Chapter 1. Atoms

Student Objectives

1.1 A Particulate View of the World: Structure Determines Properties

- Define **atoms**, **molecules**, and the science of **chemistry**.
- Represent a simple molecule, water, using spheres as atoms.

1.2 Classifying Matter: A Particulate View

- Define **matter** and distinguish between the three main states of matter: solid, liquid, and gas.
- Define and understand the difference between **crystalline** and **amorphous** solids.
- Define **mixture**, **pure substance**, **element**, **compound**, **heterogeneous**, and **homogeneous**.
- Differentiate between mixtures and pure substances; elements and compounds; and heterogeneous and homogeneous mixtures.
- Use the scheme on page 000 to classify matter.

1.3 The Scientific Approach to Knowledge

- Define and distinguish between a **hypothesis**, a **scientific law**, and a **theory**.
- Understand the role of experiments in testing hypotheses.
- State and understand the law of mass conservation as an example of scientific law.
- Understand that scientific theories are built from strong experimental evidence and that the term “theory” in science is used much differently than in pop culture.
- Understand the importance of reporting correct units with measurements.
- Know the differences between the three most common sets of units: English system, metric system, and International System (SI).
- Know the SI base units for length, mass, time, and temperature.

1.4 Early Ideas about the Building Blocks of Matter

- Describe the earliest definitions of atoms and matter (the Greeks).
- Know that greater emphasis on observation and the development of the scientific method led to the scientific revolution.

1.5 Modern Atomic Theory and the Laws That Led to It

- State and understand the law of conservation of mass (also from Section 1.4).
- State and understand the law of definite proportions.
- State and understand the law of multiple proportions.
- Know the four postulates of Dalton’s atomic theory.

1.6 The Discovery of the Electron

- Describe J. J. Thomson’s experiments with the cathode ray tube and understand how they provide evidence for the electron.
- Describe Robert Millikan’s oil-drop experiment and understand how it enables measurement of the charge of an electron.

1.7 The Structure of the Atom

- Define **radioactivity**, **nucleus**, **proton**, and **neutron**.
- Understand Thomson's plum-pudding model and how Ernest Rutherford's gold-foil experiment refuted it by giving evidence for a nuclear structure of the atom.

1.8 Subatomic Particles: Protons, Neutrons, and Electrons

- Define **atomic mass unit**, **atomic number**, and **chemical symbol**.
- Recognize chemical symbols and atomic numbers on the periodic table.
- Define **isotope**, **mass number**, and **natural abundance**.
- Determine the number of protons and neutrons in an isotope using the chemical symbol and the mass number.
- Define **ion**: **anion**, and **cation**.
- Understand how ions are formed from elements.

1.9 Atomic Mass: The Average Mass of an Element's Atoms

- Calculate atomic mass from isotope masses and natural abundances.
- Define **mass spectrometry** and understand how it can be used to measure mass and relative abundance.

1.10 Atoms and the Mole: How Many Particles?

- Understand the relationship between mass and count of objects such as atoms.
- Define **mole** and **Avogadro's number**.
- Calculate and interconvert between number of moles and atoms.
- Calculate and interconvert between number of moles and molecules.
- Calculate and interconvert between number of moles and particles.
- Calculate and interconvert between number of moles and mass.

1.11 The Origins of Atoms and Elements

- Relate how the Big Bang Theory explains formation of initial formation of hydrogen and helium.
- Explain the formation of heavy elements by fusion of lighter elements.

<u>Section Summaries</u>

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures, Tables, and Solved Examples

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

Lecture Outline

<table>
<tr><th>Terms, Concepts, Relationships, Skills</th><th>Figures, Tables, and Solved Examples</th></tr>
<tr><td>1.1 A Particulate View of the World: Structure Determines Properties
Definitions of atoms, molecules
Composition of water
Definition of chemistry
The particulate structure of matter determines properties.</td><td>Intro figure: portrayal of Disneyland ride, Adventure Thru Inner Space.
Unnumbered figures: model of H_2O, hypothetical linear H_2O molecule
Unnumbered figure: graphite and diamond structures</td></tr>
<tr><td>1.2 Classifying Matter: A Particulate View
States of matter: their definitions and some of their characteristics
gas
liquid
solid
Classification of matter
pure substance
element
compound
mixture
heterogeneous
homogeneous</td><td>Figure 1.1 The States of Matter
Figure 1.2 The Compressibility of Gases
Figure 1.3 The Classification of Matter According to Its Composition</td></tr>
</table>

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls

1.1 A Particulate View of the World: Structure Determines Properties
• Chemistry involves a great deal of what can't be seen directly, requiring representations and models.
    o The introductory figure shows hemoglobin, but the actual molecule is not a green and blue ribbon.
    o Chemists look at microscopic, macroscopic, and symbolic representations of atoms and molecules interchangeably. If you say "water," you might mean the formula H_2O, a molecular model, or a large collection of molecules (e.g., a glass of water). Students need help recognizing which representation to think about when a chemical name is used.
    o Particulate-level structure affects function: water would have different properties if the molecule was linear as opposed to bent. For example, we would expect water to be a gas at room temperature if the molecule were linear.	

1.2 The Classification of Matter • Properties of matter define its state: gas, liquid, or solid. Temperature is one example, and everyone recognizes steam, water, and ice. Ask for additional examples such as dry ice or liquid nitrogen. • Compressibility is a property that differentiates especially gases from liquids and solids. • Conceptual Connection 1.1 Pure Substances and Mixtures o Use of different shapes to represent atoms of different elements helps to reinforce the characteristics of particulate matter. • Classifying additional examples of matter (e.g., mayonnaise, Jell-O, and milk) according to the scheme demonstrates some of the challenges.	

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

1.3 The Scientific Approach to Knowledge
• Definitions of hypothesis, falsifiable, experiments, scientific law, theory
• Scientific method:
    o Observations and experiments lead to hypotheses.
    o More experiments may lead to a law and a theory.
    o A theory explains observations and laws.
• Creativity and subjectivity play important roles in science.
• Thomas S. Kuhn and Scientific Revolutions
• Scientific observations are quantifiable.
	• Unnumbered figure: painting of Antoine Lavoisier

1.4 Early Ideas about the Building Blocks of Matter • History of chemistry from antiquity (~450 BCE) • Scientific revolution (1400s-1600s)	

1.5 Modern Atomic Theory and the Laws That Led to It
• Law of conservation of mass
    o Matter is neither created nor destroyed.
    o Atoms at the start of a reaction may recombine to form different compounds, but all atoms are accounted for at the end.
    o Mass of reactants = mass of products.
• Law of definite proportions
    o Different samples of the same compound have the same proportions of constituent elements independent of sample source or size.
• Law of multiple proportions
• John Dalton's atomic theory
	• Unnumbered figure: models and photos of Na(s) and Cl_2(g) forming NaCl(s)
	• Example 1.1 Law of Definite Proportions
	• Unnumbered figure: models of CO and CO_2 illustrating the law of multiple proportions
	• Example 1.2 Law of Multiple Proportions

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.3 The Scientific Approach to Knowledge • Experiments test ideas. They are designed to support a hypothesis or to disprove it. Good scientific hypotheses must be testable or falsifiable. • Theories are developed only through considerable evidence and understanding, even though theories often are cited in popular culture as unproven or untested. • Conceptual Connection 1.2 Laws and Theories • Kuhn's book illustrates that science is not completely objective and immutable.	• Theories are *not* as easily dismissible as pop culture suggests. • Scientific knowledge constantly evolves as new information and evidence are gathered.
1.4 Early Ideas about the Building Blocks of Matter • The view of matter as made up of small, indestructible particles was ignored because more popular philosophers like Aristotle and Socrates had different views. • Leucippus and Democritus may have been proven correct, but they had no more evidence for their ideas than Aristotle did. • Observations and data led scientists to question models; the scientific method promotes the use of a cycle of such inquiry.	• Theories are not automatically accepted and may be unpopular for long periods of time. • Philosophy and religion can be supported by arguments; science requires that theories be testable and therefore falsifiable.
1.5 Modern Atomic Theory and the Laws That Led to It • That matter is composed of atoms grew from experiments and observations. • Conceptual Connection 1.3 The Law of Conservation of Mass • Investigating the law of definite proportions requires preparing or decomposing a set of pure samples of a compound like water. • Investigating the law of multiple proportions requires preparing or decomposing sets of pure samples from related compounds like NO, NO_2, and N_2O_5. • Conceptual Connection 1.4 The Laws of Definite and Multiple Proportions	• Measurements to establish early atomic theories were performed at the macroscopic level. The scientists observed properties for which they could collect data (e.g., mass or volume).

Lecture Outline

<table>
<tr><th><u>Terms, Concepts, Relationships, Skills</u></th><th><u>Figures, Tables, and Solved Examples</u></th></tr>
<tr><td>
1.6 The Discovery of the Electron

Thomson’s cathode ray tube experiments

High voltage produced a stream of particles that traveled in straight lines.
Each particle possessed a negative charge.
Thomson measured the charge-to-mass ratio of the electron.

Millikan’s oil-drop experiments

Oil droplets received charge from ionizing radiation.
Charged droplets were suspended in an electric field.
The mass and charge of each oil drop was used to calculate the mass and charge of a single electron.

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Figure 1.4 Cathode Ray Tube
Figure 1.5 Thomson’s Measurement of the Charge-to-Mass Ratio of the Electron
Unnumbered figure: properties of electrical charge
Figure 1.6 Millikan’s Measurement of the Electron’s Charge

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<tr><td>
1.7 The Structure of the Atom

Thomson’s plum-pudding model: negatively charged electrons in a sea of positive charge
Radioactivity

Alpha decay provides the alpha particles for Rutherford’s experiment.

Rutherford’s experiment

Alpha particles directed at a thin gold film deflect in all directions, including back at the alpha source.
Only a concentrated positive charge could cause the alpha particles to bounce back.

Rutherford’s nuclear theory

Most mass and all positive charge contained in a small nucleus
Most of atom by volume is empty space.
Protons: positively charged particles
Neutral particles with substantial mass also in nucleus

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Unnumbered figure: plum-pudding model
Figure 1.7 Rutherford’s Gold Foil Experiment
Figure 1.8 The Nuclear Atom
Unnumbered figure: scaffolding

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Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.6 The Discovery of the Electron • Review the attraction, repulsion, and additive nature of charges. • Discuss the physics of electric fields generated by metal plates. • A demonstration of a cathode ray tube will help students better understand Thomson's experiments. • Demonstrate how Millikan's calculation works and why he could determine the charge of a single electron. • Conceptual Connection 1.5 The Millikan Oil Drop Experiment	• Millikan did not measure the charge of a single electron; he measured the charge of a number of electrons and deduced the charge of a single electron.

1.7 The Structure of the Atom • It may be useful to give a brief description of radioactivity. Rutherford's experiment makes more sense if one knows some properties of the alpha particle and from where it comes. • Thomson identified electrons and surmised the existence of positive charge necessary to form a neutral atom. The plum-pudding model is the simplest way to account for the observations.	• Students often don't understand the *source* of alpha particles in Rutherford's experiments.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

1.8 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms
• Properties of subatomic particles
  ○ atomic mass units (amu)
    ➢ proton, neutron: ~1 amu
    ➢ electron: ~0.006 amu
  ○ charge
    ➢ relative value: –1 for electron, +1 for proton
    ➢ absolute value: 1.6×10^{-19} C
• Atomic number (number of protons): defining characteristic of an element
• Isotope: same element, different mass (different number of neutrons)
• Ion: atom with nonzero charge
  ○ anion: negatively charged (more electrons) non-metal
  ○ cation: positively charged (fewer electrons) metal
	• Table 1.1 Subatomic Particles
	• Figure 1.9 How Elements Differ
	• Figure 1.10 The Periodic Table
	• Unnumbered figure: portrait of Marie Curie
	• Unnumbered figures: Isotope notations
	• Unnumbered table: Neon Isotopes
	• Example 1.3 Atomic Numbers, Mass Numbers, and Isotope Symbols

1.9 Atomic Mass: The Average Mass of an Element's Atoms
• Average atomic mass is based on natural abundance and isotopic masses.
• Mass spectrometry
  ○ atoms converted to ions and deflected by magnetic fields to separate by mass
  ○ output data: relative mass versus relative abundance
	• Unnumbered figure: periodic table box for Cl
	• Example 1.4 Atomic Mass
	• Figure 1.11 The Mass Spectrometer
	• Figure 1.12 The Mass Spectrum of Chlorine
	• Unnumbered figure: mass spectrum of Ag

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.8 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms • Electrical charge can be demonstrated with static electricity. Two balloons charged with wool or human hair will repel each other. • Names of elements come from various sources. Tom Lehrer's "Element Song" can be found on the Internet. • Isotopic abundances are invariant in typical lab-sized samples because of such large numbers of atoms. • Conceptual Connection 1.6 Isotopes • Conceptual Connection 1.7 The Nuclear Atom, Isotopes, and Ions • The history of chemistry involves considerable cultural and gender diversity. Examples include both Lavoisiers (French), Dalton (English), Thomson (English), Marie Curie (Polish/French), Mendeleev (Russian), Millikan (American), Robert Boyle (Irish), Amedeo Avogadro (Italian).	• Students sometimes confuse the mass number as being equal to the number of neutrons, not the number of neutrons plus the number of protons. • Students logically (but mistakenly) presume that the mass of an isotope is equal to the sum of the masses of the protons and neutrons in that isotope.
1.9 Atomic Mass: The Average Mass of an Element's Atoms • The masses of isotopes must be reconciled with an element having only whole-number quantities of protons and neutrons; the values should be nearly integral since the mass of electrons is so small. • Mass spectrometry is an effective way to demonstrate where values of natural abundance are obtained. • Conceptual Connection 1.8:	• Students are tempted to calculate average atomic mass by adding together isotopic masses and dividing by the number of isotopes. • Atomic mass on the periodic table is usually not integral, even though elements have only whole numbers of protons and neutrons.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
1.10 Atoms and the Mole: How Many Particles • Mole concept and Avogadro's number • Converting between moles and number of atoms • Converting between mass and number of moles	• Unnumbered figure: pennies containing ~1 mol of Cu • Unnumbered figure: 1 tbsp of water contains ~1 mol of water • Example 1.5 Converting between Number of Moles and Number of Atoms • Figure 1.13 Molar Mass • Unnumbered figure: relative sizes of Al, C, He • Example 1.6 Converting between Mass and Amount (Number of Moles) • Unnumbered figure: conceptual plan for mass, mole, and atoms calculations • Example 1.7 The Mole Concept—Converting between Mass and Number of Atoms • Example 1.8 The Mole Concept

1.11 The Origins of Atoms and Elements
• Beginning of the universe according to the Big Bang Theory
  ○ Universe began as a collection of hot, dense matter.
  ○ Hydrogen and helium were the first elements.
  ○ Clouds of hydrogen and helium condensed to form stars and galaxies.
  ○ Heavier elements are formed from fusion of other lighter elements.
	• Unnumbered figure: Hubble Space Telescope image of the Eagle Nebula

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.10 Atoms and the Mole: How Many Particles? • Review the strategy for solving numerical problems: sort, strategize, solve, check. • Estimating answers is an important skill; the number of atoms will be very large (i.e., some large power of ten) even from a small mass or small number of moles. • Conceptual Connection 1.9 The Mole	• Many students are intimidated by estimating answers in calculations involving powers of ten.

1.11 The Origins of Atoms and Elements o Stars emit heat and light due to the large amount of energy generated by fusion	

Additional Problem for Converting between Number of Moles and Number of Atoms (Example 1.5)	**Calculate the number of moles of iron in a sample that has 3.83×10^{23} atoms of iron.**
Sort You are given a number of iron atoms and asked to find the amount of iron in moles.	**Given** 3.83×10^{23} atoms Fe **Find** mol Fe
Strategize Convert between number of atoms and number of moles using Avogadro's number.	**Conceptual Plan** Atoms $\rightarrow$ mol $\dfrac{1\text{ mol Fe}}{6.022\times10^{23}\text{ atoms Fe}}$ **Relationships Used** 6.022×10^{23} atoms = 1 mol (Avogadro's number)
Solve Follow the conceptual plan. Begin with 3.83×10^{23} Fe atoms and multiply by the ratio that equates moles and Avogadro's number.	**Solution** $3.83\times10^{23}\text{ atoms Fe}\times\dfrac{1\text{ mol Fe}}{6.022\times10^{23}\text{ atoms Fe}}$ $= 0.636\text{ mol Fe}$
Check	The sample was smaller than Avogadro's number, so the answer should be a fraction of a mole. The value of the sample has three significant figures, and the answer is provided in that form.

Additional Problem for Converting between Mass and Number of Moles (Example 1.6)	**Calculate the number of grams of silver in an American Silver Eagle coin that contains 0.288 mol of silver**
Sort You are given the amount of silver in moles and asked to find the mass of silver.	**Given** 0.288 mol Ag **Find** g Ag
Strategize Convert amount (in moles) to mass using the molar mass of the element.	**Conceptual Plan** mol Ag $\rightarrow$ g Ag $\frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}}$ **Relationships Used** 107.87 g Ag = 1 mol Ag
Solve Follow the conceptual plan to solve the problem. Start with 0.288 mol, the given number, and multiply by the molar mass of silver.	**Solution** $0.288 \cancel{\text{ mol Ag}} \times \frac{107.87 \text{ g Ag}}{1 \cancel{\text{ mol Ag}}} = 31.07 \text{ g Ag}$ $31.07 \text{ g} = 31.1 \text{ g Ag}$
Check	The magnitude of the answer makes sense since we started with an amount smaller than a mole. The molar amount and answer both have three significant figures.

Additional Problem for the Mole Concept—Converting between Mass and Number of Atoms (Example 1.7)	**What mass of iron (in grams) contains 1.20×10^{22} atoms of Fe? A paperclip contains about that number of iron atoms.**
Sort You are given a number of iron atoms and asked to find the mass of Fe.	**Given** 1.20×10^{22} atoms Fe **Find** g Fe
Strategize Convert the number of Fe atoms to moles using Avogadro's number. Then convert moles Fe into grams of iron using the molar mass of Fe.	**Conceptual Plan** Fe atoms $\rightarrow$ mol Fe $\rightarrow$ g Fe $\dfrac{1\text{ mol Fe}}{6.022\times10^{23}\text{ atoms Fe}}$ $\quad$ $\dfrac{55.85\text{ g Fe}}{1\text{ mol Fe}}$ **Relationships Used** 6.022×10^{23} atoms = 1 mol (Avogadro's number) 55.85 g Fe = 1 mol Fe
Solve Follow the conceptual plan to solve the problem. Begin with 1.20×10^{22} atoms of Fe, multiply by the ratio derived from Avogadro's number, and finally multiply by the atomic mass of Fe.	**Solution** $1.20\times10^{22}\ \cancel{\text{atoms Fe}} \times \dfrac{1\ \cancel{\text{mol Fe}}}{6.022\times10^{23}\ \cancel{\text{atoms Fe}}} \times \dfrac{55.85\text{ g Fe}}{1\ \cancel{\text{mol Fe}}} = 1.11\text{ g Fe}$
Check	The units and magnitude of the answer make sense. The sample is smaller than a mole. The number of atoms and mass both have three significant figures.

Additional Problem for the Mole Concept (Example 1.8)	**An iron sphere contains 8.55×10^{22} iron atoms. What is the radius of the sphere in centimeters? The density of iron is 7.87 g/cm^3.**
Sort You are given the number of iron atoms in a sphere and the density of iron. You are asked to find the radius of the sphere.	**Given** 8.55×10^{22} atoms Fe $d = 7.87\ \text{g/cm}^3$ **Find** radius (r) of a sphere
Strategize The critical parts of this problem are density, which relates mass to volume, and the mole, which relates number of atoms to mass: (1) Convert from the number of atoms to the number of moles using Avogadro's number; (2) Convert from the number of moles to the number of grams using the molar mass of iron; (3) Convert from mass to volume using the density of iron; (4) Find the radius using the formula for the volume of a sphere.	**Conceptual Plan** atoms Fe $\rightarrow$ mol Fe $\rightarrow$ g Fe $\rightarrow$ V (cm^3) $\dfrac{1\ \text{mol Fe}}{6.022\times10^{23}\ \text{atoms Fe}}$ $\quad\dfrac{55.85\ \text{g Fe}}{1\ \text{mol Fe}}$ $\quad\dfrac{1\ \text{cm}^3}{7.87\ \text{g Fe}}$ V (cm^3) $\rightarrow$ r (cm) $V = \dfrac{4}{3}\pi r^3$ **Relationships Used** 6.022×10^{23} atoms = 1 mol (Avogadro's number) 55.85 g Fe = 1 mol Fe d (density of Fe) = 7.87 g/cm^3 $V = 4/3\ \pi r^3$ [volume of a sphere with a radius of r]
Solve Follow the conceptual plan to solve the problem. Begin with 8.55×10^{22} Fe atoms and convert to moles, then to grams, and finally to a volume in cm^3. Solve for the radius using the rearranged equation.	**Solution** $8.55\times10^{22}\ \cancel{\text{atoms}} \times \dfrac{1\ \cancel{\text{mol Fe}}}{6.022\times10^{23}\ \cancel{\text{atoms}}} \times \dfrac{55.85\ \cancel{\text{g Fe}}}{1\ \cancel{\text{mol Fe}}}$ $\times \dfrac{1\,\text{cm}^3}{7.87\ \cancel{\text{g Fe}}} = 1.00757\ \text{cm}^3$ $r = \sqrt[3]{\dfrac{3V}{4\pi}} = \sqrt[3]{\dfrac{3\times1.00757\ \text{cm}^3}{4\pi}} = 0.622\ \text{cm}$
Check	The units (cm) are correct and the magnitude of the answer makes sense compared with previous problems.