KCl}(aq) + \mathrm{SO_2}(g) + \mathrm{H_2O}(l)$$

- Ammonium Salts, $(NH_4)_nA$

$$\mathrm{KOH}(aq) + \mathrm{NH_4Cl}(aq) \longrightarrow \mathrm{KCl}(aq) + \mathrm{NH_3}(aq) + \mathrm{H_2O}(l)$$

Process for Predicting the Products of a Gas-Evolving Reaction

1. Determine what ions each aqueous reactant has.
2. Exchange ions.
- cation from one reactant with an ion from the other.
3. Balance charges of combined ions to get formula of each product.
4. Check to see if either product is H2S.
5. Check to see if either product decomposes. If so, rewrite as $H_2O(l)$ and a gas.
6. Balance the equation.
7. Determine solubility of other product in water.

When an Aqueous Solution of Sodium Sulfite Is Added to an Aqueous Solution of Nitric Acid, a Gas Evolves.

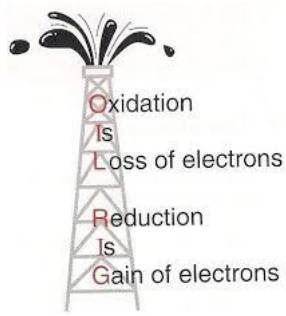
Complete the Following Reactions.

$$\mathrm{PbS}(s) + \mathrm{H_2SO_4}(aq) \longrightarrow$$

$$\mathrm{HNO_3}(aq) + \mathrm{NaHCO_3}(aq) \longrightarrow$$

Oxidation Reduction Reactions

- Redox reactions occur when one chemical species loses one or more electrons to another.
  - The species that loses electrons in the reaction is **oxidized**.
  - The species that gains electrons in the reaction is **reduced**.
  - You cannot have one without the other.



- In combustion, the O atoms in O_2 are reduced, and the non-O atoms in the other material are oxidized.
- Metals react with nonmetals to form ionic compounds.
  - The metal loses electrons and becomes a cation (oxidation).
  - The nonmetal gains electrons and becomes an anion (rduction).
  - The net result electrons are transferred from the metal to the nonmetal.

Example Metal with Nonmetal

In the reaction

$$\mathrm{Mg} + \mathrm{Cl_2} \longrightarrow \mathrm{MgCl_2}$$

The magnesium atoms are oxidized.

$$Mg- > Mg^{2+} + 2e-$$

The chlorine atoms are reduced.

$$\mathrm{Cl} + 1\,\mathrm{e}^- \longrightarrow \mathrm{Cl}^-$$

Example Combustion Reactions

- Reactions in which $O_2(g)$ is a reactant are called combustion reactions.
- Combustion reactions release lots of energy. They are exothermic.
- Combustion reactions are a subclass of oxidation–reduction reactions.
  -
  In the following reaction

$$2\,\mathrm{Mg}(s) + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{MgO}(s)$$

The magnesium atoms are oxidized.

$$\mathrm{Mg} \longrightarrow \mathrm{Mg}^{2+} + 2\,\mathrm{e}^-$$

The oxygen atoms are reduced.

$$\mathrm{O} + 2\,\mathrm{e}^- \longrightarrow \mathrm{O}^{2-}$$

Even though the following reaction does not involve ion formation, electrons are still transferred.

$$\mathrm{CH_4}(g) + 2\,\mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) + 2\,\mathrm{H_2O}(g)$$

The carbon atoms are oxidized.

$$\mathrm{C}^{4-} \longrightarrow \mathrm{C}^{4+} + 8\,\mathrm{e}^-$$

- These are *not charges*, they are called **oxidation numbers**
    - They help us see the electron transfer.

The oxygen atoms are reduced.

$$\mathrm{O} + 2\,\mathrm{e}^- \longrightarrow \mathrm{O}^{2-}$$

Recognizing Redox Reactions

- Any reaction where O_2 is a reactant or a product is a redox reaction.
- Any reaction between a metal and a nonmetal is redox.
- Any reaction where electrons are transferred is redox.
- When a free element gets combined into a compound, it will be either oxidized or reduced.

$$\mathrm{N_2}(g) + \mathrm{H_2}(g) \longrightarrow \mathrm{NH_3}(g)$$

- When a metal cation changes its charge
    - Oxidized if its charge increases or reduced if its charge decreases.

$$\mathrm{CuCl}(aq) + \mathrm{FeCl_3}(aq) \longrightarrow \mathrm{FeCl_2}(aq) + \mathrm{CuCl_2}(aq)$$

Decide Whether Each of the Following Reactions Is a Redox Reaction.

$$2\,\mathrm{Al}(s) + 3\,\mathrm{Br_2}(l) \longrightarrow 2\,\mathrm{AlBr_3}(s)$$

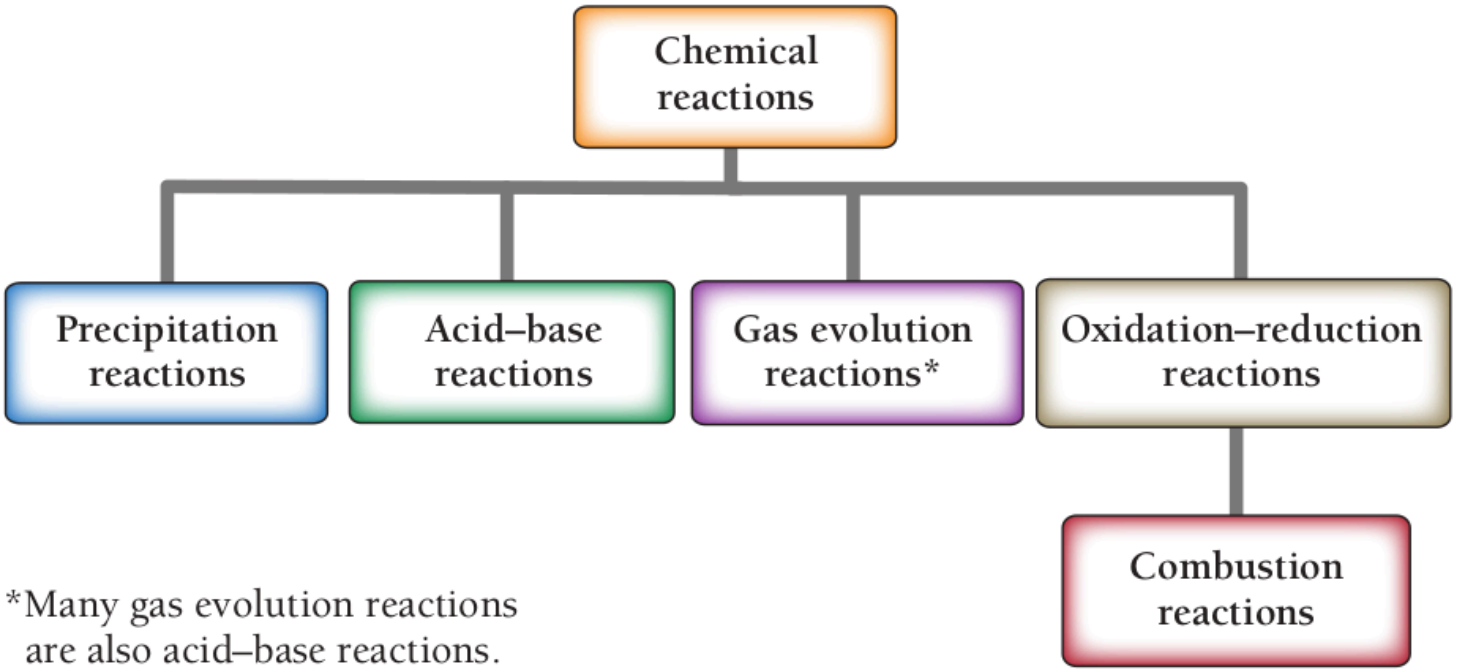
$$\mathrm{CaSO_3}(s) + 2\,\mathrm{HCl}(aq) \longrightarrow \mathrm{CaCl_2}(aq) + \mathrm{SO_2}(g) + \mathrm{H_2O}(l)$$

$$\mathrm{Fe_2O_3}(s) + \mathrm{C}(s) \longrightarrow \mathrm{Fe}(s) + 3\,\mathrm{CO}(g)$$

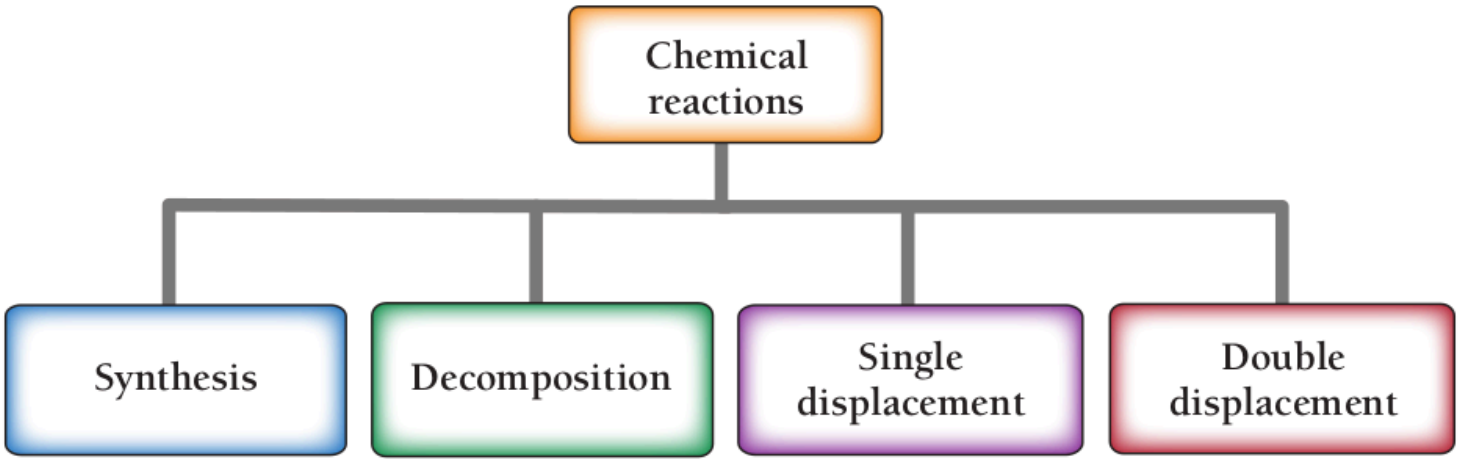
$$\mathrm{SO_2}(g) + \mathrm{O_2}(g) + \mathrm{H_2O}(l) \longrightarrow \mathrm{H_2SO_4}(aq)$$

Classifying Reactions

One way is based on the process that happens.



Another scheme classifies reactions by what the atoms do.



Type of Reaction	Generic Equation
synthesis or combination	$A + B \longrightarrow AB$
decomposition	$AB \longrightarrow A + B$
single-displacement	$A + BC \longrightarrow AC + B$
double-displacement	$AB + CD \longrightarrow AD + CB$

Synthesis Reactions

- Also known as composition or combination reactions.
  - Two (or more) reactants combine together to make one product.
  - Simpler substances combining together.

$$2\,\mathrm{CO} + \mathrm{O_2} \longrightarrow 2\,\mathrm{CO_2}$$

$$2\,\mathrm{Mg} + \mathrm{O_2} \longrightarrow 2\,\mathrm{MgO}$$

$$\mathrm{HgI_2} + 2\,\mathrm{KI} \longrightarrow \mathrm{K_2HgI_4}$$

Decomposition Reactions

- A large molecule is broken apart into smaller molecules or its elements.
  - Caused by addition of energy into the molecule.
- One reactant breaks into two or more products.

$$2\,\mathrm{FeCl_3}(s) \xrightarrow{elec} 2\,\mathrm{FeCl_2}(l) + \mathrm{Cl_2}(g)$$

$$2\,\mathrm{HgO}(s) \xrightarrow{\Delta} 2\,\mathrm{Hg}(l) + \mathrm{O_2}(g)$$

$$2\,\mathrm{O_3} \xrightarrow{h\nu} 3\,\mathrm{O_2}(g)$$

Single Displacement Reactions

- Reactions that involve one atom displacing another and replacing it in a compound.

Examples

$$\mathrm{Zn}(s) + 2\,\mathrm{HCl}(aq) \longrightarrow \mathrm{ZnCl_2}(aq) + \mathrm{H_2}(g)$$

the atom Zn displaces H from the compound.

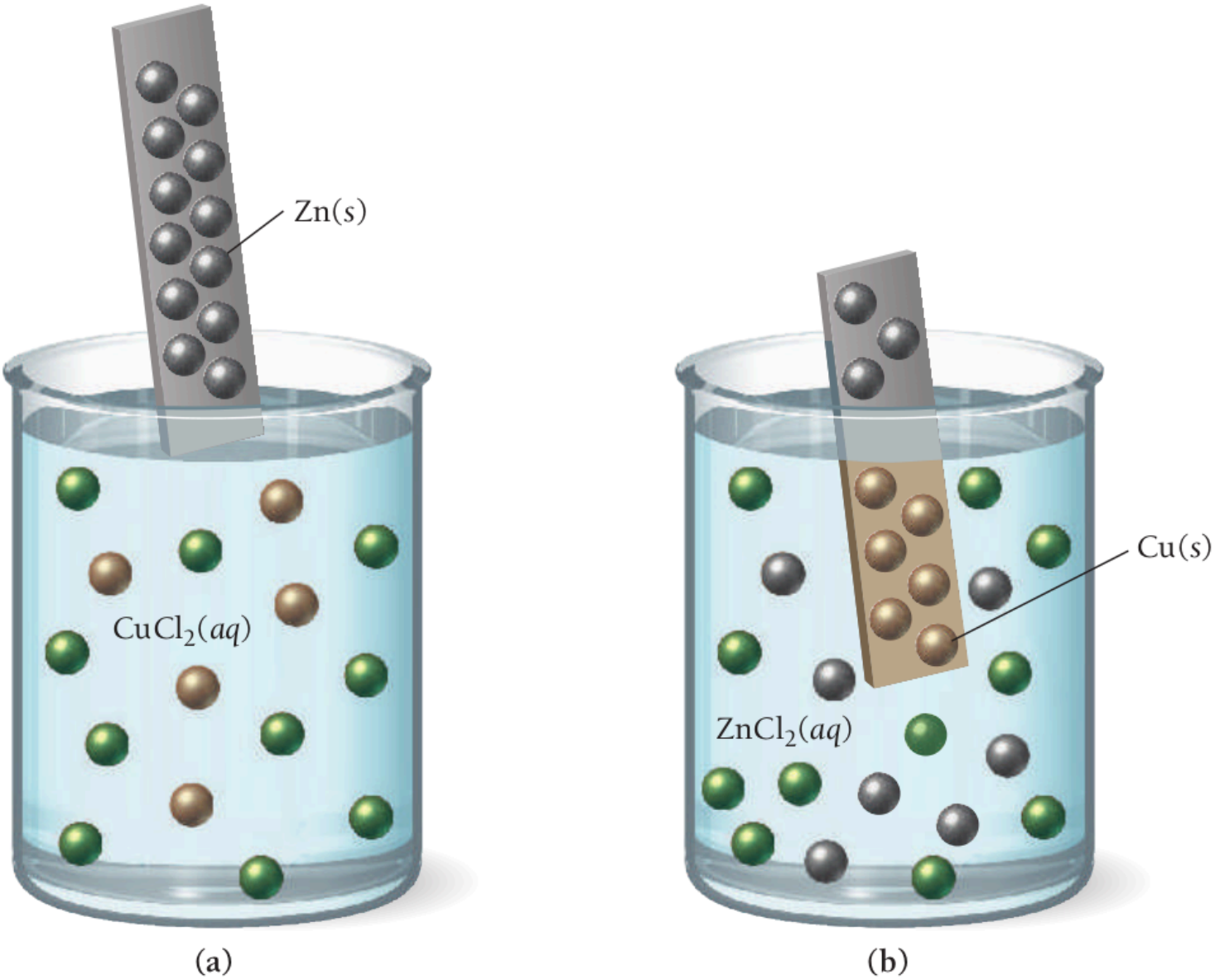
$$\mathrm{Fe_2O_3}(s) + \mathrm{Al}(s) \longrightarrow \mathrm{Fe}(s) + \mathrm{Al_2O_3}(s)$$

the Al atom displaces the Fe atoms

$$2\,\mathrm{Na}(s) + 2\,\mathrm{H_2O}(aq) \longrightarrow 2\,\mathrm{NaOH}(aq) + \mathrm{H_2}(g)$$

Na atoms displaces the H atoms

$$\mathrm{Zn}(s) + \mathrm{CuCl_2}(aq) \longrightarrow \mathrm{Cu}(s) + \mathrm{ZnCl_2}(aq)$$



Double Displacement Reactions

- Two ionic compounds exchange ions.
- May be followed by decomposition of one of the products to make a gas.
- Precipitation, acid–base, and gas evolving reactions are also double displacement reactions.

Examples of Double Displacement

$$\mathrm{CaCl_2}(aq) + \mathrm{Na_2CO_3}(aq) \longrightarrow \mathrm{CaCO_3}(s) + 2\,\mathrm{NaCl}(aq)$$

$$\mathrm{Ba(OH)_2}(s) + 2\,\mathrm{HNO_3}(aq) \longrightarrow \mathrm{Ba(NO_3)_2}(aq) + 2\,\mathrm{H_2O}(l)$$

$$\mathrm{Li_2CO_3}(aq) + 2\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{LiCl}(aq) + \mathrm{CO_2}(g) + \mathrm{H_2O}(l)$$

Classify the Following Reactions as Synthesis, Decomposition, Single Displacement, or Double Displacement.

$$3\,\mathrm{Mg}(s) + 2\,\mathrm{FeCl_3}(aq) \longrightarrow 3\,\mathrm{MgCl_2}(aq) + 2\,\mathrm{Fe}(s)$$

$$\mathrm{CO_2}(g) + \mathrm{H_2O}(l) \longrightarrow \mathrm{H_2CO_3}(aq)$$

$$3\,\mathrm{KOH}(aq) + \mathrm{H_3PO_4}(aq) \longrightarrow \mathrm{K_3PO_4}(aq) + 3\,\mathrm{H_2O}(l)$$

$$\mathrm{CaCO_3}(s) \xrightarrow{\Delta} \mathrm{CaO}(s) + \mathrm{CO_2}(g)$$

Reference Material

American Linear Units	
12 inches (in)	1 foot (ft)
3 feet	1 yard (yd)
36 inches	1 yard
63,360 inches	1 mile (mi)
5,280 feet	1 mile
1,760 yards	1 mile

Weight and Mass	
1 Ton (T)	2,000 pounds
1 pound (lb)	16 ounces (oz)
1 Ton	32,000 ounces
1 metric ton (t)	1000 kg
60 grains	1 dram

Converting American Units	
Larger unit → smaller unit	*Multiply*
smaller unit → Larger unit	*Divide*

American to Metric Units	
1 inch	2.540 centimeters
1 foot	0.305 meters
1 yard	0.914 meters
1 mile	1.609 kilometers
1 gallon	3.78 Liters
1 quart	0.95 Liter
1 pound	0.45 kilogram
1 ounce	28.35 grams
1 fluid ounce	29.57 mL
1 grain	60 milligrams (mg)
1 teaspoon (tsp)	5 mL
1 fluid dram	4 mL
1 tablespoon (tbsp)	15 mL
1 pint (pt)	500 mL (approx)
1 quart (qt)	1000 mL (approx)
1 pound (lb)	453.6 g

American Capacity	
8 fluid ounces (fl oz)	1 cup
16 fluid ounces	2 cups
2 cups	1 pint (pt)
16 fluid ounces	1 pint
2 pints	1 quart (qt)
4 quarts	1 gallon
8 pints	1 gallon
32 fluid ounces	1 quart
8 fluid dram	1 fluid ounce
3 teaspoon (tsp)	1 tablespoon (tbsp)
6 teaspoon	1 fluid ounce
2 tablespoon	1 fluid ounce
1 drop (gtt)	1 minim
60 drop	1 fluid dram
60 drop	1 teaspoon
60 minims	1 fluid dram

Metric Units

mega (M)	*	*	kilo (k)	hector (h)	deka (da)	unit (m, g, L)	deci (d)	centi (c)	milli (m)	*	*	micro (mc) (u)

When going from larger unit to smaller unit move decimal to the right →

When going from smaller unit to larger unit move decimal to the left ←

Time	
1 day	24 hours
1 hour (hr)	60 minutes (min)
1 minute	60 seconds (sec)
1 year (yr)	365.25 days
1 week	7 days
1 year	12 months (mon)
1440 minutes	1 day
3600 seconds	1 hour

Stones	
1 carat (karat)	200 mg

Metric to American Units	
1 km	0.621 miles
1 meter	1.094 yards
1 meter	3.281 feet
1 meter	39.370 inches
1 cm	0.3937 inch
1 Liter	0.26 gallon
1 Liter	1.06 quarts
1 kg	2.20 lbs
1 gram	0.035 oz
1 gram	15 grains
1 milliliter (mL)	15 minims

Temperature Formulas

$C = \frac{(F - 32)}{1.8}$	$F = 1.8 \cdot C + 32$

Medical Application (Micrograms)	
1,000,000 micrograms (mcg)	1 gram
1,000,000 micrograms	1,000 mg

1 mL = 1 cc = 1 cm^3

1 gram = 1 cm^3

Nursing students 1fl oz = 30 mL

Nursing students 1 in. = 2.5 cm

Atomic Masses of the Elements

Based on carbon-12. A number in parentheses is the atomic mass of the most stable isotope of a radioactive element.

Name	Symbol	Atomic Number	Atomic Weight
Actinium	Ac	89	(227)
Aluminum	Al	13	26.98
Americium	Am	95	(243)
Antimony	Sb	51	121.75
Argon	Ar	18	39.95
Arsenic	As	33	74.92
Astatine	At	85	(210)
Barium	Ba	56	137.33
Berkelium	Bk	97	(247)
Beryllium	Be	4	9.01
Bismuth	Bi	83	208.98
Bohrium	Bh	107	(262)
Boron	B	5	10.81
Bromine	Br	35	79.90
Cadmium	Cd	48	112.41
Calcium	Ca	20	40.08
Californium	Cf	98	(251)
Carbon	C	6	12.01
Cerium	Ce	58	140.12
Cesium	Cs	55	132.91
Chlorine	Cl	17	35.45
Chromium	Cr	24	52.00
Cobalt	Co	27	58.93
Copernicium	Cn	112	(285)
Copper	Cu	29	63.55
Curium	Cm	96	(247)
Darmstadtium	Ds	110	(271)
Dubnium	Db	105	(262)
Dysprosium	Dy	66	162.50
Einsteinium	Es	99	(252)
Erbium	Er	68	167.26
Europium	Eu	63	151.97
Fermium	Fm	100	(257)
Flerovium	Fl	114	(289)
Fluorine	F	9	19.00
Francium	Fr	87	(223)
Gadolinium	Gd	64	157.25
Gallium	Ga	31	69.72
Germanium	Ge	32	72.63
Gold	Au	79	196.97
Hafnium	Hf	72	178.49
Hassium	Hs	108	(265)
Helium	He	2	4.00
Holmium	Ho	67	164.93
Hydrogen	H	1	1.01
Indium	In	49	114.82
Iodine	I	53	126.90
Iridium	Ir	77	192.22
Iron	Fe	26	55.85
Krypton	Kr	36	83.80
Lanthanum	La	57	138.91
Lawrencium	Lr	103	(260)
Lead	Pb	82	207.2
Lithium	Li	3	6.94
Livermorium	Lv	116	(293)
Lutetium	Lu	71	174.97
Magnesium	Mg	12	24.31
Manganese	Mn	25	54.94
Meitnerium	Mt	109	(266)
Mendelevium	Md	101	(258)
Mercury	Hg	80	200.59
Molybdenum	Mo	42	95.95
Moscovium	Mc	115	(289)
Neodymium	Nd	60	144.24
Neon	Ne	10	20.18
Neptunium	Np	93	(237)
Nickel	Ni	28	58.69
Nihonium	Nh	113	(284)
Niobium	Nb	41	92.91
Nitrogen	N	7	14.01
Nobelium	No	102	(259)
Oganesson	Og	118	(294)
Osmium	Os	76	190.23
Oxygen	O	8	16.00
Palladium	Pd	46	106.42
Phosphorus	P	15	30.97
Platinum	Pt	78	195.08
Plutonium	Pu	94	(244)
Polonium	Po	84	(209)
Potassium	K	19	39.10
Praseodymium	Pr	59	140.91
Promethium	Pm	61	(147)
Protactinium	Pa	91	(231)
Radium	Ra	88	(226)
Radon	Rn	86	(222)
Rhenium	Re	75	186.21
Rhodium	Rh	45	102.91
Roentgenium	Rg	111	(280)
Rubidium	Rb	37	85.47
Ruthenium	Ru	44	101.07
Rutherfordium	Rf	104	(261)
Samarium	Sm	62	150.36
Scandium	Sc	21	44.96
Seaborgium	Sg	106	(263)
Selenium	Se	34	78.97
Silicon	Si	14	28.09
Silver	Ag	47	107.87
Sodium	Na	11	22.99
Strontium	Sr	38	87.62
Sulfur	S	16	32.06
Tantalum	Ta	73	180.95
Technetium	Tc	43	(99)
Tellurium	Te	52	127.60
Tennessine	Ts	117	(294)
Terbium	Tb	65	158.93
Thallium	Tl	81	204.38
Thorium	Th	90	(232)
Thulium	Tm	69	168.93
Tin	Sn	50	118.71
Titanium	Ti	22	47.88
Tungsten	W	74	183.85
Uranium	U	92	(238)
Vanadium	V	23	50.94
Xenon	Xe	54	131.29
Ytterbium	Yb	70	173.04
Yttrium	Y	39	88.91
Zinc	Zn	30	65.39
Zirconium	Zr	40	91.22

Periodic Table of the Elements

	GROUP 1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B
PERIOD 1	1 **H** 1.01 hydrogen								
2	3 **Li** 6.94 lithium	4 **Be** 9.01 beryllium							
3	11 **Na** 22.99 sodium	12 **Mg** 24.31 magnesium							
4	19 **K** 39.10 potassium	20 **Ca** 40.08 calcium	21 **Sc** 44.96 scandium	22 **Ti** 47.88 titanium	23 **V** 50.94 vanadium	24 **Cr** 52.00 chromium	25 **Mn** 54.94 manganese	26 **Fe** 55.85 iron	27 **Co** 58.93 cobalt
5	37 **Rb** 85.47 rubidium	38 **Sr** 87.62 strontium	39 **Y** 88.91 yttrium	40 **Zr** 91.22 zirconium	41 **Nb** 92.91 niobium	42 **Mo** 95.95 molybdenum	43 **Tc** (99) technetium	44 **Ru** 101.07 ruthenium	45 **Rh** 102.91 rhodium
6	55 **Cs** 132.91 cesium	56 **Ba** 137.33 barium	57 **La** 138.91 lanthanum	72 **Hf** 178.49 hafnium	73 **Ta** 180.95 tantalum	74 **W** 183.85 tungsten	75 **Re** 186.21 rhenium	76 **Os** 190.23 osmium	77 **Ir** 192.22 iridium
7	87 **Fr** (223) francium	88 **Ra** (226) radium	89 **Ac** (227) actinium	104 **Rf** (261) rutherfordium	105 **Db** (262) dubnium	106 **Sg** (263) seaborgium	107 **Bh** (262) bohrium	108 **Hs** (265) hassium	109 **Mt** (266) meitnerium

1 — Atomic number
H — Element symbol
1.01 — Atomic mass*
hydrogen — Element name

Lanthanide series	58 **Ce** 140.12 cerium	59 **Pr** 140.91 praseodymium	60 **Nd** 144.24 neodymium	61 **Pm** (147) promethium	62 **Sm** 150.36 samarium	63 **Eu** 151.97 europium
Actinide series	90 **Th** (232) thorium	91 **Pa** (231) protactinium	92 **U** (238) uranium	93 **Np** (237) neptunium	94 **Pu** (244) plutonium	95 **Am** (243) americium

Metals

Metalloids

Nonmetals

								18 8A
			13 3A	14 4A	15 5A	16 6A	17 7A	2 **He** 4.00 helium
			5 **B** 10.81 boron	6 **C** 12.01 carbon	7 **N** 14.01 nitrogen	8 **O** 16.00 oxygen	9 **F** 19.00 fluorine	10 **Ne** 20.18 neon
10 8B	11 1B	12 2B	13 **Al** 26.98 aluminum	14 **Si** 28.09 silicon	15 **P** 30.97 phosphorus	16 **S** 32.06 sulfur	17 **Cl** 35.45 chlorine	18 **Ar** 39.95 argon
28 **Ni** 58.69 nickel	29 **Cu** 63.55 copper	30 **Zn** 65.39 zinc	31 **Ga** 69.72 gallium	32 **Ge** 72.63 germanium	33 **As** 74.92 arsenic	34 **Se** 78.97 selenium	35 **Br** 79.90 bromine	36 **Kr** 83.80 krypton
46 **Pd** 106.42 palladium	47 **Ag** 107.87 silver	48 **Cd** 112.41 cadmium	49 **In** 114.82 indium	50 **Sn** 118.71 tin	51 **Sb** 121.75 antimony	52 **Te** 127.60 tellurium	53 **I** 126.90 iodine	54 **Xe** 131.29 xenon
78 **Pt** 195.08 platinum	79 **Au** 196.97 gold	80 **Hg** 200.59 mercury	81 **Tl** 204.38 thallium	82 **Pb** 207.2 lead	83 **Bi** 208.98 bismuth	84 **Po** (209) polonium	85 **At** (210) astatine	86 **Rn** (222) radon
110 **Ds** (281) darmstadtium	111 **Rg** (280) roentgenium	112 **Cn** (285) copernicium	113 **Nh** (284) nihonium	114 **Fl** (289) flerovium	115 **Mc** (289) moscovium	116 **Lv** (293) livermorium	117 **Ts** (294) tennessine	118 **Og** (294) oganesson

64 **Gd** 157.25 gadolinium	65 **Tb** 158.93 terbium	66 **Dy** 162.50 dysprosium	67 **Ho** 164.93 holmium	68 **Er** 167.26 erbium	69 **Tm** 168.93 thulium	70 **Yb** 173.04 ytterbium	71 **Lu** 174.97 lutetium
96 **Cm** (247) curium	97 **Bk** (247) berkelium	98 **Cf** (251) californium	99 **Es** (252) einsteinium	100 **Fm** (257) fermium	101 **Md** (258) mendelevium	102 **No** (259) nobelium	103 **Lr** (260) lawrencium

Fundamental Physical Constants

Atomic mass unit	$1\ \text{amu} = 1.660539 \times 10^{-27}\ \text{kg}$ $1\ \text{g} = 6.022142 \times 10^{23}\ \text{amu}$
Avogadro's number	$N_A = 6.022142 \times 10^{23}/\text{mol}$
Electron charge	$e = 1.602176 \times 10^{-19}\ \text{C}$
Gas constant	$R = 8.314472\ \text{J}/(\text{mol}\cdot\text{K})$ $= 0.0820582\ (\text{L}\cdot\text{atm})/(\text{mol}\cdot\text{K})$
Mass of electron	$m_e = 5.485799 \times 10^{-4}\ \text{amu}$ $= 9.109382 \times 10^{-31}\ \text{kg}$
Mass of neutron	$m_n = 1.008665\ \text{amu}$ $= 1.674927 \times 10^{-27}\ \text{kg}$
Mass of proton	$m_p = 1.007276\ \text{amu}$ $= 1.672622 \times 10^{-27}\ \text{kg}$
Pi	$\pi = 3.1415926536$
Planck's constant	$h = 6.626069 \times 10^{-34}\ \text{J}\cdot\text{s}$
Speed of light in vacuum	$c = 2.99792458 \times 10^{8}\ \text{m/s}$

Useful Geometric Formulas

Perimeter of a rectangle $= 2l + 2w$

Circumference of a circle $= 2\pi r$

Area of a triangle $= (1/2)(\text{base} \times \text{height})$

Area of a circle $= \pi r^2$

Surface area of a sphere $= 4\pi r^2$

Volume of a sphere $= (4/3)\pi r^3$

Volume of a cylinder or prism = area of base × height

Important Conversion Factors

Length: SI unit = meter (m)

- 1 m = 39.37 in.
- 1 in. = 2.54 cm (exactly)
- 1 mile = 5280 ft = 1.609 km
- 1 angstrom (Å) = 10^{-10} m

Volume: SI unit = cubic meter (m^3)

- 1 L = 1000 cm^3 = 1.057 qt (U.S.)
- 1 gal (U.S.) = 4 qt = 8 pt
  = 128 fluid ounces
  = 3.785 L

Mass: SI unit = kilogram (kg)

- 1 kg = 2.205 lb
- 1 lb = 16 oz = 453.6 g
- 1 ton = 2000 lb
- 1 metric ton = 1000 kg = 1.103 tons
- 1 g = 6.022×10^{23} atomic mass units (amu)

Pressure: SI unit = pascal (Pa)

- 1 Pa = 1 N/m^2
- 1 bar = 10^5 Pa
- 1 atm = 1.01325×10^5 Pa (exactly)
  = 1.01325 bar
  = 760 mmHg
  = 760 torr (exactly)

Energy: SI unit = joule (J)

- 1 J = 1 N·m
- 1 cal = 4.184 J (exactly)
- 1 L·atm = 101.33 J

Temperature: SI unit = kelvin (K)

- K = °C + 273.15
- °C = (5/9) (°F − 32°)
- °F = (9/5) (°C) + 32°