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SKU: 9780321971944-SOLUTIONS

Chapter 2. Atoms and Elements

Objectives

2.1 Images and Moving Individual Atoms

- Describe scanning tunneling microscopy and how atoms are imaged on surfaces.
- Define atom and element.

2.2 Modern Atomic Theory and the Laws That Led to It

- State and understand the law of conservation of mass.
- State and understand the law of definite proportions.
- State and understand the law of multiple proportions.
- Be able to calculate and distinguish the difference between the law of conservation of mass, law of definite proportions, and law of multiple proportions.
- Describe John Dalton's atomic theory.

2.3 The Discovery of the Electron

- Describe the cathode ray.
- Describe the properties of electrical charges and the electrostatic forces.
- Describe J. J. Thomson's experiments with the cathode ray tube and understand how it provides evidence for the electron.
- Describe Robert Millikan's oil drop experiment and understand how it enables measurement of the charge on an electron.

2.4 The Structure of the Atom

- Define radioactivity, alpha particles, beta particles, and gamma particles.
- Describe the plum pudding model and how nucleus, protons, and neutrons play a role in this theory.
- Discuss the nuclear theory.
- Differentiate between the nucleus, protons, and electrons.
- Describe Rutherford's experiment with radioactive alpha particles directed at gold foil and understand how it provides evidence for the structure of the atomic nucleus.

2.5 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms

- Define atomic mass units (amu), atomic number, and chemical symbol.
- Understand the role of chemical symbols and atomic numbers in a periodic chart.
- Define and understand isotopes, mass number, and natural abundance.
- Calculate the mass number from the atomic number (i.e., the number of protons or which element) and the number of neutrons.
- Define ions, anions, and cations.
- Understand how ions are formed from elements and how the number of electrons in ions differs from elements.

2.6 Finding Patterns: The Periodic Law and the Periodic Table

- Define the periodic law.

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- Understand how the periodic table groups elements that have similar properties into columns.
- Define and differentiate among metals, nonmetals, and metalloids.
- Differentiate between main-group and transition elements.
- Know the families or groups—noble gases, alkali metals, alkaline earth metals, and halogens.
- Know and understand the origin of elements that form ions with predictable charges.

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

- Understand how mass number and natural abundance of isotopes determine atomic mass.
- Calculate atomic mass from mass number and natural abundance.

2.8 Molar Mass: Counting Atoms by Weighing Them

- Understand the relationship between count and mass of objects.
- Understand the relationship between mole and Avogadro's number.
- Calculate and interconvert between number of moles and atoms.
- Calculate and interconvert between number of moles and mass.
- Utilize dimensional analysis to calculate moles to atoms and/or moles to mass.

Section Summaries

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures and Solved Problems

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

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Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved Problems
2.1 Images and Moving Individual Atoms • Description of scanning tunneling microscopy (STM) • Introduction to macroscopic and microscopic perspectives • Definitions of atom and element	• Figure 2.1 Scanning Tunneling Microscopy [A schematic of the technique and atomic resolution that results.] • Figure 2.2 Imaging Atoms [Actual images of iodine atoms on a platinum surface and Japanese characters written with atoms.]

2.2 Modern Atomic Theory and the Laws That Led to It
• Law of Conservation of Mass
  ○ Matter is neither created nor destroyed.
  ○ Atoms at start of reaction are all in the products.
  ○ Mass of reactants = mass of products.
• Law of Definite Proportions
  ○ Samples of the same compound have the same proportions of constituent elements independent of sample source or size.
  ○ Ratio of masses of O to H in H_2O is 16:2 or 8:1.
• Law of Multiple Proportions
  ○ Elements in a compound occur as ratios of whole numbers.
• John Dalton and Atomic Theory
  ○ Elements are composed of particles called atoms.
  ○ All atoms of an element have the same mass and properties and can't change into another element.
  ○ Compounds are made of atoms in ratios of small, whole numbers.
	• [unnumbered figure] A reaction of sodium and chlorine to make sodium chloride. [A specific mass of solid sodium and gaseous chloride gives the same total mass of sodium chloride salt.]
	• Solved Problem, Example 2.1 [Ratio of masses of carbon and oxygen from carbon dioxide samples.]
	• Space-filling models of carbon monoxide and carbon dioxide showing ratios of mass and atom types.
	• Solved Problem, Example 2.2 [Ratios of N to O in nitrogen dioxide and dinitrogen monoxide.]

2.3 The Discovery of the Electron
• Early experiments establish existence of subatomic particles.
• Thomson's cathode ray tube identified the electron:
  ○ High voltage produced a stream of particles that traveled in straight lines.
  ○ Particles originated independently of composition of material.
  ○ Each particle possessed a negative charge.
• Millikan's oil drops used to deduce charge on single electron:
  ○ Oil droplets received charge from ionizing radiation.
  ○ Charged drops suspended in an electric field.
  ○ The mass and charge of each oil drop were used to calculate the mass and charge of a single electron.
	• Figure 2.3 Cathode Ray Tube [A schematic drawing and photograph of a CRT.]
	• Figure 2.4 Thomson's Measurement of the Charge-to-Mass of the Electron [A schematic drawing and description of the experiment.]
	• [unnumbered figure in margin] Properties of Electrical Charge [Attraction and repulsion of charges; summing charges]
	• Figure 2.5 Millikan's Measurement of the Electron's Charge [A schematic drawing and description of the experiments.]

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Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls

2.1 Images and Moving Individual Atoms
• Scanning tunneling microscopy (STM) is a modern discovery and one that shows a microscopic view of atoms.
  ○ Measured current transformed into 3-D image.
  ○ Technique requires conducting surface like a metal.
• It is useful to reiterate the distances and analogies about size—the one used compares an atom to a grain of sand.
	• STM is not actually showing images of atoms like one might imagine seeing with a light microscope.
	• Atoms are not colored spheres; the images use color to accent the differences.

2.2 Modern Atomic Theory and the Laws That Led to It
• That matter is composed of atoms grew from experiments and observations.
• The conservation of mass would be difficult to interpret until all the reactants and products were characterized; for example, colorless and odorless gases like oxygen and carbon dioxide would be hard to detect.
  ○ Conceptual Connection 2.1 (What happens to most of the mass of a log when it burns) can be used as a 1-minute discussion within small groups.
• Investigating the law of definite proportions requires preparing or decomposing a set of pure samples of a molecule like water.
• Investigating the law of multiple proportions requires preparing or decomposing sets of pure samples from related molecules like CO and CO_2.
  ○ Conceptual Connection 2.2 (The Laws of Definite and Multiple Proportions) asks for an explanation of the difference.
	• Measurements to establish early atomic theories were performed at the macroscopic level; the scientists observed properties for which they could collect data (i.e., mass or volume).

2.3 The Discovery of the Electron • Review the attraction, repulsion, and additivity of charges. • Review the physics of electric fields generated by metal plates. • A live demonstration of a cathode ray tube would be useful; CRT televisions and computer monitors are slowly disappearing. • The mass and charge on Millikan's oil drops was very large compared to a single electron. • Demonstrate how Millikan's calculation works and why he could determine a single charge. • Conceptual Connection 2.3 (The Millikan Oil Drop Experiment) asks students to determine the number of electrons represented by a given charge value.	• Millikan did not measure a single charge on an electron; he measured a large number of them and calculated the value.

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Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved Problems

2.4 The Structure of the Atom
• Thomson's "plum pudding" model (which is a pudding with raisins throughout)—negatively charged electrons in a positively charged sphere.
• Radioactivity occurs in three forms.
  ○ gamma rays (energy)
  ○ alpha- or α-particles (positive charge)
  ○ beta- or β-particles (negative charge)
• Rutherford's experiment elucidates structure of nucleus:
  ○ Alpha particles directed at a thin gold film deflected in all directions, including back at the α source.
  ○ Only a large positive charge could cause the α-particles to bounce back.
• Rutherford's nuclear theory
  ○ Most mass and all positive charge contained in small nucleus.
  ○ Most of atom is empty space with electrons dispersed in it.
  ○ Positively charged particles are called protons and are in the nucleus.
  ○ Neutral particles having substantial mass also appear in the nucleus.
	• [unnumbered figure] Plum-pudding model
	• Figure 2.6 Rutherford's Gold Foil Experiment [A schematic drawing and description of the experiment.]
	• Figure 2.7 The Nuclear Atom [A comparison of the plum-pudding model and Rutherford's suggested alternative—the modern nucleus.]
	• [unnumbered figure] Scaffolding [occurs mostly as open space but provides solid support]

2.5 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms
• Properties of subatomic particles
  ○ atomic mass units or amu
  ▪ proton or neutron ~1 amu
  ▪ electron ~0.0006 amu
  ○ charge
  ▪ relative value electron = –1 vs. proton = +1
  ▪ absolute value = 1.6×10^{-19} C
• Element and Symbols
  ○ Atomic number defines element
  ○ Isotope: form of element with different # neutrons
  ○ Ion: form with different # electrons & protons
  ▪ anion = extra electron(s) = negatively charged
  ▪ cation = fewer electron(s) = positively charged
	• [unnumbered figure] Electrical discharge during a thunderstorm
	• Table 2.1 Subatomic Particles [mass in kg and amu; relative charge and in Coulombs]
	• Figure 2.8 How Elements Differ [micro & macro view of helium and carbon]
	• Figure 2.9 The Periodic Table [atomic number, symbol, name]
	• [unnumbered figure] Marie Curie [namesake of curium]
	• Table of isomers of neon [including number of protons, neutrons, mass, natural abundance]
	• Solved Problem, Example 2.3 Atomic Numbers, Mass Numbers, and Isotope Symbols [Cl-35 and Cr-52]

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Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.4 The Structure of the Atom • Discuss forms of radioactivity—particles and electromagnetic waves. Rutherford's experiment makes more sense if one knows some properties of an alpha particle. • Thomson identified electrons and surmised the existence of positive charges necessary to form a neutral atom. The plum-pudding model is the simplest way to account for the observations: Occam's Razor = all things being equal, the simplest explanation is the best. • A simple analogy to Rutherford's experiment could be a billiards table. You can shoot the cue ball and miss or deflect off the object ball. The analogy is not complete since the object ball and cue ball have the same mass. But the cue ball would bounce back in bumper pool if it hits a stationary post. • The figure about scaffolds supports discussion about an atom being mostly open space but still being able to have rigidity and strength in the macroscopic view. This is another example of apparent differences between the microscopic and macroscopic properties.	• Many science textbooks, particularly ones other than chemistry, still support the Bohr model—the analogy of planets orbiting around the sun. The idea of orbitals and not orbits comes later but it may be useful to start laying the groundwork.

2.5 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms • Another mass analogy is a ping-pong ball and a glass sphere of the same size (air, 0.0013 g/cm^3, glass = 2.4 g/cm^3), a ratio of ca. 1800. • Electrical charge can be demonstrated with static electricity. Two balloons can be charged with wool or human hair and will repel each other. • Names of elements come from various sources, including some controversies in recent years. Tom Lehrer's "Element Song" can be found on the Internet and shows that a mathematician can master the list of them. • Exposure to isotopes may be limited: carbon-14 dating; hydrogen, deuterium, and tritium. Medicine, particularly radiology, has many examples of radioactive forms. • Conceptual Connection 2.4 (Isotopes) can be used as a 1-minute group exercise for drawing the nucleus of two isotopes and using percent natural abundance. • Conceptual Connection 2.5 (The Nuclear Atom and Isotopes) The statements of the differences among them can serve as the basis of a small group discussion or a clicker question. • The history of chemistry involves considerable cultural and gender diversity. Examples to this point include both Lavoisier (in Ch. 1, French), Dalton (English), Thomson (English), Marie Curie (Polish/French), Mendeleev (Russian), Millikan (American).	• Students are likely to relate both the mass and size of the baseball and rice grain. • Elements, isotopes, and ions test students' understanding of differences among numbers of electrons, protons, and neutrons.

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Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved Problems

2.6 Finding Patterns: The Periodic Law and the Periodic Table
• Periodic law and the periodic table
  ○ arrange by mass
  ○ observe recurring properties
• Major divisions of the periodic table
  ○ metals, nonmetals, metalloids (& semiconductors)
  ○ main-group elements, transition elements (transition metals)
  ○ numbering systems used
• Family or group
  ○ noble gases
  ○ alkali metals
  ○ alkali earth metals
  ○ halogens
• Predicting ions
  ○ Group 1A = cation (+)
  ○ Group 2A = dication (2+)
  ○ Group 3A = trication (3+)
  ○ Group 5A = trianion (3–)
  ○ Group 6A = dianion (2–)
  ○ Group 7A = anion (–)
	• [unnumbered figure] Dmitri Mendeleev [Russian postage stamp]
	• Figure 2.10 Recurring Properties [elements with similar properties]
	• Figure 2.11 Making a Periodic Table [a simple table with color coding for similar properties]
	• Figure 2.12 Major Divisions of the Periodic Table [metals, nonmetals, metalloids] with Representative Metals, Metalloids, and Nonmetals [pictures of samples]
	• Figure 2.13 Main-Group and Transition Elements [color-coded table]
	• Figure 2.14 Elements with Predictable Charges [charge of common ions by group]
	• Solved Problem, Example 2.4 Predicting the Charge of Ions [predict charge with explanation of thought process]

2.7 Atomic Mass: The Average Mass of an Element's Atoms
• Atomic mass
  ○ natural abundance
  ○ isotopic mass
	• [unnumbered figure] Data from periodic table [atomic number, symbol, atomic mass, and name for chlorine]
	• Solved Problem, Example 2.5 Atomic Mass [calculate atomic mass from isotopic natural abundance]

2.8 Molar Mass: Counting Atoms by Weighing Them
• Counting items
• Moles and Avogadro's number
• Moles, mass, and atomic mass (in amu)
  ○ Converting between number of moles and number of atoms
  ○ Converting between mass and amount
• The mole concept
	• [unnumbered figure] Mass of 22 pennies
	• [unnumbered figure] Tablespoon of water
	• Solved Problem, Example 2.6 Converting between Number of Moles and Number of Atoms
	• [unnumbered figure] A two-pan balance to compare mass
	• Solved Problem, Example 2.7 Converting between Mass and Amount (Number of Moles)
	• Solved Problem, Example 2.8 The Mole Concept—Converting between Mass and Number of Atoms
	• Solved Problem, Example 2.9 The Mole Concept (volume occupied by atoms)

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Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.6 Finding Patterns: The Periodic Law and the Periodic Table • Alternative periodic tables were presented earlier in the history of the periodic law. • Periodic table development used to illustrate role of laws and theories discussed in Chapter 1. • Other graphic displays of the modern periodic table can be found in journals [Schwartz, *J. Chem. Educ.*, 2006, *83*, 849; Moore, *J. Chem. Educ.*, 2003, *80*, 847; Bouma, *J. Chem. Educ.*, 1989, *66,* 741.], books, and on the Web. • Periodic tables, particularly electronic ones, are arranged according to the periodic law but can compare many features (e.g., phase of matter, size of atoms, common ions). The book shows these as a series of figures.	• The periodic table is best at predicting microscopic properties although macroscopic ones are also often illustrated.
2.7 Atomic Mass: The Average Mass of an Element's Atoms • The atomic mass on the periodic table must be reconciled with an element having only whole number quantities of protons and neutrons—the values should be near integral amu values since electrons are so small. • Conceptual Connection 2.6 Atomic Mass (Non-calculator question using understanding of percent abundance) can be used for a clicker question.	• Atomic mass on the periodic table is often not an integral number of amu even though elements can have only protons and neutrons (each with ca. 1 amu of mass); electron mass is negligible.
2.8 Molar Mass: Counting Atoms by Weighing Them • Define a strategy that can be used to solve the numerical problems: Sort, Strategize, Solve, and Check. • Estimating answers is an important skill—the number of molecules will be a very large number (i.e., some power of ten) even from a small mass or number of moles. • Moles make it possible (and easy) to interconvert mass and number of molecules. • Conceptual Connection 2.7 Avogadro's Number (significance and significant figures in the value) can be used as the basis of a small group discussion. • Conceptual Connection 2.8 The Mole (estimating moles using mass and atomic mass) can be used as the basis of a small group activity.	• Most students believe that a calculator is necessary to attempt any of these problems presented in the section.

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Additional Problem for Converting between Number of Moles and Number of Atoms (Example 2.6)	**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} Fe atoms **Find** mol Fe
Strategize Convert between number of atoms and number of moles using Avogadro's number.	**Conceptual Plan** atoms $\rightarrow$ mol $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ **Relationships Used** 6.022×10^{23} = 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** $3.83 \times 10^{23} \text{ Fe atoms} \times \frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}} = 0.636 \text{ mol Fe}$
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 3 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 2.7)	**Calculate the number of grams of silver in an American Silver Eagle coin that contains 0.288 mole 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 3 significant figures.

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Additional Problem for the Mole Concept—Converting between Mass and Number of Atoms (Example 2.8)	**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} Fe atoms **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 $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ **Relationships Used** $6.022 \times 10^{23} = 1 \text{ mol}$ (Avogadro's number) $55.85 \text{ g Fe} = 1 \text{ mol Fe}$
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** $1.20 \times 10^{22} \text{ Fe atoms} \times \frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ = 1.11 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 3 significant figures.

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Additional Problem for the Mole Concept (Example 2.9)	**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} Fe atoms $d = 7.87$ 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** Fe atoms $\rightarrow$ mol Fe $\rightarrow$ g Fe $\rightarrow$ V (cm^3) $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ $\frac{1 \text{ cm}^3}{7.87 \text{ g Fe}}$ V (cm^3) $\rightarrow$ r (cm) $V = \frac{4}{3}\pi r^3$ **Relationships Used** $6.022 \times 10^{23} = 1$ 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** $8.55 \times 10^{22} \text{ atoms} \times \frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ atoms}} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \times \frac{1 \text{ cm}^3}{7.87 \text{ g Fe}} = 1.00757 \text{ cm}^3$ $r = \sqrt[3]{\frac{3V}{4\pi}} = \sqrt[3]{\frac{3 \times 1.00757 \text{ cm}^3}{4\pi}} = 0.622$ cm
Check	The units (cm) are correct, and the magnitude of the answer makes sense compared with previous problems.

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