CHAPTER 2
ATOMS, MOLECULES, AND IONS

Chapter Learning Goals for Students

Section 2.1 Use symbols to represent element names.

Section 2.2 Identify the location of metals, nonmetals, and semimetals on the periodic table.

Indicate the atomic number, group number, and period number for an element whose position in the periodic table is given.

Identify groups as main group, transition metal group, or inner transition metal group.

Section 2.3 Specify the location and give examples of elements in the alkali metal, alkaline earth metal, halogen, and noble gas groups.

Use the properties of an element to classify it as metal, nonmetal, or semimetal and give its location in the periodic table.

Section 2.4 Determine the mass of the products in a reaction using the law of mass conservation.

Section 2.5 Demonstrate the law of multiple proportions using mass composition of two compounds of the same elements.

Section 2.6 Describe Thomson's cathode-ray experiment and what it contributed to the current model of atomic structure.

Describe Millikan's oil drop experiment and what it contributed to the current model of atomic structure.

Section 2.7 Describe Rutherford's gold foil experiment and what it contributed to the current model of atomic structure.

Describe the structure and size of the atom.

Calculate the number of atoms in a sample given the size of the atom.

Section 2.8 Determine the mass number, atomic number, and number of protons, neutrons, and electrons from an isotope symbol.

Write the isotope symbol for elements.

Section 2.9 Calculate atomic weight given the functional abundance and mass of each isotope.

Convert between grams and numbers of moles or atoms using molar mass and Avogadro's number.

Identify an element given the mass and number of atoms or moles.

Section 2.10 Classify molecular representations of matter as a mixture, pure substance, element, or compound.

Convert between structural formulas, ball-and-stick models, and chemical formulas.

Section 2.11 Classify bonds as ionic or covalent.

Determine the number of electrons and protons from chemical symbol and charge.

Match the molecular representation of an ionic compound with its chemical formula.

Section 2.12 Convert between name and formula for binary ionic compounds.

Convert between formula and name for ionic compounds with polyatomic ions.

Convert between name and formula for binary molecular compounds.

Lecture Outline

2.1. Chemistry and the Elements[1]

A. Element – fundamental substance that can't be chemically changed or broken down into anything simpler
   1. Chemical symbol –represents a specific element
      a. Capitalize first letter
      b. Lower case used for second letter if present

B. Periodic table – tabular organization of all 114 elements

2.2. Elements and the Periodic Table[2]

A. Creation of the periodic table – ideal example of how scientific theory comes into being
   1. Random observations
   2. Organization of data in ways that make sense
   3. Consistent hypothesis emerges
      a. Explains known facts
      b. Makes predictions about unknown phenomena

B. Mendeleev's hypothesis about organizing known chemical information – meets criteria for a good hypothesis
   1. Listed the known elements by atomic weight
   2. Grouped them together according to their chemical reactivity
   3. Was able to predict the properties of unknown elements – *eka*–aluminum, *eka*–silicon

C. Periodic table – grid of the elements arranged in 7 horizontal rows and 18 vertical columns
   1. Periods – seven horizontal rows in the periodic table
   2. Groups – 18 vertical columns in the periodic table
      a. Groups numbered 1A $\rightarrow$ 8A and 1B $\rightarrow$ 8B (or 1 $\rightarrow$18)
      b. Actually have 32 groups – lanthanides (14 elements after lanthanum) and actinides (14 elements after actinium) not included in the group numbers
      c. Elements in a given group have similar chemical properties.

D. The periodic table of the elements is the most important organizing principle of chemistry.
   1. Regular progression in size of the seven periods – reflects a similar regularity in atomic structure
   2. Main Group (or Representative) Elements – Groups 1A-8A; (two larger groups on the left and six larger groups on the right of the table)
   3. Transition Metal Elements – Groups 1B-8B; (10 smaller groups in the middle of the table)

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[1]Test Item File Questions: Multiple Choice 1 – 3, Algorithmic 1 – 5, Short Answer 1 – 2
[2]Test Item File Questions: Multiple Choice 4 – 7, Algorithmic 9, Short Answer 3 – 5

4. Inner Transition Metal (or Rare Earth) Elements (14 groups shown separately at the bottom of the table)

2.3. Some Common Groups of Elements and Their Properties[3]

A. Property – any characteristic used to describe or identify matter
   1. Physical properties – characteristics not involving chemical change of the sample
   2. Chemical properties – properties that do change chemical makeup of the sample
   3. Intensive properties – sample size-independent properties
   4. Extensive properties – sample size-dependent properties

B. Elements within a group have similar chemical properties.
   1. Group 1A – Alkali metals
      a. Lustrous, silvery metals
      b. React rapidly with water to form highly alkaline products
   2. Group 2A – Alkaline earth metals
      a. Lustrous, silvery metals
      b. Less reactive than alkali metals
   3. Group 7A – Halogens
      a. Corrosive, nonmetallic elements
      b. Salt formers
   4. Group 8A – Noble gases: gases with low reactivity

C. Three major classes of elements
   1. Metals
      a. Largest category of elements
      b. Found on the left side of the periodic table (left of the heavy zigzag line)
         i. Solids (except mercury)
         ii. Malleable
         iii. Ductile – can be drawn into thin wires without breaking
         iv. Conduct heat and electricity
   2. Nonmetals
      a. Relatively small number
      b. Found on the right side of the periodic table (right of the heavy zigzag line)
         i. Gases, liquids, or solids
         ii. Brightly colored
         iii. Brittle solids
         iv. Poor conductors of heat and electricity
   3. Semimetals (metalloids)
      a. Elements adjacent to the zigzag boundary between metals and nonmetals
      b. Properties fall between metals and nonmetals
         i. Brittle
         ii. Poor conductors of heat and electricity

2.4. Observations Supporting Atomic Theory: The Conservation of Mass and the Law of Definite Proportions[4]

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.

2.5. The Law of Multiple Proportions and Dalton's Atomic Theory[5]

A. Dalton's Atomic Theory
   1. Elements made of tiny particles called atoms

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[3]Test Item File Questions: Multiple Choice 8 – 15, Algorithmic 10 – 33, Short Answer 6 – 8,
[4]Test Item File Questions: Multiple Choice 16 – 19, Algorithmic 34 – 38, Short Answer 9
[5]Test Item File Questions: Multiple Choice 20 – 28, Algorithmic 39, Short Answer 10

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.

2.6. Atomic Structure: Electrons[6]

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. $e = 1.602\ 177 \times 10^{-19}$ 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

2.7. Atomic Structure: Protons and Neutrons[7]

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

2.8. Atomic Numbers[8]

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.

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[6]Test Item File Questions: Multiple Choice 29 – 30, Short Answer 11
[7]Test Item File Questions: Multiple Choice 31 – 34, Algorithmic 40, Short Answer 12
[8]Test Item File Questions: Multiple Choice 35 – 50, Algorithmic 41 – 49, Short Answer 13 – 18

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.

2.9. Atomic Weights and the Mole[9]

A. Unified atomic mass unit (u) previously the atomic mass unit (amu)
  1. Exactly 1/12th the mass of an atom of ${}^{12}_{6}C$
  2. 1 u = $1.660\ 539 \times 10^{-24}$ g
  3. Unit also called a Dalton (Da)
B. Isotopic mass
  1. Mass of an atom in atomic mass units
  2. Numerically close to the atom's mass number
C. Atomic massweight values – weighted averages for the naturally occurring mixtures of isotopes
  1. Average atomic mass of an element (atomic weight) = Σ(mass of each isotope $\times$ the fraction of the isotope)
  2. Σ used for the term "the sum of"
  3. Use atomic weights 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 weight.
  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 weight 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

2.10. 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
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:

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[9]Test Item File Questions: Multiple Choice 51 – 55, Algorithmic 50 – 54, Short Answer 19 – 22
[10]Test Item File Questions: Multiple Choice 56, Algorithmic 55 – 62, Short Answer 23

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 2.15 and 2.16
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
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. Ionic compounds containing polyatomic ions named by following rules for naming binary ionic compounds.

[11]Test Item File Questions: Multiple Choice 57 – 60, Algorithmic 63 – 79, Short Answer 24 – 26
[12]Test Item File Questions: Multiple Choice 61 – 77, Algorithmic 80 – 111, Short Answer 27 – 28

1. Identify cation
2. Identify anion
    a. Names, formulas, and charge numbers of the most common polyatomic anions found in textbook Table 2.5
    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*.

C. Binary molecular compounds – molecular compounds (nonmetals) containing only two elements
1. One of the elements is more cationlike – takes the name of the element
2. One of the elements is more anionlike – 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 2.6) – 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

How is the Principle of Atom Economy Used to Minimize Waste in a Chemical Synthesis?

CHAPTER 2
ATOMS, MOLECULES, AND IONS

Teaching Tips, Points of Emphasis, and Common Misconceptions

Section 2.1 Elements *can* be broken into simpler substances (protons, neutrons, electrons, etc.), but the energy required is greater than available under ordinary chemical conditions.

Section 2.2 Iodine is a silvery-gray solid that sublimes to a purple gas. The brown solution that was once used as an antiseptic medicine is known as tincture of iodine (iodine dissolved in alcohol).

Section 2.2 Because elements in the same group tend to exhibit similar chemical behavior, a group is also called a *family*.

Section 2.3 While elements within the same group tend to have similar chemical properties, due to its small size the first member of each main group exhibits a chemistry that differs most from the other members of the group.

Section 2.3 Hydrogen is a unique element. Although listed under group 1A, hydrogen is not a metal under ordinary conditions. In its ionic chemistry, hydrogen sometimes behaves as though it belongs in group 7A. In covalent compounds, hydrogen behaves as though it belongs between boron and carbon.

Section 2.5 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 2.7 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 2.9 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 2.9 The atomic weight of an element is a weighted, not a simple, average.

Section 2.9 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 2.10 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 2.2 James L. Marshall, "A Living Periodic Table," *J. Chem. Educ.,* Vol. 77, **2000**, 979-983. Describes a portable and permanent collection of 87 elements.

Section 2.8 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 P.G. Nelson, "Important Elements," *J. Chem. Educ*., Vol. 68, 1991, 732-737.

Section 2.1 Vivi Ringes, "Origin of the Names of Chemical Elements," *J. Chem. Educ*., Vol. 66, **1989**, 731-738.

Section 2.2 Milton J. Wieder, "It's Elementary," *J. Chem. Educ.,* Vol. 78, **2001**, 468-469.

Section 2.4 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 2.5 Doris Eckey, "A Millikan Oil Drop Analogy," *J. Chem. Educ.,* Vol. 73, **1996**, 237-238.

Section 2.7 Mary V. Lorentz, "Bowling Balls and Beads, A Concrete Analogy to the Rutherford Experiment," *J. Chem. Educ.,* Vol. 65, **1988**, 1082.

Section 2.7 Barrie M. Peake, "The Discovery of the Electron, Proton, and Neutron," *J. Chem. Educ.,* Vol. 66, **1989**, 738.

Section 2.8 William Spindel and Takanobu Ishida, "Isotope Separation," *J. Chem. Educ.,* Vol. 68, **1991**, 312-318.

Section 2.9 Arthur M. Last and Michael J. Webb, "Using Monetary Analogies to Teach Average Atomic Mass," *J. Chem. Educ.,* Vol. 70, **1993**, 234-235.

Section 2.9 John J. Fortman, "Pictorial Analogies IV: Relative Atomic Weights," *J. Chem. Educ.,* Vol. 70, **1993**, 235-236.

Section 2.9 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 2.12 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.

Media References

Section 2.1 Names of Elements activity 1 from the Instructor Resource Center DVD

Section 2.1 Names of Elements activity 2 from the Instructor Resource Center DVD

Section 2.2 Periodic Property movie from the Instructor Resource Center DVD

Section 2.3 Sodium and Potassium in Water movie from the Instructor Resource Center DVD

Section 2.3 Physical Properties of the Halogens movie from the Instructor Resource Center DVD

Section 2.3 Periodic Table Groups activity from the Instructor Resource Center DVD

Section 2.4 Electrolysis of Water movie from the Instructor Resource Center DVD

Section 2.4 Conservation of Mass activity from the Instructor Resource Center DVD

Section 2.25 Multiple Proportions movie from the Instructor Resource Center DVD

Section 2.5 Multiple Proportions activity from the Instructor Resource Center DVD

Section 2.6 Millikan Oil Drop Experiment movie from the Instructor Resource Center DVD

Section 2.6 Separation of Rays activity from the Instructor Resource DVD

Section 2.7 Rutherford Experiment movie from the Instructor Resource Center DVD

Section 2.8 Atomic Number activity from the Instructor Resource Center DVD

Section 2.58 Carbon Isotopes activity from the Instructor Resource Center DVD

Section 2.9 Isotopes of Hydrogen activity from the Instructor Resource Center 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

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