CHAPTER 2
ATOMS, MOLECULES, AND IONS

Chapter Learning Goals for Students

Section 2.1 Explain the law of a) conservation of mass and b) definite proportions.

Section 2.2 Explain the law of multiple proportions.

For two different compounds comprised of the same two elements, show that the law of multiple proportions is obeyed.

Section 2.3 Explain information about the atom revealed by the experiments of a) Thomson and b) Millikan, c) Rutherford.

Describe the atom in terms of the composition, mass, and volume of the nucleus relative to the mass and volume occupied by the electrons.

Section 2.4 Explain the information about the atom revealed by the experiments of Rutherford.

Describe the atom in terms of the composition, mass, and volume of the nucleus relative to the mass and volume occupied by the electrons.

Perform calculations using atomic size.

Section 2.5 State the difference between atoms and elements.

Write and interpret isotope symbols, determining the number of protons, neutrons, and electrons.

Section 2.6 Given the mass and natural abundance of all isotopes of a given element, calculate the average atomic mass of that element.

For any element, calculate a) the mass in grams of a single atom, the number of moles of atoms, and c) the number of atoms in a given number of grams.

Section 2.7 Define nucleon, nuclide, and nuclear reaction, summarizing the differences between nuclear reactions and chemical reactions and how they are described by their respective nuclear equations and chemical equations.

Section 2.8 Define radioactivity and radioisotope.

Write balanced equations for nuclear reactions, identifying the types of radiation and radioisotopes involved.

Section 2.9 Use the neutron/proton plot for stable isotopes to determine whether a given nuclide is expected to be stable or unstable.

Identify nuclear decay processes and the identity of isotopes involved in nuclear decay processes.

Section 2.10 Classify substances as a) pure substances or mixtures and b) elements or chemical compounds.

State the difference between a) atoms and molecules, b) covalent bonds and ionic bonds, c) chemical formulas and structural formulas, and d) ball-and-stick models and space-filling models.

Relate molecular models to structural formulas and chemical formulas.

Section 2.11 Know charges of monatomic and polyatomic anions and cations.

Classify substances as ionic or molecular and bonding as ionic bonding or covalent bonding.

Section 2.12 Give the formulas and names of a) common polyatomic ions, b) ionic compounds, and c) binary molecular compounds.

Lecture Outline

2.1. The Conservation of Mass and the Law of Definite Proportions[1]

A. Element – a substance that cannot be further broken down

B. Law of Mass Conservation – Mass is neither created nor destroyed in chemical reactions.

C. Law of Definite Proportions – Different samples of a pure chemical substance always contain the same proportion of elements by mass; elements do not combine chemically in random proportion.

2.2. The Law of Multiple Proportions and Dalton's Atomic Theory[2]

A. Dalton's Atomic Theory
1. Elements made of tiny particles called atoms
2. Each element characterized by the mass of its atoms
    a. Atoms of the same element – identical masses
    b. Atoms of different elements – different masses
3. Atoms combine in small, whole-number ratios – form new substances, called compounds
4. Chemical reactions only rearrange the way that atoms are combined; atoms themselves are not changed.

B. Law of Multiple Proportions – If two elements combine in different ways to form different substances, the mass ratios are small, whole-number multiples of each other.

2.3. Atomic Structure: Electrons[3]

A. Thomson – found that cathode rays consist of tiny, negatively charged particles called electrons
1. Electrons are emitted from electrodes made of two thin pieces of metal.
2. Many different metals may be used to make electrodes – all contain electrons.
3. Cathode rays can be deflected by bringing either a magnet or an electrically charged plate near the tube. Deflection depends on the
    a. strength of the deflecting magnetic or electric field.
    b. size of the negative charge on the electron.
    c. mass of the electron.
4. Charge-to-mass ratio, e/m of the electron = $1.758\ 819 \times 10^8$ C/g

B. Millikan determined that the charge on a drop of oil was always a small, whole-number multiple of *e*. (See Fig. 2.4 in the textbook.)

---

[1] Test Item File Questions: Multiple Choice 1-7, 141-144; Algorithmic 1-4; Short Answer 1
[2] Test Item File Questions: Multiple Choice 8-14, 145-149; Algorithmic 5; Short Answer 2
[3] Test Item File Questions: Multiple Choice 15-16; Short Answer 3

1. $e = 1.602\ 177 \times 10^{-19}$ C
2. Knowing the values for *e/m* and *e* for an electron, *m* can be calculated.
   a. $m = 9.109\ 390 \times 10^{-28}$ g

2.4. Atomic Structure: Protons and Neutrons[4]

A. Rutherford directed a beam of alpha particles at a thin gold foil.
   1. Alpha (α) particles
      a. 7000 times more massive than electrons
      b. Have a positive charge twice the magnitude of, but opposite in sign to, the charge on an electron
   2. Beam of α particles
      a. Most pass through the thin gold foil
      b. A few deflected at large angles
B. Nuclear model of the atom
   1. Nucleus
      a. A tiny central core in an atom where the mass of the atom is concentrated
      b. Contains the atom's positive charges
   2. Electrons move in space a relatively large distance away from the nucleus.
C. Nucleus composed of two kinds of particles
   1. Protons
      a. Mass = $1.672\ 623 \times 10^{-24}$ g
      b. Positively (+) charged
      c. Number of protons = number of electrons in a neutral atom
   2. Neutrons
      a. Neutron mass ≈ proton mass
      b. Charge = 0

2.5. Atomic Numbers[5]

A. Elements differ from one another according to the number of protons in their atoms.
B. Atomic number (Z) = the number of protons in an atom = the number of electrons in an atom
C. Most nuclei also contain neutrons.
D. Mass number (A) = number of protons + number of neutrons in an atom
E. Isotopes – atoms with identical atomic numbers but different mass numbers
   1. Mass number written as left superscript
   2. Atomic number (Z) written as left subscript (The atomic number is sometimes left off because all atoms of an element always contain the same number of protons.)
   3. Number of neutrons in an isotope calculated from $A - Z$
   4. Number of neutrons in an atom has little effect on chemical properties of the atom.

2.6. Atomic Masses and the Mole[6]

A. Atomic mass unit (amu)
   1. Exactly 1/12th the mass of an atom of ${}^{12}_{6}C$
   2. 1 amu = $1.660\ 539 \times 10^{-24}$ g
B. Isotopic mass
   1. Mass of an atom in atomic mass units
   2. Numerically close to the atom's mass number
C. Atomic mass values – weighted averages for the naturally occurring mixtures of isotopes
   1. Atomic mass of an element = Σ(mass of each isotope × the fraction of the isotope)
   2. Σ used for the term "the sum of"
   3. Use atomic masses to count number of atoms by weighing a sample of the element.
D. One mole of any element is the amount whose mass in grams is numerically equal to its atomic mass.

---

[4] Test Item File Questions: Multiple Choice 17-21; Short Answer 4
[5] Test Item File Questions: Multiple Choice 22-35; Algorithmic 6-12; Short Answer 5-10
[6] Test Item File Questions: Multiple Choice 36-44; Short Answer 11-14

1. Mole (mol) = 6.022×10^{23} of anything
   a. Avogadro's number (abbreviated N_A)
   b. 1 mol N atoms = 6.022×10^{23} atoms of N
   c. Likewise, 1 mol electrons = 6.022×10^{23} electrons
2. Importance of the mole – provides a relationship between numbers of atoms and masses of atoms

E. Molar mass of an element
1. One mole of any element has a mass equal to its atomic mass in grams
2. Mass of 6.022×10^{23} atoms of an element
3. Serves as a conversion factor between numbers of atoms and mass

2.7. Nuclear Chemistry: The Change of One Element Into Another[7]

A. An atom is characterized by its atomic number, Z, and its mass number, A.
1. Z – written as a subscript to the left of the element symbol; gives the number of protons in the nucleus
2. A – written as a superscript to the left of the element symbol, gives the total number of nucleons
3. Nucleon – a general term for both protons (p) and neutrons (n)
4. Uranium-238

   $^{238}_{92}\mathrm{U}$

   a. 92 protons
   b. 238 – 92 = 146 neutrons
   c. 238 nucleons

B. Isotopes – atoms with identical atomic numbers, but different mass numbers

C. Nuclide – nucleus of a specific isotope

D. Nuclear reaction – a reaction that changes the nucleus

$$^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th}$$

E. Differences between nuclear reactions and chemical reactions
1. Nuclear reactions – a change in an atom's nucleus
   a. In nuclear reactions electrons are beta particles

   $^{0}_{-1}\mathrm{e} = \beta^-$

   b. Superscript 0 – the mass of an electron is essentially zero when compared with that of a proton or neutron
   c. Subscript (–1) – the charge of an electron is –1
2. Chemical reaction – a change in distribution of the outer shell electrons around an atom
3. Different isotopes of an element
   a. Have essentially the same behavior in chemical reactions
   b. Have a different behavior in nuclear reactions
4. Changes in temperature or pressures or the addition of a catalyst do not affect the rate of a nuclear reaction.
5. The nuclear reaction of an atom is essentially the same regardless of whether the atom is in an element or a compound.
6. The energy changes accompanying nuclear reactions are far greater than those accompanying chemical reactions.

2.8. Radioactivity[8]

A. Radioisotope – an unstable isotope
1. Radioactive nuclei
2. Radioactive – the spontaneous decomposition (decay) of an unstable isotope to form a more stable isotope of the same element or of another element and emission of radiation

---

[7] Test Item File Questions: Multiple Choice 45-51; Algorithmic13; Short Answer 15-16

[8] Test Item File Questions: Multiple Choice 52-72, 153-162; Algorithmic 14-15; Short Answer 17-20

3. Nuclear reaction – reaction that changes the nucleus as described by a nuclear equation

B. Three common types of radiation with different properties
1. Alpha (α)
2. Beta (β)
3. Gamma (γ)

C. Alpha (α) radiation – a stream of particles that consists of two protons and two neutrons
1. Repelled by a positively charged electrode
2. Attracted by a negatively charged electrode
3. A helium ion

$$^{4}_{2}He^{2+} = \alpha$$

4. Emission of an α particle reduces the mass number of the nucleus by four and reduces the atomic number by two.

$$^{238}_{92}U \rightarrow {}^{4}_{2}He + {}^{234}_{90}Th$$

5. Common for heavy radioactive isotopes (radionuclides)

D. Beta (β) radiation
1. Attracted by a positively charged electrode
2. Repelled by a negatively charged electrode
3. In nuclear reactions electrons are beta particles

$$^{0}_{-1}e = \beta^{-}$$

4. Superscript 0 – the mass of an electron is essentially zero when compared with that of a proton or neutron
5. Subscript (–1) – the charge of an electron is –1
6. Occurs when a neutron in the nucleus spontaneously decays into a proton plus an electron, which is then ejected.
7. Emission of a beta particle does not affect the mass number but increases the atomic number by one.

$$^{14}_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$$

E. Gamma (γ) radiation – electromagnetic radiation of very high energy and short wavelength
1. Not attracted by a positively charged electrode
2. Not repelled by a negatively charged electrode
3. No mass
4. Stream of high-energy photons
5. Almost always accompanies α and β emission – mechanism for release of energy
6. γ emission often not shown in nuclear equations – doesn't change either the mass number or the atomic number of the product nucleus

F. Positron emission and electron capture
1. Positron emission – conversion of a proton in the nucleus into a neutron plus an ejected positron, a particle with the same mass as an electron but opposite charge:

$$^{0}_{1}e = \beta^{+}$$

   a. Decrease in atomic number of the product nucleus
   b. No change in mass number
2. Electron capture – a proton in the nucleus captures an inner-shell electron, which is converted into a neutron
   a. No change in mass number of the product nucleus
   b. Atomic number decreases by one

G. Balanced nuclear reaction
1. Not balanced in usual chemical sense – nuclei are not the same on both sides of the reaction
2. Sums of nucleons on both sides are equal.
3. Sums of charges on nuclei and any elementary particles on both sides are equal.

4. Not concerned with ionic charges on atoms
    a. Irrelevant to nuclear disintegration
    b. Disappear during reaction

2.9. Nuclear Stability[9]

A. Stable radioactive isotope – one that can be prepared and whose half-life can be measured
  1. Unstable – those isotopes that can't be prepared or that decay too rapidly for their half-lives to be measured
  2. Nonradioactive (stable indefinitely) – isotopes that do not undergo radioactive decay

B. Neutron/proton ratio in the nucleus determines if an isotope is radioactive.

C. Plot of number of neutrons (*y*-axis) versus number of protons (*x*-axis) – textbook Figure 2.8
  1. Stable nuclides – fall in a curved band called the peninsula of nuclear stability
  2. Sea of instability
     a. Area on either side of the peninsula of nuclear stability
     b. Represents the large number of unstable neutron/proton combinations
  3. Island of stability – area predicted to exist for a few super heavy nuclides near 114 protons and 184 neutrons

D. Generalizations from plot
  1. Every element has at least one radioactive isotope.
  2. Hydrogen is the only element whose most abundant stable isotope contains more protons than neutrons.
  3. The ratio of neutrons to protons gradually increases for elements heavier than calcium.
  4. All isotopes beyond bismuth-209 are radioactive.
  5. Nonradioactive isotopes generally have an even number of neutrons.

E. Neutrons function as a kind of nuclear glue that holds nuclei together by overcoming proton–proton repulsions.

F. Magic numbers of protons or neutrons — 2, 8, 20, 28, 50, 82, 126.
  1. Give rise to particularly stable nuclei
  2. Nucleus with a magic number of either protons or neutrons is unusually stable.
  3. Analogous to chemical stability brought about by an octet of electrons

G. Trends apparent from the peninsula of nuclear stability
  1. Elements with an even atomic number have a larger number of nonradioactive nuclides than do elements with an odd atomic number.
  2. Radioactive nuclei on the right side of the peninsula undergo nuclear disintegration by positron emission, electron capture, or alpha emission.
     a. Lower neutron/proton ratio
     b. Processes increase the neutron/proton ratio.
  3. Nuclei on the left side of the peninsula emit beta particles.
     a. Higher neutron/proton ratios
     b. Process decreases the neutron/proton ratio.

H. Some radioisotopes can't reach a nonradioactive nucleus in a single emission – undergo a decay series of disintegrations

2.10. Mixtures and Chemical Compounds; Molecules and Covalent Bonds[10]

A. Different kinds of matter on earth classified as either pure substances or mixtures – textbook Figure 2.10

B. Pure substance
  1. Elements
  2. Compounds

---

[9] Test Item File Questions: Multiple Choice 73-86; Algorithmic 16-18; Short Answer 21

[10] Test Item File Questions: Multiple Choice 87-91, 132-136, 163-164, 166, 168, 171, 178-180, 182-184; Algorithmic19-22; Short Answer 22

C. Chemical compounds – pure substance formed from the combination of atoms of two or more different elements
  1. Have constant composition
  2. Composition indicated by a chemical formula
    a. Lists the symbols of individual constituent elements
    b. Number of each atom is given by subscript
  3. Formed when atoms undergo chemical combination in a specific manner
  4. Transformation from elements to compound = chemical reaction

D. Chemical bonds – connections that join atoms together in a compound
  1. Formed by atom's electrons
  2. Classified as:
    a. Covalent bonds – occur between two nonmetals
    b. Ionic bonds – occur between a metal and a nonmetal

E. Covalent bond – two atoms share electrons
  1. Molecule – unit of matter that results when two or more atoms joined by covalent bonds
  2. Structural formula
    a. Shows specific connections between atoms
    b. Contains more information than the chemical formula
  3. Some elements exist as molecules: H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2

2.11. Ions and Ionic Bonds[11]

A. Ionic Bond – complete transfer of one or more electrons from one atom to another
  1. Formed between metals and nonmetals
    a. Metal – gives up electrons
    b. Nonmetal – accepts electrons
  2. Ions – charged particles resulting from loss or gain of electrons
    a. Cation – positively charged particle resulting from loss of one or more electrons
    b. Anion – negatively charged particle resulting from gain of one or more electrons
    c. Polyatomic ions
      i. Charged, covalently bonded groups of atoms
      ii. Charged molecules – specific numbers and kinds of atoms joined in a definite way by covalent bonds
  3. Ionic solids
    a. Cations and anions packed together in a regular manner
    b. Charges cancel

2.12. Naming Chemical Compounds[12]

A. Binary Ionic Compounds – ionic compounds containing only two elements, a cation and an anion
  1. Identify cation first, then anion
    a. Cation
      i. Same name as the element
      ii. Remember metals form cations
    b. Anion
      i. First part of its name from the element
      ii. Adds the ending *ide*
      iii. Remember nonmetals form anions
  2. Common main group and transition metal ions – textbook Figures 2.14 and 2.15
    a. Elements within a group often form similar kinds of ions.
    b. Main-group metal cations: charge = group number
    c. Main-group nonmetal anions: charge = group number – 8

---

[11] Test Item File Questions: Multiple Choice 92-114, 137-140, 165, 167, 169-170, 172-177, 181; Algorithmic 23-37; Short Answer 23-25

[12] Test Item File Questions: Multiple Choice 115-131; Algorithmic 38-50; Short Answer 26-27

- d. Some metals form more than one kind of cation.
  - i. Charge indicated by a Roman numeral in parentheses
  - ii. Common for transition metal complexes
3. Electrical neutrality – cations and anions combine in such a manner that the overall charge on a compound is equal to zero
   - a. Total positive charge = total negative charge
   - b. Determine number of positive charges on the cation by counting the number of negative charges on the anion (and vice versa)
4. Formulas for ionic compounds always contain the smallest whole number ratio of cation to anion.

B. Binary molecular compounds – molecular compounds (nonmetals) containing only two elements
1. One of the elements is more cation-like – takes the name of the element
2. One of the elements is more anion-like – takes an *ide* ending
3. Character depends on relative positions of the two elements in the periodic table
   - a. More cationlike – farther left in the periodic table
   - b. More anionlike – farther right in the periodic table
4. To specify the numbers of each element present, use numerical prefixes (textbook Table 2.2) – the *mono* prefix not used for the atom named first
5. When naming binary molecular compounds that contain hydrogen, it is necessary to indicate whether the molecule is in the gaseous or aqueous (in water) state.
   - a. If molecule a gas
     - i. Use above rules.
     - ii. Indicate gaseous state with (*g*) after formula
   - b. If molecule in aqueous solution
     - i. Name compound as a binary acid (see below)
     - ii. Indicate aqueous state with (*aq*) after formula

C. Ionic compounds containing polyatomic ions named by following rules for naming binary ionic compounds.
1. Identify cation
2. Identify anion
   - a. Names, formulas, and charge numbers of the most common polyatomic anions found in textbook Table 2.3
   - b. Most names end with the suffix *ite* or *ate*
   - c. Several pairs of ions related by presence or absence of hydrogen atom
3. Oxoanions – an atom of the same element combined with different numbers of oxygen atoms
   - a. Learn the name and formula of the ion whose name ends with *ate* (including the charge on the anion).
   - b. Add one O; add prefix *per*.
   - c. Remove one O; change *ate* to *ite*.
   - d. Remove two O's; add prefix *hypo* and change *ate* to *ite*.

CHAPTER 2

ATOMS, MOLECULES, AND IONS

Teaching Tips, Points of Emphasis, and Common Misconceptions

Section 2.2 Dalton did not know about isotopes – atoms of the same element that have different masses due to different numbers of neutrons. Challenge students to determine which of Dalton's postulates are incorrect.

Section 2.4 Demonstrate the relative insignificance of an electron's mass by calculating its percent contribution to the approximate mass of a simple atom, such as carbon-12.

Section 2.6 The use of carbon-12 to define the amu is another example of a system of measurement based on an arbitrary standard. In earlier times the amu was based on oxygen-16 and hydrogen-1.

Section 2.6 The average atomic mass of an element is a weighted, not a simple, average.

Section 2.6 An effective example showing the difference between simple average and weighted average is the use of a hypothetical student's test scores of 100, 100, 100, 60. The weighted average = 90; since there are only two different grades, 100 and 60, the simple average is 80.

Section 2.7 Energy changes for nuclear reactions are on the order of a million times greater than energy changes for chemical reactions.

Section 2.8 An alpha (α) particle ($^{4}_{2}He^{2+}$) is a helium nucleus, but the +2 charge is not used in writing nuclear equations.

Section 2.8 A beta (β) particle ($^{0}_{-1}e$ or β^{-}) is an electron.

Section 2.8 For the purpose of balancing nuclear equations, the mass of the electron is taken to be 0 and its "nuclear charge" is taken to be –1.

Section 2.8 Electron capture: proton + core electron $\rightarrow$ neutron.

Section 2.8 For a nuclear equation to be balanced: sum of atomic numbers of reactants (subscripts) = sum of atomic numbers of products, and sum of nucleons of reactants (superscripts) = sum of nucleons of products. Ionic charges are ignored.

Section 2.9 Transuranium elements, those with atomic numbers higher than uranium, do not occur naturally but are produced by nuclear transmutation reactions.

Section 2.10 The element oxygen is written O_2, P_4, and S_8 are other examples of elements that occur as polyatomic species.

Section 2.11 NaCl is a *simplest formula* or *empirical formula*. It does not reveal that the compound is made up of ions ($Na^{+}Cl^{-}$), nor that it is an extended system $(NaCl)_x$.

Section 2.11 *Polyatomic ions* are also known as *molecular ions* because they are molecules with a charge.

Section 2.11 Some students think HCl is ionic because it dissolves in water to form H^+ and Cl^-. The interaction of a molecule and a solvent can substantially change the properties of the molecule.

Section 2.12 In compounds, elements on the left of the periodic table tend to be relatively positive, and elements on the right tend to be relatively negative.

Lecture/Laboratory Demonstration References

Section 2.5 Arthur B. Ellis, Edward A. Adler, and Frederick H. Juergens, "Dramatizing Isotopes: Deuterated Ice Cubes Sink," *J. Chem. Educ.,* Vol. 67, **1990**, 159-160.

Literature References

Section 2.1 John J. Fortman, "Analogical Demonstration," *J. Chem. Educ.,* Vol. 69, **1992**, 323-324. Discusses the law of conservation of mass and the law of multiple proportions.

Section 2.3 Doris Eckey, "A Millikan Oil Drop Analogy," *J. Chem. Educ.,* Vol. 73, **1996**, 237-238.

Section 2.4 Mary V. Lorentz, "Bowling Balls and Beads, A Concrete Analogy to the Rutherford Experiment," *J. Chem. Educ.,* Vol. 65, **1988**, 1082.

Section 2.4 Barrie M. Peake, "The Discovery of the Electron, Proton, and Neutron," *J. Chem. Educ.,* Vol. 66, **1989**, 738.

Section 2.5 William Spindel and Takanobu Ishida, "Isotope Separation," *J. Chem. Educ.,* Vol. 68, **1991**, 312-318.

Section 2.6 Arthur M. Last and Michael J. Webb, "Using Monetary Analogies to Teach Average Atomic Mass," *J. Chem. Educ.,* Vol. 70, **1993**, 234-235.

Section 2.6 John J. Fortman, "Pictorial Analogies IV: Relative Atomic Weights," *J. Chem. Educ.,* Vol. 70, **1993**, 235-236.

Section 2.6 Josefina Arce de Sanabia, "Relative Atomic Mass and the Mole: A Concrete Analogy to Help Students Understand These Abstract Concepts," *J. Chem. Educ.,* Vol. 70, **1993**, 233-234.

Section 2.7 Enrique A. Hughes and Anita Zalts, "Radioactivity in the Classroom," *J. Chem. Educ.,* Vol. 77, **2000**, 613-614.

Section 2.7 "Nuclear Chemistry: State of the Art for Teachers," a series of articles in the October 1994 issue of the *Journal of Chemical Education*.

Section 2.7 S. G. Hutchinson and F. I. Hutchinson, "Radioactivity: in Everyday Life," *J. Chem. Educ.,* Vol. 74, **1997**, 501-505.

Section 2.7	Charles H. Atwood, "Teaching Aids for Nuclear Chemistry," *J. Chem. Educ.,* Vol. 71, **1994**, 845-847.
Section 2.8	C. Ronneau, "Radioactivity: A Natural Phenomenon," *J. Chem. Educ.,* Vol. 67, **1990**, 736-737.
Section 2.8	Robert Suder, "Beta Decay Diagram," *J. Chem. Educ.,* Vol. 66, **1989**, 231.
Section 2.9	Donald J. Olbris and Judith Herzfeld, "Nucleogenesis! A Game with Natural Rules for Teaching Nuclear Synthesis and Decay," *J. Chem. Educ.,* Vol. 76, **1999**, 349-352.
Section 2.9	Darleane C. Hoffman and Diana M. Lee, "Chemistry of the Heaviest Elements – One Element at a Time," *J. Chem. Educ.,* Vol. 76, **1999**, 332-347.
Section 2.12	Gerhard Lind, "Teaching Inorganic Nomenclature: A Systematic Approach," *J. Chem. Educ.,* Vol. 69, **1992**, 613-614.
Section 2.12	Steven J. Hawkes, "A Mnemonic for Oxy-Anions," *J. Chem. Educ.,* Vol. 67, **1990**, 149.
Interlude	V. E. Viola, "Teaching Nuclear Science: A Cosmological Approach," *J. Chem. Educ.,* Vol. 71, **1994**, 840-844.

Media References

Section 2.1	Electrolysis of Water movie from the Instructor Resource Center DVD
Section 2.1	Conservation of Mass activity from the Instructor Resource Center DVD
Section 2.2	Multiple Proportions movie from the Instructor Resource Center DVD
Section 2.2	Multiple Proportions activity from the Instructor Resource Center DVD
Section 2.3	Millikan Oil Drop Experiment movie from the Instructor Resource Center DVD
Section 2.3	Separation of Rays activity from the Instructor Resource DVD
Section 2.4	Rutherford Experiment movie from the Instructor Resource Center DVD
Section 2.5	Atomic Number activity from the Instructor Resource Center DVD
Section 2.5	Carbon Isotopes activity from the Instructor Resource Center DVD
Section 2.6	Isotopes of Hydrogen activity from the Instructor Resource Center DVD
Section 2.9	Nuclear Stability activity from the Instructor Resource Center DVD
Section 2.9	Half-Life activity from the Instructor Resource DVD
Section 2.10	Mixtures and Compounds activity from the Instructor Resource Center DVD

Section 2.12 Main-Group Ions activity from the Instructor Resource Center DVD

Section 2.12 Nonmetal Anions activity from the Instructor Resource Center DVD

Section 2.12 Transition Metal Ions activity from the Instructor Resource Center DVD

Section 2.12 Naming Polyatomic Ions activity 1 from the Instructor Resource Center DVD

Section 2.12 Naming Polyatomic Ions activity 2 from the Instructor Resource Center DVD

Section 2.12 Naming Polyatomic Ions activity 3 from the Instructor Resource Center DVD