Chemistry Structure and Properties 2nd edition Tro Solutions Manual

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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.

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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.

Section Summaries

Lecture Outline

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

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

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Lecture Outline

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

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
	• 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

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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.
  ○ The introductory figure shows hemoglobin, but the actual molecule is not a green and blue ribbon.
  ○ 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.
  ○ 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 ○ 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.	

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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:
  ○ Observations and experiments lead to hypotheses.
  ○ More experiments may lead to a law and a theory.
  ○ 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
  ○ Matter is neither created nor destroyed.
  ○ Atoms at the start of a reaction may recombine to form different compounds, but all atoms are accounted for at the end.
  ○ Mass of reactants = mass of products.
• Law of definite proportions
  ○ 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

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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).

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Lecture Outline

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

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.
	• 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

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
	• 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.

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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

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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.

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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

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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 ○ Stars emit heat and light due to the large amount of energy generated by fusion	

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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 $\frac{1\,\text{mol Fe}}{6.022 \times 10^{23}\ \text{atoms Fe}}$ **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** $3.83 \times 10^{23}\ \text{atoms Fe} \times \frac{1\,\text{mol Fe}}{6.022 \times 10^{23}\ \text{atoms Fe}}$ $= 0.636\ \text{mol Fe}$
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.

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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 g = 31.1 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.

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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 $\frac{1\,\text{mol Fe}}{6.022\times10^{23}\,\text{atoms Fe}}$ $\quad$ $\frac{55.85\,\text{g Fe}}{1\,\text{mol Fe}}$ **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** $1.20\times10^{22}\ \cancel{\text{atoms Fe}} \times \frac{1\ \cancel{\text{mol Fe}}}{6.022\times10^{23}\ \cancel{\text{atoms Fe}}}$ $\times \frac{55.85\,\text{g Fe}}{1\ \cancel{\text{mol Fe}}} = 1.11\,\text{g Fe}$
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.

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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 $d = 7.87$ 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) $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ atoms Fe}}$ $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ $\frac{1 \text{ cm}^3}{7.87 \text{ g Fe}}$ V (cm^3) $\rightarrow$ r (cm) $V = \frac{4}{3}\pi r^3$ **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** $8.55 \times 10^{22} \cancel{\text{atoms}} \times \frac{1 \cancel{\text{ mol Fe}}}{6.022 \times 10^{23} \cancel{\text{ atoms}}} \times \frac{55.85 \cancel{\text{ g Fe}}}{1 \cancel{\text{ mol Fe}}}$ $\times \frac{1 \text{ cm}^3}{7.87 \cancel{\text{ g Fe}}} = 1.00757 \text{ cm}^3$ $r = \sqrt[3]{\frac{3V}{4\pi}} = \sqrt[3]{\frac{3 \times 1.00757 \text{ cm}^3}{4\pi}} = 0.622$ cm
Check	The units (cm) are correct and the magnitude of the answer makes sense compared with previous problems.

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