FUNDAMENTALS OF CHEMISTRY

David E. Goldberg

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BROOKLYN COLLEGE

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Preface

Chemistry is a dynamic, rapidly changing field that comprises an extraordinarily interesting subject to study and a challenging one to teach. Today, as perhaps never before, the world is looking to the field of chemistry to provide answers to some of the most challenging problems that confront each of us as citizens of the global community. We live in a world in which we are confronted with products, problems, and changes that our parents and grandparents never dreamed of. And we are learning that what happens on the other side of the world can have a profound impact on us and our lives. Problems such as global climate change, loss of stratospheric ozone, and human-created ecological disasters such as the 1991 oil well fires in Kuwait have become subjects of daily discussion in newspapers and on television. Each discipline of science, but especially chemistry, will contribute to solving these problems, as well as others not yet encountered. In addition, the field of chemistry will continue to yield new processes and materials that will open up whole new vistas of opportunities and benefits.

As citizens of the global community, we need a solid foundation in scientific principles—including chemical principles—to help us better understand the world around us, as well as to contribute positively to that world. It is my hope that students using this book will develop a foundation of chemical principles with which they can begin to understand the processes that make up the world and underlie life and will use that foundation to succeed and prosper, not only in any subsequent chemistry courses but in life as well.

Audience and Philosophy

This book is intended primarily to serve beginning chemistry students, who typically have had no instruction, or limited instruction, in chemistry. These students often need help not only with mathematical manipulations, but also with reading and writing scientific material with precision. The factor label method (dimensional analysis) is used to help the student translate word problems into easily solved algebraic expressions. In many places, a problem is stated in parts to lead the student stepwise through a solution; later the same problem appears as a whole, worded as it might appear on an examination. There is a series of figures that build on preceding ones to reveal the fundamental unity of the concept of the mole. Many problems are worded so as to show the student that very different questions may sound similar and that the same question may be presented in very different words; these will encourage the student to try to understand concepts and not to memorize solutions. When different terms that sound alike are presented and explained together, the student can more easily learn both (see Problems 5.1, 5.2, and 5.3, for example).

Frequent use of analogies to daily life helps students understand that chemistry problems are not significantly different from everyday problems, even though they may seem more difficult because of their unfamiliarity. For example, calculations involving dozens of pairs of socks and moles of diatomic molecules may be carried out by the same methods (Problems 7.3 and 7.4). Oxidizing and reducing agents can be compared conceptually to dish towels and wet dishes (Example 15.12). Specific heat calculations are like those involving room rates at a hotel (Example 13.9).

Modern nomenclature is used throughout the text (the Stock system for inorganic compounds and IUPAC nomenclature for organic compounds), but classical group numbering is used in the periodic table since this numbering is an aid to learning many elementary concepts (the number of electrons in an atom's outermost shell, for example).

Pedagogical Devices

The text includes a variety of pedagogical devices. These were chosen and designed to answer the question "If I were a student, what would help me organize and understand the material covered in this book?"

1. *Key Terms and Symbols*. Each chapter begins with a listing of the key terms and the symbols or abbreviations that are used in the chapter. Sectional references are provided. Students have the opportunity to use these lists to review their understanding of the important terms and symbols before examinations.
2. *Chapter Outline*. There is also an outline at the beginning of each chapter. The outline allows students to tell at a glance how the chapter is organized and what major topics are included.
3. *Learning Objective*. At the start of each section, a learning objective is presented to alert students to the key concept covered in the section. These objectives are another valuable study tool for students when they are reviewing chapter material for examinations.
4. *Boldfaced Key Terms*. Key terms appear in **boldface** when they are introduced within the text and are immediately defined in context. These terms are also listed, with sectional references, at the start of the chapter. All key terms are defined in the glossary.
5. *Items of Interest*. Throughout the text, boxes titled "Item of Interest" relate the subject matter to the real world.
6. *Marginal Comments*. Marginal comments are designed to alert the student to a key point, a helpful hint, or a safety caution.
7. *Tables*. Numerous strategically placed tables list and summarize important information, making it readily accessible for efficient study.
8. *Enrichment Boxes*. Throughout the text, boxes titled "Enrichment" highlight special topics that take the text material to a more extended level. Students will find them to be a lively and interesting feature as they investigate the processes of chemistry.

9. *Examples and Practice Problems:* The book has a wealth of Examples that show the student the step-by-step solution to the problem presented, which is directly related to the preceding textual information. The Practice Problems that follow the Examples give students the opportunity to solve a similar problem immediately. Solutions to Practice Problems are presented at the end of the book to provide immediate feedback.

10. *Photographs.* A wide array of visually appealing and informative photographs is used to help students understand chemical and physical phenomena and pique their interest.

11. *Illustrations.* Because a picture is worth a thousand words, each chapter is amply illustrated with accurate, colorful diagrams that clarify difficult concepts and enhance learning.

12. *Flow Diagrams.* To help students understand the steps in problem solving, flow diagrams have been included at key locations throughout the text. These diagrams allow students to visualize the process of solving a problem.

13. *Summary.* At the end of each chapter, there is a summary of the major concepts covered. Each section is reviewed in paragraph form. The summary, along with the chapter outline and section objectives, provides a complete overview of the chapter material.

14. *Items for Special Attention.* Appearing at the end of every chapter, this unique section highlights and emphasizes key concepts that often confuse students. This section anticipates students' questions and problem areas and helps them avoid many pitfalls.

15. *Self-Tutorial Problems.* This end-of-chapter section presents problems in simple form, many from everyday life, and emphasizes the importance of identifying the information needed to answer questions, thus encouraging the advancement of students' analytical skills.

16. *Problems.* The problems in this end-of-chapter set are grouped under headings that match the chapter's section titles. This organization allows students to practice the problem-solving skills and methods associated with each important concept presented in the text.

17. *General Problems.* The final set of problems in each chapter is more difficult than the others and is not classified by topic. Many of these problems require knowledge of two or more concepts. Similar in scope to the type of questions students will be confronted with on tests, these problems provide students with an excellent means by which to judge their knowledge of the chapter's contents.

18. *Appendixes.* The book contains a complete set of appendixes, which include the solutions to the in-text Practice Problems (Appendix 4) and the end-of-chapter problems (Appendix 5). A short review of scientific algebra and a unique presentation detailing the use of the electronic calculator (Appendix 1) will help students overcome any mathematical deficiencies. Lists of symbols, abbreviations, prefixes and suffixes, and mathematical equations (Appendixes 2 and 3) make the book more user-friendly.

19. *Glossary.* A complete glossary of all important terms is found at the end of the text.

The Learning System

Chapter Outline

Each chapter begins with an outline. These allow students to tell at a glance how the chapter is organized and what major topics are included.

14 Solutions

Outline

14.1 The Solution Process
14.2 Saturated, Unsaturated, and Supersaturated Solutions
14.3 Molality
14.4 Mole Fraction
14.5 Colligative Properties

Key Terms *(Key terms are defined in the Glossary.)*

boiling-point elevation (14.5)
colligative properties (14.5)
freezing-point depression (14.5)
ideal solution (14.5)
molal (14.3)
molality (14.3)
mole fraction (14.4)
nonvolatiie (14.5)
osmotic pressure (14.5)
Raoult's law (14.5)
saturated solution (14.2)
solubility (14.2)
supersaturated solution (14.2)
unsaturated solution (14.2)
vapor-pressure lowering (14.5)
volatile (14.5)

Symbols

k_b (boiling-point elevation constant) (14.5)
k_f (freezing-point depression constant) (14.5)
m (molal) (14.3)
m (molality) (14.3)
$P°$ (vapor pressure of pure substance) (14.5)
ΔP (vapor-pressure lowering) (14.5)
π (osmotic pressure) (14.5)
Δt_b (boiling-point elevation) (14.5)
Δt_f (freezing-point depression) (14.5)
X_A (mole fraction of component A) (14.4)

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Key Terms and Symbols

Each chapter begins with a listing of the key terms and symbols or abbreviations that are presented in the chapter. Section references are provided. Using these lists, students have the opportunity to review their understanding of important terms and symbols prior to examinations.

Learning Objective

At the start of each section, a learning objective is presented to alert students to the key concept covered in the section. The objectives are another valuable study tool when students review chapter material for examinations.

13.3 Changes of Phase

OBJECTIVE To learn about the processes involved when a substance changes from one state to another

Any process that results in a change of state (or phase) for a sample of matter is called a **phase change.** When a solid changes to a liquid as a result of a rise in temperature, the process is called **melting,** or **fusion.** The process of changing a liquid to a gas is called **evaporation,** or **vaporization.** A gas in contact with its liquid is often called a **vapor.** Changing a solid directly into a gas is called sublimation. Changing a liquid to a solid is **freezing,** and changing a gas to either a solid or a liquid is called **condensation.** These states and changes are summarized in Figure 13.8.

As an example of these phase changes, we will investigate the vaporization process in some detail. It takes energy to cause a liquid to evaporate. We can put a pan of water on a stove and add heat to it to cause rapid vaporization. When you come out of a swimming pool on a hot breezy day, you are aware that energy is used to evaporate a liquid. The breeze causes evaporation, which uses energy; as a result, your temperature is lowered. You grab for a towel to stop the evaporation and resulting cooling.

A process called **distillation** is often used to purify liquids. The liquid is heated to vaporize it, and the vapor is cooled to condense it back to a liquid in a different place (Figure 13.9). Impurities that are less easily vaporized are left behind in the original container, and those that are more easily vaporized are discarded before the desired product is collected.

On a molecular level, the change from a liquid to a gas is accompanied by a separation of the molecules from close proximity to far apart. Only the most energetic of the liquid molecules have sufficient energy to overcome the attractive forces and go into the gas phase. Thus, the process of vaporization leaves behind the less energetic molecules. As you learned in Section 12.8, the absolute temperature is directly proportional to the average kinetic energy of gas molecules. It turns out that the same thing is also true for liquid (and solid) molecules.

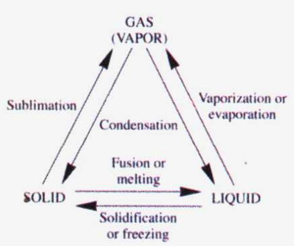


Figure 13.8 Terminology of Phase Changes

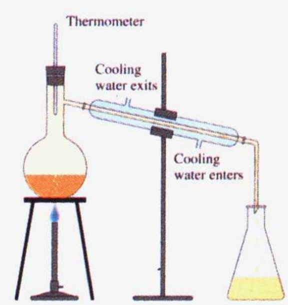


Figure 13.9 Distillation Apparatus

Bold-Faced Words

New terms appear in boldface when they are introduced within the text and are immediately defined in context. Key terms are introduced, with section references, at the start of the chapter. All key terms are defined in the Glossary.

Item of Interest

Throughout the text, Item of Interest boxes are included to relate the subject matter to the real world.

5.1 Chemical Formulas 91

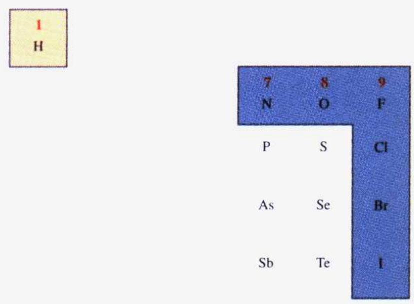


Figure 5.3
The Seven Elements That Form Diatomic Molecules
Note that the shape formed by six of these elements in the periodic table is like a seven starting at atomic number 7.

You must remember the seven elements that occur as **diatomic molecules** (molecules with two atoms) *when they are not combined with other elements*. Fortunately, these elements are easy to remember because, except for hydrogen, they form a shape like a seven in the periodic table starting at the element with atomic number 7 (Figure 5.3).

The collection of atoms represented by a formula is called a **formula unit.** A chemical formula consists of symbols of element(s), often with **subscripts** that tell how many atoms of each element are present per formula unit. The subscript *follows* the symbol of the element it multiplies. Parentheses may be

ITEM OF INTEREST

Hydrogen molecules (H_2) are so much more stable than separated hydrogen atoms that the reaction of the atoms to form molecules produces a lot of heat:

$$2\,H \rightarrow H_2 + \text{heat}$$

Production of a given number of H_2 molecules from hydrogen atoms produces more heat than the production of the same number of CO_2 molecules from burning carbon (charcoal) in oxygen. Construction workers take advantage of the reaction of atomic hydrogen to weld steel pieces together in the absence of oxygen. That condition is desirable because oxygen might make the steel rust. Where do the hydrogen atoms come from? They are produced by electrical discharge in a welding torch, such as is diagrammed in Figure 5.4.

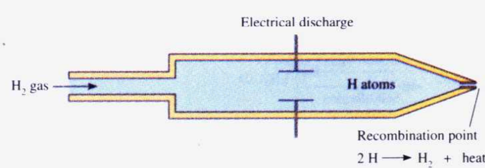


Figure 5.4
A Welding Torch
Hydrogen gas is piped into a special tube in which many of the H_2 molecules are separated into atoms by an electrical discharge. The gas flow is adjusted to the rate at which most of the H atoms recombine into molecules just as they exit from the torch. The heat produced at that point is intense enough to weld pieces of steel together.

Marginal Comments

Marginal comments are designed to alert the student to a key point, a helpful hint, or a safety precaution.

Practice Problem 5.2 How many atoms of each element are present in one formula unit of each of the following compounds? (a) $Co(NO_3)_2$ (b) $CoCO_3$

In reading formulas aloud, the number is simply stated for any subscript following a symbol, as in H_2O: "H two O." To express parentheses followed by a subscript 2, the words "taken twice" are used; for a subscript 3, the words "taken three times" are used, and so on. The centered dot is read "dot." Here are some examples to illustrate these conventions:

$(NH_4)_2SO_4$	"N H four taken twice S O four"
$(NH_4)_3PO_4$	"N H four taken three times P O four"
$CH_3(CH_2)_4CH_3$	"C H three, C H two taken four times, C H three"
$CuSO_4 \cdot 5H_2O$	"C u S O four dot five H two O"

5.2 Ionic Bonding

OBJECTIVE To learn about how the octet rule governs the formation of ionic compounds and how to predict the formulas of compounds of main group metals with nonmetals

The electrons in atoms are arranged in groups having nearly the same energies. These energy levels are often referred to as shells. The first shell of any atom can hold a maximum of two electrons; the second shell can hold a maximum of eight electrons; and the other shells can hold a maximum of eight electrons when they are the outermost shell, but a greater number when they are not (Table 5.1). The **outermost shell** is the shell with the highest value of n that contains electrons. The details of the electronic structures of atoms are covered in Chapter 4.

The noble gases are composed of stable atoms; no reactions of the first three (He, Ne, Ar) have been discovered, and the others (Kr, Xe, Rn) are almost completely unreactive. The stability of the noble gases is due to the eight electrons in the outermost shell of each atom (two electrons in the case of helium). In fact, eight electrons in the outermost shell is a stable configuration for most atoms. Atoms other than those of the noble gases tend to form ionic and/or covalent bonds with other atoms to achieve this electronic configuration. The eight electrons in the outermost shell are called an **octet.** The tendency of atoms to be stable with eight electrons in the outermost shell is called the **octet rule.** In some compounds, not all the atoms obey the octet rule. Some exceptions to the octet rule will be mentioned in Section 5.4.

Atoms tend to accept, donate, or share electrons to achieve the electronic structure of the nearest noble gas.

TABLE 5.1 Maximum Electron Occupancy of Shells

Shell Number	Maximum Occupancy as the Outermost Shell	Maximum Occupancy as an Inner Shell
1	2	2
2	8	8
3	8	18
4	8	32
5	8	50*
6	8	72*
7	8*	98*

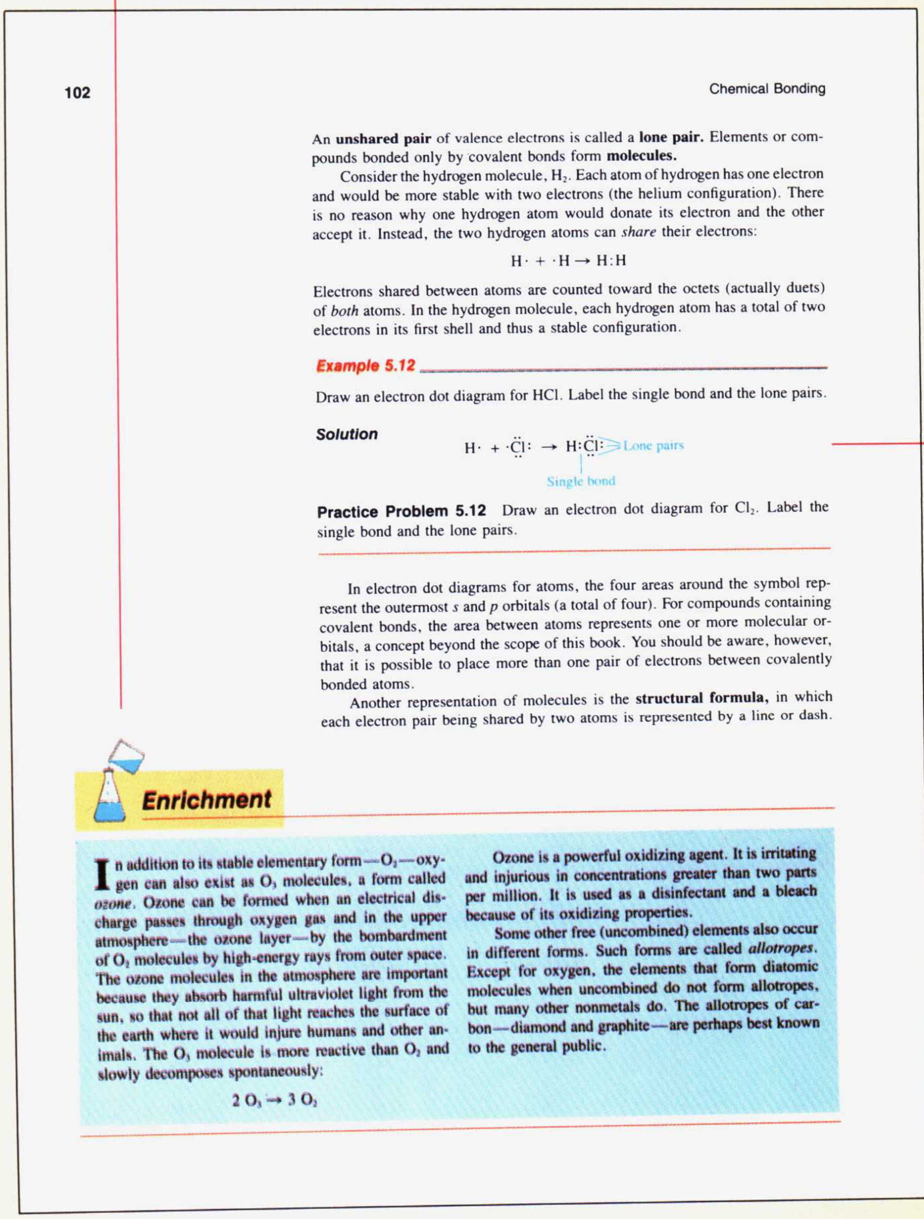
*More than the number of electrons available in any atom

Tables

Numerous strategically placed tables list and summarize important information, making it readily accessible for efficient study.

Enrichment Boxes

Throughout the text, Enrichment boxes discuss special topics at a more extended level. Students will find them to be a lively and interesting feature as they investigate the processes of chemistry.



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An **unshared pair** of valence electrons is called a **lone pair.** Elements or compounds bonded only by covalent bonds form **molecules.**

Consider the hydrogen molecule, H_2. Each atom of hydrogen has one electron and would be more stable with two electrons (the helium configuration). There is no reason why one hydrogen atom would donate its electron and the other accept it. Instead, the two hydrogen atoms can *share* their electrons:

$$H\cdot + \cdot H \rightarrow H:H$$

Electrons shared between atoms are counted toward the octets (actually duets) of *both* atoms. In the hydrogen molecule, each hydrogen atom has a total of two electrons in its first shell and thus a stable configuration.

Example 5.12

Draw an electron dot diagram for HCl. Label the single bond and the lone pairs.

Solution

Practice Problem 5.12 Draw an electron dot diagram for Cl_2. Label the single bond and the lone pairs.

In electron dot diagrams for atoms, the four areas around the symbol represent the outermost *s* and *p* orbitals (a total of four). For compounds containing covalent bonds, the area between atoms represents one or more molecular orbitals, a concept beyond the scope of this book. You should be aware, however, that it is possible to place more than one pair of electrons between covalently bonded atoms.

Another representation of molecules is the **structural formula,** in which each electron pair being shared by two atoms is represented by a line or dash.

Enrichment

In addition to its stable elementary form—O_2—oxygen can also exist as O_3, molecules, a form called *ozone*. Ozone can be formed when an electrical discharge passes through oxygen gas and in the upper atmosphere—the ozone layer—by the bombardment of O_2 molecules by high-energy rays from outer space. The ozone molecules in the atmosphere are important because they absorb harmful ultraviolet light from the sun, so that not all of that light reaches the surface of the earth where it would injure humans and other animals. The O_3 molecule is more reactive than O_2 and slowly decomposes spontaneously:

$$2\ O_3 \rightarrow 3\ O_2$$

Ozone is a powerful oxidizing agent. It is irritating and injurious in concentrations greater than two parts per million. It is used as a disinfectant and a bleach because of its oxidizing properties.

Some other free (uncombined) elements also occur in different forms. Such forms are called *allotropes*. Except for oxygen, the elements that form diatomic molecules when uncombined do not form allotropes, but many other nonmetals do. The allotropes of carbon—diamond and graphite—are perhaps best known to the general public.

Examples and Practice Problems

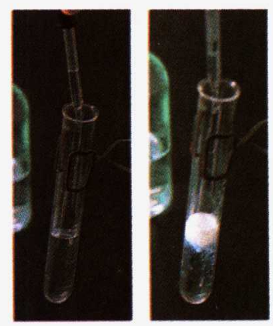
A wealth of examples throughout the text provides step-by-step solutions to problems. The examples show the student how the textual information presented is used to solve problems. Practice Problems follow most examples and give students the opportunity to solve similar problems. Solutions are provided at the end of the book to allow for immediate feedback.

Photographs

A wide array of visually appealing and informative photographs helps students understand chemical and physical phenomena.

10.1 Writing Net Ionic Equations 201

Figure 10.2
The Reaction of Silver Nitrate and Sodium Chloride



Since the Na^+ and NO_3^- ions appear on both sides of this equation (are unchanged by the reaction), they are called **spectator ions.** They may be eliminated from the equation:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

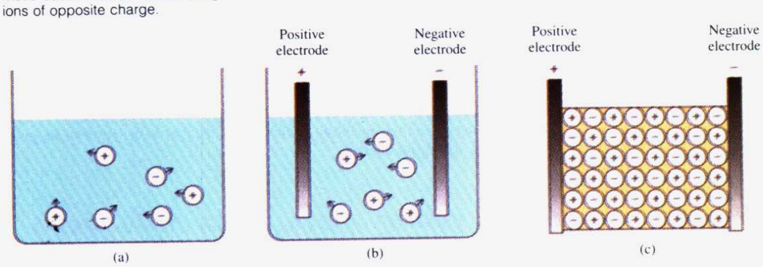
This equation is an example of a **net ionic equation.** All the spectator ions are omitted from a net ionic equation.

Thus, to write net ionic equations, we first write compounds that are both soluble and ionic in the form of their separate ions. A listing of water-soluble compounds was given in Table 8.3. In addition to the compounds listed there, all strong acids are water soluble. Compounds that dissociate or ionize in aqueous solution include the following:

1. Strong acids (HCl, $HClO_3$, $HClO_4$, HBr, HI, HNO_3, H_2SO_4)
2. All salts (compounds containing metal or ammonium ions)
3. All soluble metal hydroxides

All other compounds (for example, gases, other covalent compounds, and all solids) are written as compounds in an ionic equation. We then eliminate the ions that appear unchanged on the two sides of the equation to obtain the net ionic equation.

Figure 10.3
Mobility of Ions in Solution
(a) Ions in dilute solutions are free to move independently of other ions. In the absence of electrodes, they move in random directions. (b) Under the influence of the charges on electrodes, the ions move toward the electrode of opposite charge. (c) In contrast, even if charged electrodes are present, ions in solids cannot move because of the surrounding ions of opposite charge.



Illustrations

Because a picture is worth a thousand words, each chapter is amply illustrated with accurate, colorful diagrams that clarify difficult concepts and enhance learning.

Flow Diagrams

To help students understand the steps in problem solving, flow diagrams have been used throughout the text. These diagrams will help students visualize the process of solving a problem.

Example 7.5

Knowing that the molecular mass (formula mass) of CO_2 is 44.0 amu, show that CO_2 has a molar mass equal to 44.0 g/mol.

Solution

$$\frac{44.0\text{ amu}}{1\text{ molecule}}\underbrace{\left(\frac{6.02\times10^{23}\text{ molecules}}{1\text{ mol}}\right)}_{\text{Avogadro's number}}\underbrace{\left(\frac{1.00\text{ g}}{6.02\times10^{23}\text{ amu}}\right)}_{\text{From Example 7.4}}=\frac{44.0\text{ g}}{1\text{ mol}}$$

Practice Problem 7.5 The formula mass of $HClO_4$ is 100.5 amu. What is its molar mass (in grams per mole)?

We can use Avogadro's *number* as a conversion factor to convert moles to *numbers* of formula units, and vice versa. We can use the molar *mass* to convert moles to *masses*, and vice versa. (See Figure 7.1.)

Example 7.6

(a) Calculate the number of molecules in 2.00 mol of C_2H_4.

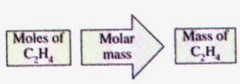
(b) Calculate the mass of 2.00 mol of C_2H_4.

Solution

(a) $$2.00\text{ mol }C_2H_4\left(\frac{6.02\times10^{23}\text{ molecules }C_2H_4}{1\text{ mol }C_2H_4}\right)$$

$$=12.0\times10^{23}\text{ molecules }C_2H_4=1.20\times10^{24}\text{ molecules }C_2H_4$$

(b) The molar mass of C_2H_4 is 2(12.0 g) + 4(1.008 g) = 28.0 g.



$$2.00\text{ mol }C_2H_4\left(\frac{28.0\text{ g }C_2H_4}{1\text{ mol }C_2H_4}\right)=56.0\text{ g }C_2H_4$$

Practice Problem 7.6

(a) Calculate the number of moles of C_3H_8 in 17.0 g of C_3H_8.

(b) Calculate the number of moles of C_3H_8 in a sample containing 6.95×10^{24} C_3H_8 molecules.

Example 7.7

Calculate the number of molecules of CH_4 in 23.45 g of CH_4.

Solution

The molar mass of CH_4 is 12.01 g + 4(1.008 g) = 16.04 g. (The molar mass is calculated to four significant digits because the data in the problem are given to four significant digits.)

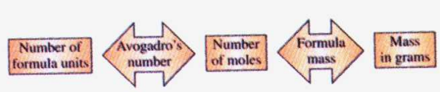


Figure 7.1
Some Conversions Involving Moles

Summary

At the end of each chapter, there is a summary of the major concepts covered. Major topics are reviewed in paragraph form. The summary, along with the chapter outline and section objectives, provides a complete overview of the chapter material.

Summary

The Brønsted theory of acids and bases extends the definition of acids and bases, which allows an explanation of why most salts dissolved in water do not form neutral solutions. Acids are defined as proton donors, and bases are proton acceptors. An excess of H_3O^+ ions over OH^- ions makes an aqueous solution acidic, and an excess of OH^- ions over H_3O^+ ions makes the solution basic. Neutral solutions have equal concentrations of these two ions. Some substances, most notably water and also acid salts such as $NaHCO_3$, can act as either acids or bases, depending on what other species is present. According to the Brønsted theory, an acid reacts with a base to produce a conjugate base of the original acid and a conjugate acid of the original base. The stronger the acid, the weaker is its conjugate base. Strong acids, for example, have feeble conjugate bases. Solutions of salts in water often test acidic or basic because one of their ions reacts with the water more than the other does. (Section 17.1)

For reactions of weak acids or weak bases with water, the specialized equilibrium constant is denoted K_a or K_b, respectively. Neither K_a nor K_b includes the concentration of water in its equilibrium constant expression. If you are given initial concentrations and one equilibrium concentration, you can calculate the equilibrium constant. If you are given initial concentrations and the value of the equilibrium constant, you can calculate the equilibrium concentrations. (Section 17.2)

Water can react with itself to a very limited extent, and in every aqueous solution there is at least *some* concentration of hydronium ion and of hydroxide ion, enough to satisfy the equation

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Given a hydronium ion concentration, you can calculate the hydroxide ion concentration using this equation, and vice versa. The pH, defined as $-\log [H_3O^+]$, is a convenient way of expressing the acidity of a solution. (Section 17.3)

A buffer solution resists changes in its pH due to addition of small quantities of strong acid or strong base. A buffer solution may be formed by combining a weak acid and its conjugate base or a weak base and its conjugate acid. The resistance of the buffer solution to pH changes is based on LeChâtelier's principle. For example, if a small quantity of H_3O^+ ions is added to a solution containing the conjugate base of a weak acid, the base will react with most of the added H_3O^+ ions to form more conjugate weak acid, and the pH will not change much. If a small quantity of OH^- ions is added to a solution containing an unionized weak acid, it will react with the acid producing more conjugate base. The added OH^- ions are no longer present to increase the pH. Given a quantity of strong acid or base that is added to a buffer solution, you first calculate how much of a reactant that is present will react completely with the added quantity, giving consideration to the limiting quantity (Chapter 9). You next deduce the new concentrations of the weak acid and its conjugate base and then proceed with the equilibrium calculation. (Section 17.4)

Items for Special Attention

Appearing at the end of every chapter, this unique section highlights and emphasizes key concepts that often confuse students. This section anticipates students' questions and problem areas.

Self-Tutorial Problems

This section presents simple problems, many from everyday life, and emphasizes the importance of identifying the information needed to answer questions, thus encouraging the development of students' analytical skills. At the end of the book are complete solutions to all the problems in the text.

226 Molarity

ITEMS FOR SPECIAL ATTENTION

- Be sure to use mol as an abbreviation for mole—*not* M or m, which are used for other quantities related to moles. Otherwise, you may mix up quantities and concentrations.
- Remember that volume of *solution*, not volume of solvent, is used in the definition of molarity.
- Concentrations are *not* added when solutions are mixed.

Self-Tutorial Problems

11.1 If 25.0 mL of a solution is poured from 100 mL of a 1.07-M sample, what is the concentration of the 25.0-mL portion?

11.2 What is the final concentration when 2.00 L of 3.00 M sugar solution is added to 3.00 L of 3.00 M sugar solution?

11.3 What is the difference between (a) "dilute the solution to 5.00 L with water" and (b) "dilute the solution with 5.00 L of water"?

11.4 (a) If exactly one-thousandth of a 1.000-L sample of 2.000 M solution is poured into a small beaker, how many milliliters of solution are in the beaker?
(b) How many millimoles of solute are in the beaker?
(c) What is the concentration of the solution in the beaker?
(d) What is the concentration in millimoles per milliliter in the beaker?

11.5 Which of the following combinations of solutions will result in a chemical reaction, which will result in a combination of the numbers of moles of a common ion, and which will result in a mere dilution?
(a) $NaOH(aq) + HNO_3(aq)$
(b) $NaOH(aq) + NaCl(aq)$
(c) $KOH(aq) + NaCl(aq)$

11.6 (a) If 3 dozen couples get married at city hall on a certain weekend, how many brides are there? How many grooms?
(b) What is the concentration of the cation in a 3.0 M solution of NaCl? What is the concentration of the anion?

11.7 What is the concentration of each ion in the following solutions?
(a) 1.0 M solution of NaCl
(b) 1.0 M solution of $CaCl_2$
(c) 1.0 M solution of $AlCl_3$
(d) 1.0 M solution of $NaClO_3$
(e) 1.0 M solution of $Ca(ClO_3)_2$
(f) 1.0 M solution of $Al(ClO_3)_3$

11.8 What is the final volume of solution if 2.0 L of solution is diluted (a) with 3.0 L of solvent or (b) to 3.0 L with solvent?

Problems

11.1 Definition and Uses of Molarity

11.9 Calculate the molarity of a solution containing 3.00 mol of solute in 250 mL of solution.

11.10 Calculate the molarity of a solution of which 100 mL contains 25.0 g of formaldehyde, CH_2O.

11.11 Calculate the molarity of a solution containing 32.10 mmol of solute in 300.0 mL of solution.

11.12 Calculate the number of moles of solute in 3.01 L of 2.22 M solution.

11.13 Calculate the number of grams of NaF in 2.00 L of 3.00 M NaF solution.

11.14 Calculate the volume of 3.09 M solution that contains 2.07 mol of solute.

11.15 Calculate the number of milliliters of 2.00 M NaCl solution that contains 12.4 g of NaCl.

11.16 What is the final concentration if 2.0 L of 1.2 M solution is diluted (a) with 3.0 L of solvent or (b) to 3.0 L with solvent?

11.17 Calculate the molarity of a solution prepared by diluting 25.0 mL of 3.53 M solution to 75.0 mL.

11.18 Calculate the volume of solution prepared by diluting 25.0 mL of 3.53 M solution to 1.37 M.

11.19 Calculate the volume of 3.00 M solution required to make 7.00 L of 1.25 M solution by dilution with water.

11.2 Molarities of Ions

11.20 Calculate the concentration of each ion in each of the following solutions:
(a) 1.25 M NaCl
(b) 3.11 M $CuCl_2$
(c) 0.955 M K_2SO_4
(d) 1.35 M $Al_2(SO_4)_3$
(e) 0.100 M $(NH_4)_2SO_4$

11.21 Find the concentration of each type of ion in solution after 50.0 mL of 2.00 M Na_2SO_4 is diluted to 125 mL.

11.22 In which of the following combinations of solutions will there be a chemical reaction? Which have ions in common? In which are the ions all different and unreactive?

Problems

This set of problems is sorted by section, that is, the major topics presented in the chapter. These problems do not merely ask for a restatement of the information in the chapter, and students will find them interesting as well as challenging.

General Problems

The final set of problems provided for each chapter is more difficult than the previous sets, is not classified by topic, and frequently requires knowledge of two or more concepts. Similar in scope to the type of questions students will be confronted with on tests, these problems provide students with an excellent means by which to judge their knowledge of the chapter's contents.

General Problems **227**

(a) $NaCl(aq) + CaCl_2(aq)$

(b) $NaCl(aq) + KNO_3(aq)$

(c) $NaCl(aq) + AgNO_3(aq)$

(d) $Ba(OH)_2(aq) + H_2SO_4(aq)$

11.23 Calculate the concentration of each ion in solution after 2.08 L of 1.37 M NaCl is mixed with 1.62 L of 1.05 M $AlCl_3$ and the resulting solution is diluted to 5.00 L.

11.24 Calculate the concentration of each type of ion in solution after 50.0 mL of 1.00 M NaCl and 50.0 mL of 1.00 M Na_2SO_4 are mixed. Assume that the final volume is 100 mL.

11.25 If 0.250 mol of $Al_2(SO_4)_3$ and 0.250 mol of Na_2SO_4 are dissolved in enough water to make 250 mL of solution, what is the concentration of each ion in the solution?

11.26 Calculate the concentration of each ion in 0.200 M $Hg_2(NO_3)_2$ solution.

11.27 What is the concentration of each type of ion in solution after 45.5 mL of 0.465 M $HClO_3$ is added to 50.0 mL of 0.750 M NaOH? Assume that the final volume is the sum of the original volumes.

11.3 Titration

11.28 Calculate the concentration of an H_2SO_4 solution if 25.00 mL is completely neutralized by 43.57 mL of 3.000 M NaOH solution.

11.29 Calculate the concentration of a sulfuric acid solution if 25.00 mL is completely neutralized by 43.57 mL of 3.000 M sodium hydroxide solution.

11.30 When 7.850 g of potassium hydrogen phthalate (symbolized as KHPh; molar mass = 204.2 g/mol) is titrated with KOH solution, it takes 42.07 mL of the base to achieve the end point. Calculate the concentration of the KOH solution.

$$KHPh(aq) + KOH(aq) \rightarrow K_2Ph(aq) + H_2O(l)$$

11.31 An antacid tablet contains 15.0 g of $NaHCO_3$. What volume of 4.12 M stomach acid (HCl) can this tablet neutralize?

General Problems

11.32 Calculate the concentration of each ion in solution after 2.08 L of 1.37 M NaCl is mixed with 1.62 L of 1.05 M $ZnCl_2$ and then diluted to 5.00 L.

11.33 Calculate the concentration of each ion in solution after 4.08 L of 1.37 M NaOH is mixed with 1.62 L of 1.05 M H_2SO_4 and then diluted to 8.00 L.

11.34 (a) Calculate the concentration of each type of ion in 75.90 mL of a solution containing 35.74 mmol of NaCl plus 14.76 mmol of HCl.

(b) Calculate the concentration of each type of ion in solution after 50.50 mL of 1.000 M HCl is added to 25.40 mL of 1.407 M NaOH. Assume that the volume of the final solution is the sum of the volumes of the two original solutions.

(c) Explain how parts (a) and (b) are related.

11.35 Calculate the concentration of each type of ion in solution after 50.50 mL of 1.000 M HCl is added to 25.40 mL of 1.407 M NaOH. Use a net ionic equation in solving this problem. Assume that the final volume is equal to the sum of the volumes of the two original solutions.

11.36 When an alkali metal oxide is treated with water, it reacts with the water to form hydroxide ions. What concentration of hydroxide ions is present if 0.250 mol of solid Li_2O is treated with water and the final volume is 500 mL.

11.37 Calculate the concentration of H^+ ion produced when H_2S is bubbled into 0.500 M Cu^{2+} solution, causing precipitation of all the copper(II) ion as CuS. Assume no change in the volume of the solution.

$$Cu^{2+}(aq) + H_2S(g) \rightarrow CuS(s) + 2\ H^+(aq)$$

11.38 Calculate the concentration of CH_2O in a solution prepared by mixing 1.50 L of 2.05 M CH_2O and 1.80 L of 0.894 M CH_2O and diluting the mixture to 4.00 L with water.

11.39 Calculate the concentrations of acetate ion and acetic acid in solution after 100 mL of 2.00 M $HC_2H_3O_2$ and 100 mL of 1.00 M NaOH are mixed. Assume that the final volume is 200 mL and that the excess acetic acid yields no acetate ions to the final solution (since it is a *weak* acid).

$$OH^-(aq) + HC_2H_3O_2(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$$

11.40 Calculate the percentage of $CaCO_3$ in a 5.00-g sample of limestone if 27.20 mL of 3.000 M HCl is required to react completely with the $CaCO_3$. Assume that the rest of the limestone sample is inert.

11.41 In a certain experiment, 25.00 mL of 2.000 M H_3PO_4 was titrated to a certain end point with 33.85 mL of 2.954 M NaOH. Write the equation for the chemical reaction that occurred.

11.42 What volume of 2.954 M NaOH would be required to completely neutralize the H_3PO_4 in Problem 11.41?

11.43 There are several acids in vinegar. Calculate the total concentration of acids in a 10.0-mL sample of vinegar if it takes 23.40 mL of 0.1000 M NaOH to neutralize the acids. Assume that each acid contains only one ionizable hydrogen atom per formula unit.

11.44 Calculate the total concentration of all ions in each of the following solutions:

(a) 0.100 M NaCl (b) 0.100 M $BaCl_2$

(c) 24.7 g of $(NH_4)_2SO_4$ in 155 mL of solution

Supplemental Materials

An extensive supplemental package has been designed to support this book. It includes the following elements:

1. *Instructor's Manual*. The instructor's manual contains the printed test item file, a list of transparencies, and suggestions on how to organize the course.
2. *Student Study Guide*. The student study guide offers students a variety of exercises, self-tests, and hints to promote their comprehension of the basics as well as the more difficult concepts.
3. *Transparencies*. A set of 50 color transparencies is available to help the instructor coordinate the lecture with key illustrations from the text.
4. *Customized Transparency Service*. If adopters are interested in acetates of text figures not included in the standard transparency set, those acetates will be custom-made upon request. Contact your local Wm. C. Brown Publishers sales representative for more information.
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