CHAPTER 2
ATOMS, MOLECULES, AND IONS

Chapter Learning Goals

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.

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.

Perform calculations using atomic size.

Section 2.5 State the difference between atoms and elements.

Given the symbol for an isotope of an atom or ion, determine 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, b) the number of atoms in a given number of grams, and c) the number of grams for a mole of an element or a compound.

Section 2.7 State the difference between a) compounds and mixtures and b) heterogeneous and homogeneous mixtures.

Section 2.8 State the difference between a) atoms and molecules, b) covalent bonds and ionic bonds, c) chemical formulas, d) structural formulas, and molecular models, and e) ball-and-stick models and space-filling models.

Relate structural formulas to chemical formulas.

Identify which substances are ionic and which are molecular.

Section 2.9 Identify which substances are acids and which are bases.

Section 2.10 Give the formulas and names of a) common polyatomic ions, b) ionic compounds, c) binary molecular compounds, d) common acids.

Lecture Outline

2.1. Conservation of Mass and the Law of Definite Proportions

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. The ratio of atoms is the same in each molecule of a compound.

2.2. Dalton's Atomic Theory and the Law of Multiple Proportions

A. Dalton's Atomic Theory
   1. Elements made of tiny particles called atoms
   2. Each element is 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. Mass is conserved in chemical reactions.

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. The Structure of Atoms: Electrons

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

2.4. The Structure of Atoms: Protons and Neutrons

A. Rutherford directed a beam of alpha particles at a thin gold foil.
   1. Alpha (α) particles – Helium nuclei, He^{2+}
      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 Number

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 a neutral atom
C. All nuclei (except for Hydrogen) 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 Mass

A. Atomic mass unit (amu)
    1. Exactly 1/12th the mass of an atom of ${}^{12}_{6}C$
    2. 1 amu = $1.660\ 54 \times 10^{-24}$ 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.

2.7. Compounds and Mixtures

A. Different kinds of matter on earth classified as either pure substances or mixtures – textbook Fig. 2.7
B. Pure substance include both Elements and Compounds.
C. Chemical compounds – pure substance formed from the combination of atoms of two or more different elements
    1. Have constant composition – Law of Definite Proportions
    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 reactions
    1. Atoms from two or more different elements combine to form compounds.
    2. New materials created with properties completely unlike those of their constituent elements.
    3. Chemical equation – format used to write a chemical reaction
        a. Reactants
            i. Starting substances in chemical reaction (beginning materials).
            ii. Written on left side of arrow
        b. Products
            i. New materials formed (ending materials).
            ii. Written on right side of arrow
        c. Arrow placed between reactants and products to indicate a transformation
E. Mixtures
    1. Two or more substances blended together in random proportion.
    2. Individual substances unchanged chemically

a. Homogeneous mixture (solution) – uniform properties throughout
b. Heterogeneous mixture – regions with differing compositions

2.8. Molecules, Ions, and Chemical Bonds

A. 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 (sharing of electrons)
b. Ionic bonds – occur between a metal and a nonmetal (transfer of electrons)

B. 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 alone
3. Some elements exist as molecules: H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2, P_4, S_8

C. 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 (also called molecular 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

2.9. Acids and Bases

A. Two important ions: H^+ and OH^-
1. Fundamental to the concept of acids and bases

B. Acid
1. Substance that provides H^+ ions when dissolved in water
2. Produces H^+ along with the corresponding anion (conjugate base)

C. Base
1. Substance that provides OH^- ions when dissolved in water
2. Produces OH^- along with the corresponding cation (conjugate acid)

2.10. Naming Chemical Compounds

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 and polyatomic anions
2. Common main group and transition metal ions – textbook Figures 2.11 and 2.12
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. 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 – classes of anions formed by 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*.

D. Oxoacids – contain oxygen in addition to hydrogen and another element
   1. Yields H^+ and a polyatomic oxoanion when dissolved in H_2O (textbook Table 2.4)
   2. Names related to the names of the corresponding oxoanions
      a. Replace suffix *ite* with *ous acid*
      b. Replace suffix *ate* with *ic acid*

E. Binary acids – contain hydrogen and one other element or a polyatomic anion that is not an oxoanion
   1. Aqueous solutions
   2. Use the prefix *hydro* followed by the first part of the name of the other element, the suffix *ic* and acid.
   3. When writing the formula, indicate aqueous solution by (*aq*).