Chemistry The Central Science 14th edition Brown Solutions Manual

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Chapter 2. Atoms, Molecules, and Ions

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Chapter 2. Atoms, Molecules, and Ions

Common Student Misconceptions

- Students have problems with the concept of amu.
- Students often think that mass number and atomic number can be used interchangeably.
- Students think that the term *isotope* is synonymous with being a harmful, radioactive substance.
- Beginning students often do not see the difference between empirical and molecular formulas.
- Students think that polyatomic ions can easily dissociate into smaller ions.
- Students often fail to recognize the importance of the periodic table as a tool for organizing and remembering chemical facts.
- Students often cannot relate the charges on common monatomic ions to the position of their parent atoms in the periodic table.
- Students often do not realize that an ionic compound *can* consist of nonmetals only, e.g., $(NH_4)_2SO_4$.
- Students often confuse the guidelines for naming ionic compounds with those for naming binary molecular compounds.
- Students routinely underestimate the importance of this chapter.

Teaching Tips

- It is critical that students learn the names and formulas of common and polyatomic ions as soon as possible. They sometimes need to be told that this information will be used throughout their careers as chemists (even if that career is only one semester).
- Remind students that *families* or *groups* are the columns in the periodic table; *periods* are the rows.
- Emphasize to students that the subscripts in the molecular formula of a substance are always an integral multiple of the subscripts in the empirical formula of that substance.

Lecture Outline

2.1 The Atomic Theory of Matter[1]

- Greek Philosophers: Can matter be subdivided into fundamental particles?
- Democritus (460–370 BC): All matter can be divided into indivisible *atomos*.
- Dalton: proposed atomic theory with the following postulates:
  - Elements are composed of atoms.
  - All atoms of an element are identical.
  - In chemical reactions, atoms are not changed into different types of atoms. Atoms are neither created nor destroyed.
  - Compounds are formed when atoms of elements combine.
- **Atoms** are the building blocks of matter.
- *Law of constant composition*: The relative kinds and numbers of atoms are constant for a given compound.
- **Law of conservation of mass** (*matter*): During a chemical reaction, the total mass before the reaction is equal to the total mass after the reaction.
  - Conservation means something can be neither created nor destroyed. Here, the term applies to matter (mass). Later we will apply it to energy (Chapter 5).
- **Law of multiple proportions**: If two elements, A and B, combine to form more than one compound, then the mass of B combines with the mass of A in a ratio of small whole numbers.
- Dalton's theory *predicted* the law of multiple proportions.

[1] "Analogical Demonstration" from Further Readings

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FORWARD REFERENCES

- The law of conservation of mass (matter) falls under the first law of thermodynamics (Chapter 5).

2.2 The Discovery of Atomic Structure

- By 1850 scientists knew that atoms consisted of charged particles.
- **Subatomic particles** are those particles that make up the atom.
- Recall the law of electrostatic attraction**:** like charges repel, and opposite charges attract.

Cathode Rays and Electrons[2,3,4,5]

- **Cathode rays** were first discovered in the mid-1800s during studies of electrical discharge through partially evacuated tubes (cathode-ray tubes, or CRTs).
  - Computer terminals were once popularly referred to as CRTs (cathode-ray tubes).
  - Cathode rays = radiation produced when high voltage is applied across the tube.
- The voltage causes negative particles to move from the negative electrode (cathode) to the positive electrode (anode).
- The path of the electrons can be altered by the presence of a magnetic field.
- Consider cathode rays leaving the positive electrode through a small hole.
  - If they interact with a magnetic field perpendicular to an applied electric field, then the cathode rays can be deflected by different amounts.
  - The amount of deflection of the cathode rays depends on the applied magnetic and electric fields.
  - In turn, the amount of deflection also depends on the charge-to-mass ratio of the electron.
  - In 1897 Thomson determined the charge-to-mass ratio of an electron.
    - Charge-to-mass ratio: 1.76×10^8 C/g
    - C is a symbol for coulomb; it is the SI unit for electric charge.
- Millikan Oil Drop Experiment (1909)
  - Goal: find the charge on the electron to determine its mass.
  - Oil drops are sprayed above a positively charged plate containing a small hole.
  - As the oil drops fall through the hole, they acquire a negative charge.
  - Gravity forces the drops downward. The applied electric field forces the drops upward.
  - When a drop is perfectly balanced, the weight of the drop is equal to the electrostatic force of attraction between the drop and the positive plate.
  - Millikan carried out the above experiment and determined the charges on the oil drops to be multiples of 1.60×10^{-19} C.
  - He concluded the charge on the electron must be 1.60×10^{-19} C.
- Knowing the charge-to-mass ratio of the electron, we can calculate the mass of the electron:

$$\text{Mass} = \frac{1.60 \times 10^{-19}\,\text{C}}{1.76 \times 10^{8}\,\text{C/g}} = 9.10 \times 10^{-28}\,\text{g}$$

Radioactivity[6]

- **Radioactivity** is the spontaneous emission of radiation.
- Consider the following experiment:
  - A radioactive substance is placed in a lead shield containing a small hole so that a beam of radiation is emitted from the shield.
  - The radiation is passed between two electrically charged plates and detected.

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[2] Figure 2.4
[3] "A Millikan Oil Drop Analogy" from Further Readings
[4] "Marie Curie's Doctoral Thesis: Prelude to a Nobel Prize" from Further Readings
[5] Figure 2.5
[6] "The Curie-Becquerel Story" from Further Readings

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- Three spots are observed on the detector:
  1. a spot deflected in the direction of the positive plate,
  2. a spot that is not affected by the electric field, and
  3. a spot deflected in the direction of the negative plate.
- A large deflection toward the positive plate corresponds to radiation that is negatively charged and of low mass. This is called β-radiation (consists of electrons).
  - No deflection corresponds to neutral radiation. This is called γ-radiation (similar to X-rays).
- A small deflection toward the negatively charged plate corresponds to high-mass, positively charged radiation. This is called α-radiation (positively charged core of a helium atom.)
  - X-rays and γ–radiation are examples of true electromagnetic radiation, whereas α- and β-radiation are actually streams of particles—helium nuclei and electrons, respectively.

The Nuclear Model of the Atom[7,8,9,10]

- The plum pudding model is an early picture of the atom.
- The Thomson *plum-pudding model* pictures the atom as a sphere with small electrons embedded in a positively charged mass.
- Rutherford carried out the following "gold foil" experiment:
  - A source of α-particles was placed at the mouth of a circular detector.
  - The α-particles were shot through a piece of gold foil.
  - Both the gold nucleus and the α-particle were positively charged, so they repelled each other.
  - Most of the α-particles went straight through the foil without deflection.
  - If the Thomson model of the atom was correct, then Rutherford's result was impossible.
- Rutherford modified Thomson's model, postulating the **nuclear** model as follows:
  - Assume that the atom is spherical, but the positive charge must be located at the center with a diffuse negative charge surrounding it.
  - In order for the majority of α-particles that pass through a piece of foil to be undeflected, the majority of the atom must consist of a low-mass, diffuse negative charge—the electron.
  - To account for the small number of large deflections of the α-particles, the center or **nucleus** of the atom must consist of a dense positive charge.
  - Later experiments led to the discovery of **protons** (positive particles) and **neutrons** (neutral particles) in the nucleus.

FORWARD REFERENCES

- Radioactive decay will be further discussed in Chapter 14 as an example of first-order kinetics.
- Radioactivity will be further discussed in Chapter 21.

2.3 The Modern View of Atomic Structure[11,12,13]

- The atom consists of positive, negative, and neutral entities (**protons**, **electrons** and **neutrons**).
- Protons and neutrons are located in the nucleus of the atom, which is small. Most of the mass of the atom is due to the nucleus.
- Electrons are located outside the nucleus. Most of the volume of the atom is due to electrons.

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[7] Figure 2.7
[8] "Bowling Balls and Beads: A Concrete Analogy to the Rutherford Experiment" from Further Readings
[9] Figure 2.9
[10] "Turning Plastic into Gold" from Live Demonstrations
[11] "The Discovery of the Electron, Proton, and Neutron" from Further Readings
[12] Figure 2.10
[13] Interactive Sample Exercise 2.1

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- The quantity 1.602×10^{-19} C is called the **electronic charge**.
  - The charge on an electron is -1.602×10^{-19} C; the charge on a proton is $+1.602 \times 10^{-19}$ C; neutrons are uncharged.
  - Atoms have an equal number of protons and electrons, so they have no net electric charge.
- The **angstrom** is a convenient non-SI unit of length used to denote atomic dimensions.
  - Since most atoms have radii around 1×10^{-10} m, we define 1 Å = 1×10^{-10} m.
- Masses are so small that we define the **atomic mass unit**, amu.
  - 1 amu = 1.66054×10^{-24} g
  - The mass of a proton is 1.0073 amu, that of a neutron is 1.0087 amu, and that of an electron is 5.486×10^{-4} amu.

Atomic Numbers, Mass Numbers, and Isotopes[14,15,16,17]

- **Atomic number** (Z) = number of protons in the nucleus.
- **Mass number** (A) = total number of nucleons in the nucleus (i.e., protons and neutrons).
- By convention, for element X, we write ${}^{A}_{Z}\mathrm{X}$.
  - Thus, isotopes have the same Z but different A.
    - There can be a variable number of neutrons for the same number of protons. Isotopes have the same number of protons but different numbers of neutrons.
- All atoms of a specific element have the same number of protons.
  - **Isotopes** of a specific element differ in the number of neutrons.

FORWARD REFERENCES

- The concept of an isotope (specifically ^{12}C) will be useful when defining the mole in Chapter 3.
- Atomic numbers will be mentioned in the context of deriving electron configurations (Chapter 6), drawing Lewis structures (Chapter 8), and understanding molecular orbitals (Chapter 9).
- Atomic structure ideas will be applied to the understanding of nuclear reactions in Chapter 21.

2.4 Atomic Weights

The Atomic Mass Scale[18,19]

- Consider 100 g of water:
  - Upon decomposition, 11.1 g of hydrogen and 88.9 g of oxygen are produced.
  - The mass ratio of O to H in water is 88.9/11.1 = 8.
  - Therefore, the mass of O is $2 \times 8 = 16$ times the mass of H.
  - If H has a mass of 1, then O has a *relative mass* of 16.
  - We can measure atomic masses using a mass spectrometer.
  - We know that ^{1}H has a mass of 1.6735×10^{-24} g and ^{16}O has a mass of 2.6560×10^{-23} g.
  - *Atomic mass units* (amu) are convenient units to use when dealing with extremely small masses of individual atoms.
- 1 amu = 1.66054×10^{-24} g and 1 g = 6.02214×10^{23} amu
- By definition, the mass of ^{12}C is exactly 12 amu.

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[14] "The Origin of Isotope Symbolism" from Further Readings
[15] "Isotope Separation" from Further Readings
[16] "Dramatizing Isotopes: Deuterated Ice Cubes Sink" from Live Demonstrations
[17] Interactive Sample Exercise 2.3
[18] "Revisiting Molar Mass, Atomic Mass, and Mass Number: Organizing, Integrating, and Sequencing Fundamental Chemical Concepts" from Further Readings
[19] "Relative Atomic Mass and the Mole: A Concrete Analogy to Help Students Understand These Abstract Concepts" from Further Readings

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Average Weight[20,21,22]

- We average the masses of isotopes to give average atomic masses.
- Naturally occurring C consists of 98.93% ^{12}C (12 amu) and 1.07% ^{13}C (13.00335 amu).
- The average mass of C is as follows:
  - (0.9893)(12 amu) + (0.0107)(13.00335 amu) = 12.01 amu
- **Atomic weight** (AW) is also known as *average atomic mass*.
- *Atomic weights* are listed on the periodic table.

The Mass Spectrometer[23,24]

- A **mass spectrometer** is an instrument that allows for direct and accurate determination of atomic (and molecular) weights.
- The sample is charged as soon as it enters the spectrometer.
- The charged sample is accelerated using an applied voltage.
- The ions are then passed into an evacuated tube and through a magnetic field.
- The magnetic field causes the ions to be deflected by different amounts, depending on their mass.
- The ions are then detected.
  - A graph of signal intensity vs. mass of the ion is called a *mass spectrum*.

FORWARD REFERENCES

- Being able to locate atomic weights on the periodic table will be crucial in calculating molar masses in Chapter 3 and beyond.

2.5 The Periodic Table[25,26,27,28,29,30,31,32,33,34,35,36,37,38]

- The **periodic table** is used to organize the elements in a meaningful way.
- As a consequence of this organization, there are *periodic* properties associated with the periodic table.
- Rows in the periodic table are called **periods**.

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[20] "Using Monetary Analogies to Teach Average Atomic Mass" from Further Readings
[21] "Pictorial Analogies IV: Relative Atomic Weights" from Further Readings
[22] Interactive Sample Exercise 2.4
[23] "Mass Spectrometry for the Masses" from Further Readings
[24] Figure 2.11
[25] "Periodic Tables of Elemental Abundance" from Further Readings
[26] Figure 2.14
[27] "Seventh Row of the Periodic Table Is Now Complete with Addition of Four Elements" from Further Readings
[28] "A Second Note on the Term 'Chalcogen'" from Further Readings
[29] "Developing and Playing Chemistry Games to Learn about Elements, Compounds, and the Periodic Table: Elemental Periodica, Compoundica, and Groupica" from Further Readings
[30] "A Game-Based Approach to Learning the Idea of Chemical Elements and Their Periodic Classification" from Further Readings
[31] "An Effective Method of Introducing the Periodic Table as a Crossword Puzzle at the High School Level" from Further Readings
[32] "The Proper Place for Hydrogen in the Periodic Table" from Further Readings
[33] "Which Elements Are Metalloids?" from Further Readings
[34] "The Periodic Table: Key to Past 'Elemental' Discoveries—A New Role in the Future?" from Further Readings
[35] "An Educational Card Game for Learning Families of Chemical Elements" from Further Readings
[36] "Periodic Graphics: The Compositions of U.S. Coins" from Further Readings
[37] "Cheminoes: A Didactic Game to Learn Chemical Relationships between Valence, Atomic Number, and Symbol" from Further Readings
[38] Interactive Sample Exercise 2.5

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- Columns in the periodic table are called **groups**.
  - Several numbering conventions are used (i.e., groups may be numbered from 1 to 18, or from 1A to 8A and 1B to 8B).
- Some of the groups in the periodic table are given special names.
  - These names indicate the similarities between group members.
  - Examples:
    - Group 1A: alkali metals
    - Group 2A: alkaline earth metals
    - Group 7A: halogens
    - Group 8A: noble gases
- **Metallic elements**, or **metals**, are located on the left-hand side of the periodic table (most of the elements are metals).
  - Metals tend to be malleable, ductile, and lustrous and are good thermal and electrical conductors.
- **Nonmetallic elements**, or **nonmetals**, are located on the top right-hand side of the periodic table.
  - Nonmetals tend to be brittle as solids, tend to be dull in appearance, and do not conduct heat or electricity well.
- Elements with some properties similar to those of metals and some properties similar to those of nonmetals are called **metalloids** and are located at the interface between the metals and the nonmetals.
  - These include the elements B, Si, Ge, As, Sb and Te.

FORWARD REFERENCES

- Additional information associated with the unique location of an element in the periodic table will be covered in Chapter 6 (electron configurations), Chapter 7 (periodic properties), Chapter 8 (tendency to form ionic or covalent bonds) and Chapter 16 (relative acid strength).

2.6 Molecules and Molecular Compounds

- A **molecule** consists of two or more atoms bound tightly together.

Molecules and Chemical Formulas

- Each molecule has a **chemical formula**.
- The chemical formula indicates
  1. which atoms are found in the molecule, and
  2. in what proportion they are found.
- A molecule made up of two atoms is called a **diatomic molecule**.
  - Different forms of an element, which have different chemical formulas, are known as allotropes.
  - Allotropes differ in their chemical and physical properties.
    - Examples: ozone (O_3) and "normal" oxygen (O_2)
- Compounds composed of molecules are **molecular compounds**.
  - These contain at least two types of atoms.
  - Most molecular substances contain only nonmetals.

Molecular and Empirical Formulas

- **Molecular formulas**
  - These formulas give the actual numbers and types of atoms in a molecule.
  - Examples: H_2O, CO_2, CO, CH_4, H_2O_2, O_2, O_3, and C_2H_4
- **Empirical formulas**
  - These formulas give the relative numbers and types of atoms in a molecule (they give the lowest whole-number ratio of atoms in a molecule).
  - Examples: H_2O, CO_2, CO, CH_4, HO, CH_2

Picturing Molecules

- Molecules occupy three-dimensional space.
- However, we often represent them in two dimensions.

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- The **structural formula** gives the connectivity between individual atoms in the molecule.
- The structural formula may or may not be used to show the three-dimensional shape of the molecule.
- If the structural formula does show the shape of the molecule, then either a perspective drawing, a ball-and-stick model, or a space-filling model is used.
  - **Perspective drawings** use dashed lines and wedges to represent bonds receding and emerging from the plane of the paper.
  - **Ball-and-stick models** show atoms as contracted spheres and the bonds as sticks.
    - The angles in the ball-and-stick model are accurate.
  - **Space-filling models** give an accurate representation of the 3-D shape of the molecule.

FORWARD REFERENCES

- More detailed discussion of bonding in molecules and molecular shapes can be found in Chapters 8 and 9, respectively.

2.7 Ions and Ionic Compounds

- If electrons are added to or removed from a neutral atom, an **ion** is formed.
- When an atom or molecule loses electrons, it becomes positively charged.
  - Positively charged ions are called **cations**.
- When an atom or molecule gains electrons, it becomes negatively charged.
  - Negatively charged ions are called **anions**.
- In general, metal atoms tend to lose electrons, and nonmetal atoms tend to gain electrons.
- When molecules lose electrons, **polyatomic ions** are formed (e.g., SO_4^{2-}, NH_4^+).

Predicting Ionic Charges[39]

- An atom or molecule can lose more than one electron.
- Many atoms gain or lose enough electrons to have the same number of electrons as the nearest noble gas (group 8A).
- The number of electrons an atom loses is related to its position on the periodic table.
- Anions can also be viewed as particles originating from acids, and thus as having negative charges equal to the number of (acidic) hydrogen atoms in molecules of those acids (e.g., HNO_3 has 1 H atom, so NO_3^- has a charge of -1).

Ionic Compounds[40,41]

- A great deal of chemistry involves the transfer of electrons between species.
- Example:
  - To form NaCl, the neutral sodium atom, Na, must lose an electron to become a cation: Na^+.
  - The electron cannot be lost entirely, so it is transferred to a chlorine atom, Cl, which then becomes an anion: Cl^-.
  - The Na^+ and Cl^- ions are attracted to form an ionic NaCl lattice, which crystallizes.
- NaCl is an example of an **ionic compound** consisting of positively charged cations and negatively charged anions.
  - Important: note that there are no easily identified NaCl molecules in the ionic lattice. Therefore, we cannot use molecular formulas to describe ionic substances.
- In general, ionic compounds are combinations of metals and nonmetals, whereas molecular compounds are composed of nonmetals only.
  - There are exceptions; notably, $(NH_4)_2SO_4$ and other ammonium salts are ionic.

[39] Figure 2.19
[40] Figure 2.20
[41] Interactive Sample Exercise 2.9

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- Writing empirical formulas for ionic compounds:
  - You need to know the ions of which the compound is composed.
  - The formula must reflect the electrical neutrality of the compound.
  - You must combine cations and anions in a ratio in such a way that the total positive charge is equal to the total negative charge.
  - The empirical formula should be the smallest possible whole-number ratio of the two elements.
  - Example: Consider the formation of Mg_3N_2:
    - Mg loses two electrons to become Mg^{2+}.
    - Nitrogen gains three electrons to become N^{3-}.
    - For a neutral species, the number of electrons lost and the number gained must be equal.
    - However, Mg can lose electrons only in twos, and N can accept electrons only in threes.
    - Therefore, Mg needs to lose six electrons (2×3) and N gains those six electrons (3×2).
    - That is, 3 Mg atoms need to form 3 Mg^{2+} ions (total 3×2 positive charges), and 2 N atoms need to form 2 N^{3-} ions (total 2×3 negative charges).
    - Therefore, the formula is Mg_3N_2.

Chemistry and Life: Elements Required by Living Organisms[42]

- Of the known elements, only about 29 are required for life.
- Water accounts for at least 70% of the mass of most cells.
- More than 97% of the mass of most organisms comprises just six elements (O, C, H, N, P, and S).
- Carbon is the most common element in the solid components of cells.
- The most important elements for life are H, C, N, O, P, and S (red).
- The next most important ions are Na^+, Mg^{2+}, K^+, Ca^{2+}, and Cl^- (blue).
- The other required 18 elements are needed only in trace amounts (green); they are *trace elements*.

FORWARD REFERENCES

- Formulas (including correct charges) of ions will be important in writing metathesis and net ionic equations in Chapter 4 (sections 4.2–4.3).
- Periodic trends in ionization energy (in gas phase) as well as ionic radii (in crystals) will be covered in Chapter 7.
- The nature of bonding between ions and charges of most monoatomic ions will be rationalized in terms of electron configurations in Chapter 8 (section 8.2).
- Common types of ionic structures will be discussed in Chapter 11.
- Solubility of ionic solids will be covered qualitatively in Chapter 4 (section 4.2) and quantitatively in Chapter 17 (section 17.4).
- The fate of ionic solids when dissolved in water will be briefly discussed in Chapter 4 (section 4.1) and elaborated on in Chapter 13 (section 13.1); ion-dipole forces will be explained in Chapter 11 (section 11.2).
- The loss of electrons to form monoatomic metal cations (oxidation) and the gain of electrons to form monoatomic nonmetal anions (reduction) will be further discussed in Chapter 4 (section 4.4).
- Atoms of the same element appearing in several different ions (as well as molecules), and hence having different oxidation numbers, will be the basis of redox reactions in Chapter 20.
- The role of metal cations in the formation of metal complexes will be discussed in Chapter 23.

[42] Figure 2.21

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2.8 Naming Inorganic Compounds[43,44,45,46,47,48,49]

- **Chemical nomenclature** is the naming of substances.
- Common names are traditional names for substances (e.g., water, ammonia).
- Systematic names are based on a systematic set of rules.
  - Divided into organic compounds (those containing C, usually in combination with H, O, N, or S) and inorganic compounds (all other compounds).

Names and Formulas of Ionic Compounds[50]

1. Positive Ions (Cations)

- Cations formed from a metal have the same name as the metal.
  - Example: Na^+ = sodium ion.
  - Ions formed from a single atom are called *monatomic ions*.
- Many transition metals exhibit variable charge.
  - If the metal can form more than one cation, then the charge is indicated in parentheses in the name.
    - Examples: Cu^+ = copper(I) ion; Cu^{2+} = copper(II) ion.
  - An alternative nomenclature method uses the endings **-ous** and **-ic** to represent the lower and higher charged ions, respectively.
    - Examples: Cu^+ = cuprous ion; Cu^{2+} = cupric ion.
- Cations formed from nonmetals end in **-ium**.
  - Examples: NH_4^+ = ammonium ion; H_3O^+ = hydronium ion.

2. Negative Ions (Anions)[51,52]

- Monatomic anions (with only one atom) use the ending **-ide**.
  - Example: Cl^- is the chloride ion.
- Some polyatomic anions also use the -ide ending.
  - Examples: hydroxide, cyanide, and peroxide ions.
- Polyatomic anions (with many atoms) containing oxygen are called **oxyanions**.
  - Their names end in **-ate** or **-ite**. (The one with more oxygen ends in **-ate**.)
  - Examples: NO_3^- is nitr**ate**; NO_2^- is nitr**ite**.
- Polyatomic anions containing oxygen with more than two members in the series are named as follows (in order of **decreasing oxygen**):

per-…**-ate**	example:	ClO_4^-	**per**chlor**ate**
-ate		ClO_3^-	chlor**ate**
-ite		ClO_2^-	chlori**te**
hypo-...**-ite**		ClO^-	**hypo**chlori**te**

- Polyatomic anions containing oxygen with additional hydrogens are named by adding hydrogen or bicarbonate (one H), dihydrogen (two H) etc., to the name as follows:
  - CO_3^{2-} is the carbon**ate** anion.
  - HCO_3^- is the hydrogen carbonate (or **bi**carbonate) anion.

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[43] "Teaching Inorganic Nomenclature: A Systematic Approach" from Further Readings
[44] "Nomenclature Made Practical; Student Discovery of the Nomenclature Rules" from Further Readings
[45] "Using Product Content Labels to Engage Students in Learning Chemical Nomenclature" from Further Readings
[46] "Chemical Alias: An Engaging Way to Examine Nomenclature" from Further Readings
[47] "Flow Chart for Naming Inorganic Compounds" from Further Readings
[48] "Using Games to Teach Chemistry: An Annotated Bibliography" from Further Readings
[49] "ChemOkey: A Game to Reinforce Nomenclature" from Further Readings
[50] "The Proper Writing of Ionic Charges" from Further Readings
[51] "A Mnemonic for Oxy-Anions" from Further Readings
[52] Figure 2.23

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- PO_4^{3-} is the phosphate ion.
- $H_2PO_4^-$ is the **dihydrogen** phosphate anion.

3. Ionic Compounds

- These are named by the cation and then the anion.
- Examples:
  - $CaCl_2$ = calcium chloride
  - $(NH_4)_3PO_4$ = ammonium phosphate
  - $KClO_4$ = potassium perchlorate

Names and Formulas of Acids[53]

- Acids are substances that yield hydrogen ions when dissolved in water (Arrhenius definition).
  - The names of acids are related to the names of anions:
  - **-ide** becomes **hydro-**…**-ic** acid; example: HCl **hydro**chlor**ic** acid
  - **-ate** becomes **-ic** acid; $HClO_4$ perchlor**ic** acid
  - **-ite** becomes **-ous** acid. HClO hypochlor**ous** acid

Names and Formulas of Binary Molecular Compounds

- *Binary* molecular compounds have two elements.
- The most metallic element (i.e., the one farthest to the left on the periodic table) is usually written first. The exception is NH_3.
- If both elements are in the same group, the lower one is written first.
- Greek prefixes are used to indicate the number of atoms (e.g., *mono-, di-*, *tri-*).
  - The prefix *mono-* is never used with the first element (i.e., carbon monoxide, CO).
- Examples:
  - Cl_2O is **di**chlorine *mon*oxide.
  - N_2O_4 is **di**nitrogen *tetr*oxide.
  - NF_3 is nitrogen *tri*fluoride.
  - P_4S_{10} is **tetra**phosphorus *deca*sulfide.

FORWARD REFERENCES

- Nomenclature will be required throughout the text.
- Acids will be mentioned again in Chapter 4 and further discussed in Chapters 16 and 17.

2.9 Some Simple Organic Compounds

- **Organic chemistry** is the study of carbon-containing compounds.
  - *Organic compounds* are those that contain carbon and hydrogen, often in combination with other elements.

Alkanes

- Compounds containing only carbon and hydrogen are called **hydrocarbons**.
- In **alkanes** each carbon atom is bonded to four other atoms.
- The names of alkanes end in *-ane*.
  - Examples: methane, ethane, propane, butane.

Some Derivatives of Alkanes

- When *functional groups*, specific groups of atoms, are used to replace hydrogen atoms on alkanes, new classes of organic compounds are obtained.
  - **Alcohols** are obtained by replacing a hydrogen atom of an alkane with an –OH group.
  - Alcohol names derive from the name of the alkane and have an *-ol* ending.
    - Examples: methane becomes methanol; ethane becomes ethanol.

[53] Figure 2.25

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- **Isomers** are compounds with the same molecular formula but different arrangements of atoms.
  - Examples of *structural isomers* are 1-propanol and 2-propanol; they differ with regard to which carbon has the OH functional group.
- Carbon atoms often form compounds with long chains of carbon atoms.
  - Properties of alkanes and derivatives change with changes in chain length.
  - *Polyethylene*, a material used to make many plastic products, is an alkane with thousands of carbons.

FORWARD REFERENCES

- Simple organic compounds will be used throughout the text to illustrate weak acid behavior (e.g., acetic acid in Chapters 16 and 17), weak base behavior (e.g., amines in Chapters 16 and 17), resonance (e.g., benzene in Chapter 9), molecular polarity (e.g., CH_3Cl vs. CCl_4 in Chapter 9), solubility of organic compounds in water or organic solvents (e.g., pentane in Chapter 13), to mention just a few.
- Nonpolar organic compounds will be mentioned again when London dispersion forces are discussed in Chapter 11.
- This section introduces organic chemistry, which will be elaborated on in Chapter 24.

Further Readings:

1. John J. Fortman, "Analogical Demonstration," *J. Chem. Educ*., Vol. 69, **1992**, 323–324. This reference includes demonstrations of the concepts of the conservation of mass in chemical reactions, the Law of Multiple Proportions, etc.

2. Doris Eckey, "A Millikan Oil Drop Analogy," *J. Chem. Educ*., Vol. 73, **1996**, 237–238.

3. Robert L. Wolke, "Marie Curie's Doctoral Thesis: Prelude to a Nobel Prize," *J. Chem. Educ*., Vol. 65, **1988**, 561–573.

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

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

6. Harold F. Walton, "The Curie-Becquerel Story," *J. Chem. Educ*., Vol. 69, **1992**, 10–15.

7. William Spindel and Takanobu Ishida, "Isotope Separation," *J. Chem. Educ*., Vol. 68, **1991**, 312–318. An article describing methods used to isolate important isotopes.

8. William B. Jensen, "The Origin of Isotope Symbolism," *J. Chem. Educ.*, Vol. 88, **2011**, 22–23.

9. Stephen DeMeo, "Revisiting Molar Mass, Atomic Mass, and Mass Number: Organizing, Integrating, and Sequencing Fundamental Chemical Concepts," *J. Chem. Educ.*, Vol. 83, **2006**, 617–620.

10. 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.

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

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Chemistry The Central Science 14th edition Brown Solutions Manual

12. John H. Fortman, "Pictorial Analogies IV: Relative Atomic Weights," *J. Chem. Educ*., Vol. 70, **1993**, 235–236.

13. Jared D. Persinger, Geoffrey C. Hoops, and Michael J. Samide, "Mass Spectrometry for the Masses," *J. Chem. Educ*., Vol. 81, **2004**, 1169–1171.

14. Steven I. Dutch, "Periodic Tables of Elemental Abundance," *J. Chem. Educ.*, Vol. 76, **1999**, 356–358.

15. Jyllian Kemsley, "Seventh Row of the Periodic Table Is Now Complete with Addition of Four Elements," *Chem. Eng. News*, Vol. 94, **2016**, 6.

16. Werner Fischer, "A Second Note on the Term 'Chalcogen,'" *J. Chem. Educ*., Vol. 78, **2001**, 1333.

17. Antonio Joaquin Franco Mariscal, Jose Maria Olivia Martinez, and Serafin Bernal Marquez, "An Educational Card Game for Learning Families of Chemical Elements," *J. Chem. Educ*., Vol. 89, **2012**, 1044–1046.

18. Eylem Bayir, "Developing and Playing Chemistry Games to Learn about Elements, Compounds, and the Periodic Table: Elemental Periodica, Compoundica, and Groupica," *J. Chem. Educ*., Vol. 91, **2014**, 531–535.

19. Antonio Joaquin Franco-Mariscal, Jose Maria Oliva-Martinez, Angel Blanco-Lopez, and Enrique Espana-Ramos, "A Game-Based Approach to Learning the Idea of Chemical Elements and Their Periodic Classification," *J. Chem. Educ*., Vol. 93, **2016**, 1173–1190.

20. Sushama D. Joag, "An Effective Method of Introducing the Periodic Table as a Crossword Puzzle at the High School Level," *J. Chem. Educ*., Vol. 91, **2014**, 864–867.

21. Marshall W. Cronyn, "The Proper Place for Hydrogen in the Periodic Table," *J. Chem. Educ*., Vol. 80, **2003**, 947–950.

22. Rene E. Vernon, "Which Elements Are Metalloids?" *J. Chem. Educ*., Vol. 90, **2013**, 1703–1707.

23. Darleane C. Hoffman, "The Periodic Table: Key to Past 'Elemental' Discoveries—A New Role in the Future?" *J. Chem. Educ.,* Vol. 86, **2009**, 1122–1128.

24. Andy Brunning, "Periodic Graphics: The Compositions of U.S. Coins," *Chem. Eng. News*, Vol. 94, **2016**, 29.

25. Luis F. Moreno, Gina Hincapie, and Maria Victoria Alzate, "Cheminoes: A Didactic Game to Learn Chemical Relationships between Valence, Atomic Number, and Symbol," *J. Chem. Educ*., Vol. 91, **2014**, 872–875.

26. M. Araceli Calvo Pascual, "Using Product Content Labels to Engage Students in Learning Chemical Nomenclature," *J. Chem. Educ*., Vol. 91, **2014**, 757–759.

27. Mikhail Kurushkin and Maria Mikhaylenko, "Chemical Alias: An Engaging Way to Examine Nomenclature," *J. Chem. Educ*., Vol. 92, **2015**, 1678–1680.

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Chemistry The Central Science 14th edition Brown Solutions Manual

28. Gerhard Lind, "Teaching Inorganic Nomenclature: A Systematic Approach," *J. Chem. Educ*., Vol. 69, **1992**, 613–614.

29. Michael C. Wirtz, Joan Kaufmann, and Gary Hawley, "Nomenclature Made Practical: Student Discovery of the Nomenclature Rules," *J. Chem. Educ*., Vol. 83, **2006**, 595–598.

30. Nusret Kavak, "ChemOkey: A Game to Reinforce Nomenclature," *J. Chem. Educ*., Vol. 89, **2012**, 1047–1049.

31. David Robson, "Flow Chart for Naming Inorganic Compounds," *J. Chem. Educ*., Vol. 60, **1983**, 131–132.

32. Jeanne V. Russell, "Using Games to Teach Chemistry. An Annotated Bibliography," *J. Chem. Educ*., Vol. 76, **1999**, 481–484. This is the first article in a special issue that contains many articles describing games and puzzles that may be used to teach chemistry.

33. William B. Jensen, "The Proper Writing of Ionic Charges," *J. Chem. Educ*., Vol. 89, **2012**, 1084–1085.

34. Steven J. Hawkes, "A Mnemonic for Oxy-Anions," *J. Chem. Educ*., Vol. 67, **1990**, 149.

Live Demonstrations:

1. Arthur B. Ellis, Edward A. Adler, and Frederick H. Juergens, "Dramatizing Isotopes: Deuterated Ice Cubes Sink," *J. Chem. Educ*., Vol. 67, **1990**, 159–160. Differences in density of $H_2O(l)$ and $D_2O(s)$ are used to demonstrate the effects of isotopic substitution.

2. Robert B. Gregory and Ed Vitz, "Turning Plastic into Gold: An Analogy to Demonstrate the Rutherford Gold Foil Experiment," *J. Chem. Educ*., Vol. 84, **2007**, 626–628.

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