

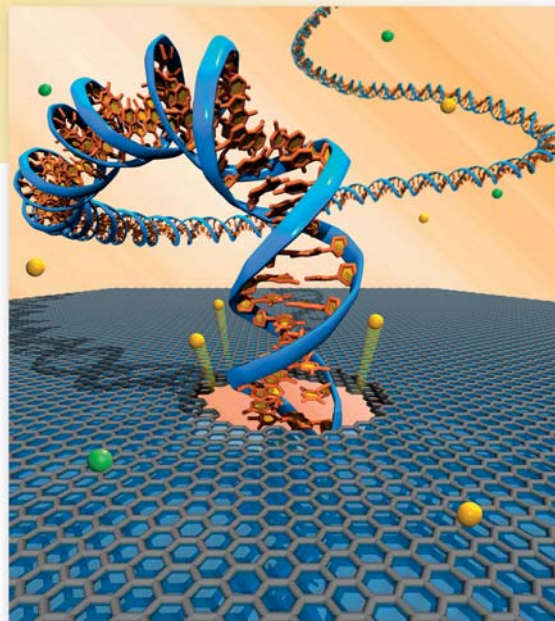
# Chapter 1

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## Chemistry The Study of Change

### Chapter Contents

- 1.1 Chemistry: A Science for the Twenty-First Century
- 1.2 The Study of Chemistry
- 1.3 The Scientific Method
- 1.4 Classifications of Matter 1.A.1 2.A.3
- 1.5 The Three States of Matter 1.A.1 2.A.1 2.A.2
- 1.6 Physical and Chemical Properties of Matter
- 1.7 Measurement
- 1.8 Handling Numbers
- 1.9 Dimensional Analysis in Solving Problems
- 1.10 Real-World Problem Solving: Information, Assumptions, and Simplifications



By applying electric fields to push DNA molecules through pores created in graphene, scientists have developed a technique that someday can be used for fast sequencing the four chemical bases according to their unique electrical properties.

### Introduction

Chemistry surrounds us. It determines the myriad of interactions needed for our bodies to function. Its laws determine the function of the food we eat and the water we drink. It is in our daily routines. Consider the car or bus ride to school. As a result of chemical interactions, a vehicle starts when the ignition is turned on and accelerates when the gas pedal is depressed. A mini explosion occurs within each cylinder and that energy is transferred to turn the wheels of the car. The tires grip the road with a prescribed air pressure within. Exhaust fumes are cleaned up by a catalytic converter. A solution of alcohol and water is sprayed on the windshield, improving visibility. Halogen headlights, an interaction of matter and energy, show us the road in the early morning hours. The car or bus is a traveling road show of chemistry! The *challenge* of chemistry is to connect each of these visible events with the invisible particles that cause them to happen.

In Advanced Placement chemistry, these particles—atoms, ions and molecules—are introduced to you in **Big Idea 1**, the Atomic Theory. As explained in **Big Idea 2**, these particles will often group together and demonstrate behavior dictated by their arrangements and attractions for one another. Such arrangements may be reorganized through chemical and physical interactions as described in **Big Idea 3**. Interactions occur at different speeds. Some are instantaneous while others are slow. **Big Idea 4** investigates how particle collisions determine this speed. Reactions may heat or cool their surroundings. The role of energy in the outcome of a reaction is understood through thermodynamics, as indicated in **Big Idea 5**. In nature, many reactions—particularly biological ones—occur in solution where interactions are reversible. Equilibrium occurs when an interaction and its reverse occur at the same speed. **Big Idea 6** provides a comprehensive survey of those principles of equilibrium chemistry.

The six Big Ideas of AP Chemistry are an instructional guide. When followed, one step at a time, you will better understand how your world works—from headlights to taillights and everything in between.

# Chapter 2

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## Atoms, Molecules, and Ions



Illustration depicting Marie and Pierre Curie at work in their laboratory. The Curies studied and identified many radioactive elements.

### Chapter Contents

- 2.1 The Atomic Theory 1.A.1 1.D.1
- 2.2 The Structure of the Atom 1.B.1 1.D.1
- 2.3 Atomic Number, Mass Number, and Isotopes 1.D.2
- 2.4 The Periodic Table 1.C.1
- 2.5 Molecules and Ions 1.A.1 1.B.1
- 2.6 Chemical Formulas 1.A.1 2.C.2
- 2.7 Naming Compounds
- 2.8 Introduction to Organic Compounds

### Introduction

The behavior of the metallic element, sodium, is dramatic. It interacts vigorously when placed in water. Yet when a single electron is removed from an atom of sodium, it becomes a positively charged particle that is present in ordinary table salt. And it dissolves readily—and *unspectacularly*—in water. Its behavior to some degree is copied by the elements of lithium and potassium, found directly above and below sodium on the Periodic Table. Yet fluorine and calcium, which are one atomic number lower and higher than sodium, have very different chemical behaviors compared to sodium.

To account for chemical behavior we turn to the structure of the atom. In this respect, the number and placement of electrons is especially crucial. Although every element has a unique behavior, there are patterns among them. The Periodic Table helps us see and use these patterns (EK.1.C.1). Finally, a system by which we name and symbolize particles exists so that sodium the atom (Na) will not be confused with sodium the ion ( $\text{Na}^+$ ) or carbon monoxide (CO) —a lethal gas—is not mistaken for carbon dioxide ( $\text{CO}_2$ ), the gas we exhale when breathing and swallow when drinking carbonated beverages.

# Chapter 3

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## Mass Relationships in Chemical Reactions

### Chapter Contents

- 3.1** Atomic Mass 1.A.3
- 3.2** Avogadro's Number and the Molar Mass of an Element 1.A.3
- 3.3** Molecular Mass 1.A.3
- 3.4** The Mass Spectrometer 1.D.2
- 3.5** Percent Composition of Compounds 1.A.1 1.A.2 1.A.3
- 3.6** Experimental Determination of Empirical Formulas 1.A.1 1.A.2 1.A.3 1.E.2
- 3.7** Chemical Reactions and Chemical Equations 1.A.1 1.E.1 1.E.2 3.A.1 3.C.1
- 3.8** Amounts of Reactants and Products 1.A.1 1.A.3 1.E.1 1.E.2 3.A.1 3.A.2
- 3.9** Limiting Reagents 1.A.3 1.E.1 1.E.2 3.A.1 3.A.2
- 3.10** Reaction Yield 1.A.1 1.A.3 1.E.1 1.E.2 3.A.1 3.A.2



Fireworks are chemical reactions noted for the spectacular colors rather than the energy or useful substances they produce.

### Introduction

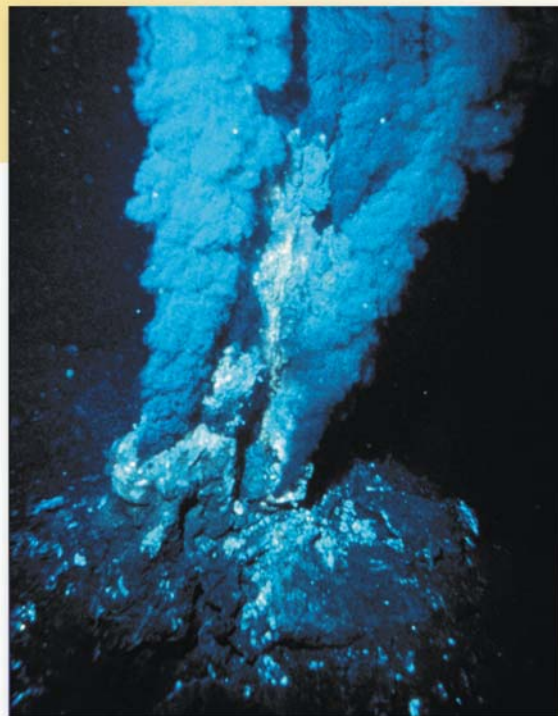
Chemists and cooks have much in common. They perform procedures, which call for a set of ingredients in prescribed amounts. The amounts can be varied but only proportionally. Their starting materials are changed—often with the help of heat—producing new substances with new properties but containing all of the original elements in their original quantities. Whether it's baking chocolate chip cookies or creating a fireworks display, it is crucial that the cook and the chemist know exactly what starting materials they are using and in what amounts they are to be combined. For this to happen, the cook follows a recipe. The chemist uses a chemical equation to determine starting materials and quantities.

To use an equation, the chemist must know how to count particles. Counting particles is challenging since atomic sizes are very small and numerous in any visible sample. To help, a large count unit called the “mole” is used. The mole unit is linked to the mass of a substance, allowing the chemist to go from a macroscopic measurement of mass to the microscopic level of particles (EK.1.A.3). While a cook uses measuring cups and teaspoons, a chemist uses masses, moles, and a chemical equation to determine the amounts of reactant needed to produce a desired amount of product. And from the actual amount of product formed, the efficiency of the reaction can be determined (EK.1.E.2.a-d).

# Chapter 4

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## Reactions in Aqueous Solutions



Black smokers form when superheated water, rich in minerals, flows out onto the ocean floor through the lava from an ocean volcano. The hydrogen sulfide present converts the metal ions to insoluble metal sulfides.

### Chapter Contents

- 4.1 General Properties of Aqueous Solutions 2.A.1 2.A.3 2.B.2 2.D.1 3.A.1 6.A.1
- 4.2 Precipitation Reactions 1.A.1 3.A.1 3.C.1
- 4.3 Acid-Base Reactions 1.E.1 3.A.1 3.B.2
- 4.4 Oxidation-Reduction Reactions 1.E.1 3.A.1 3.B.1 3.B.3
- 4.5 Concentration of Solutions 1.A.3 2.A.3
- 4.6 Gravimetric Analysis 1.A.2 1.E.2 3.A.1 3.A.2
- 4.7 Acid-Base Titrations 1.A.2 1.E.2 3.A.2 3.B.2
- 4.8 Redox Titrations 1.A.2 1.E.2 3.A.2 3.B.3

### Introduction

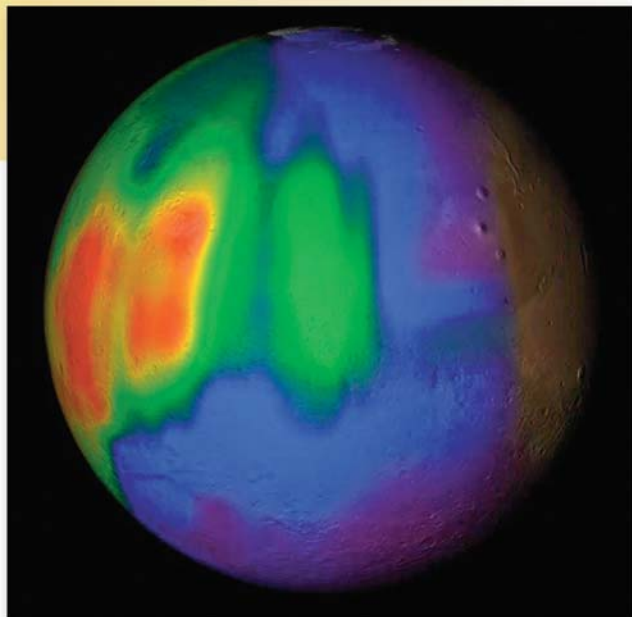
Reactions occurring in water can be classified into three categories: acid-base reactions, precipitation reactions, and oxidation-reduction (redox) reactions. To find examples of these look no further than your own body, your bathtub, and your backyard. An acid-base reaction takes place when an antacid—a mild base—neutralizes an over secretion of stomach acid. A precipitation reaction occurs when calcium ions combine with stearate ions to produce a ring of soap scum around your bathtub or shower. An oxidation-reduction reaction occurs when the iron of a garden tool is exposed to air—in the presence of moisture—to form rust. Notice that water is a common requirement for each reaction. And since water is in our bodies, our bathtubs, and in the overnight dew in our yards, it is no surprise that these reactions are a common part of our everyday world.

Specific groups of particles are involved in each reaction type (EK. 3.B.2; EK.3.B.3). How these particles interact to produce the effects above can be represented by molecular, ionic, and net ionic equations (EK.3.A.1). Equations can be used—along with the lab techniques of titration and gravimetric analysis—to determine quantities of reactants and products involved in these reactions (EK.3.A.2). Questions concerning the acid content of a food, the concentration of lead ions in drinking water, and the blood alcohol content of a suspected drunk driver can be answered through the use of these analytical techniques.

# Chapter 5

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## Gases



Water vapor and methane have recently been detected in significant amounts in the Martian atmosphere. (The concentration increases from purple to red.) The methane could be released by geothermal activity, or it could be produced by bacteria, fueling speculation that there may be life on Mars.

### Chapter Contents

- 5.1 Substances That Exist as Gases 2.A.2 2.B.3 2.C.2
- 5.2 Pressure of a Gas 2.A.2
- 5.3 The Gas Laws 1.E.1 2.A.2
- 5.4 The Ideal Gas Equation 2.A.2
- 5.5 Gas Stoichiometry 1.E.2 2.A.2 3.A.2
- 5.6 Dalton's Law of Partial Pressures 2.A.2
- 5.7 The Kinetic Molecular Theory of Gases 2.A.2 5.A.1
- 5.8 Deviation from Ideal Behavior 2.A.2

### Introduction

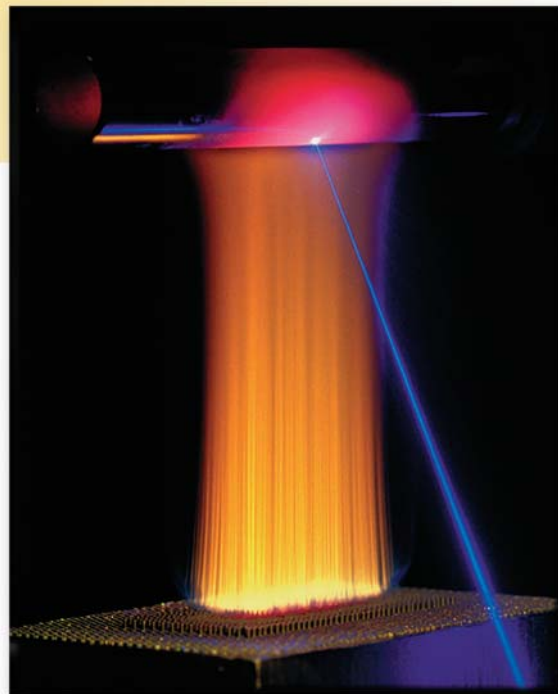
We live at the bottom of a sea of colorless gases and often take them for granted. But make no mistake; gases have properties that can be measurable, dramatic, and life altering. For example, tucked behind the steering wheel of a modern car lies a canister containing a pellet of sodium azide ( $\text{Na}_3\text{N}$ ). When activated by the sudden deceleration of the car during a collision, the pellet decomposes into several products—one of which is nitrogen gas. In less than a second, this gas inflates a bag which pops out of the dashboard or steering wheel, meeting the driver before he or she impacts the steering wheel and thereby lessening the severity of injuries. The amount of gas needed to properly inflate the bag is determined by stoichiometry of the decomposition reaction of the sodium azide and gas laws describing the relationship between pressure, temperature, and volume. These gas laws also describe how our lungs allow us to inhale and exhale effectively, why auto tires increase pressure on a hot, summer day and why a kernel of popcorn will pop when heated.

While gas laws *describe and predict* the behavior of gases (EK.2.A.2), the kinetic theory of matter *explains* their behavior (EK.2.A.2). This theory proposes that gas particles are essentially independent of each other under standard temperatures and pressures, and have free and random movement. Our Ideal Gas Law is derived from this model.

# Chapter 6

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## Thermochemistry



The analysis of particles formed from burning methane in a flame is performed with a visible laser.

### Chapter Contents

- 6.1** The Nature of Energy and Types of Energy 5.B.2 5.C.1 5.D.1
- 6.2** Energy Changes in Chemical Reactions 3.C.2 5.B.1 5.B.2 5.B.3 5.C.2 5.D.2
- 6.3** Introduction to Thermodynamics 3.C.2 5.A.2 5.B.1 5.B.2 5.B.3
- 6.4** Enthalpy of Chemical Reactions 1.E.2 3.C.2 5.B.1 5.B.2 5.B.3
- 6.5** Calorimetry 3.C.2 5.A.2 5.B.1 5.B.2 5.B.3 5.B.4
- 6.6** Standard Enthalpy of Formation and Reaction 3.C.2 5.C.2
- 6.7** Heat of Solution and Dilution 3.C.2 5.B.1 5.B.2 5.B.3 5.C.2 5.D.2

### Introduction

Reactions that release energy are familiar and can be useful. The burning of propane provides heat, which can be used for cooking. Yet there are reactions which consume energy as well. If you place your hands around a container, which has salt dissolving in water, you may feel a cold sensation—a sign that the dissolving reaction is absorbing heat from your hands (EK.3.C.2; EK.5.B.3). Both energy release and energy consuming reactions involve the breaking and the formation of chemical bonds. Despite having this in common, one releases energy while the other absorbs. How can this happen? What makes one reaction type different from the other?

Breaking bonds requires an input of energy. Creating new bonds results in an output of energy. The change we feel is the net energy of the two opposing processes (EK.5.C.2). This net energy is the heat that cooks the hamburger on your grill and is the heat loss of your hands to the container of dissolving salt. These net energies can be measured through a device called a calorimeter (calorie-meter!). Using the Law of Conservation of Energy (EK.5.B.2), the energy change of the calorimeter is linked to the energy change of a reaction (EK.5.B.4). This energy can be included in an equation representing a reaction and used along with moles and masses to determine the amount of energy released or consumed by a reaction.

# Quantum Theory and the Electronic Structure of Atoms



“Neon light” is a generic term for atomic emission involving various noble gases, mercury, and phosphor. The UV light from excited mercury atoms causes phosphor-coated tubes to fluoresce white light and other colors.

## Chapter Contents

- 7.1 From Classical Physics to Quantum Theory 1.C.2 1.D.1 1.D.3
- 7.2 The Photoelectric Effect 1.B.1 1.D.3
- 7.3 Bohr’s Theory of the Hydrogen Atom 1.C.2 1.D.3
- 7.4 The Dual Nature of the Electron 1.C.2 1.D.1 1.D.3
- 7.5 Quantum Mechanics 1.C.2 1.D.1
- 7.6 Quantum Numbers 1.B.2 1.C.2
- 7.7 Atomic Orbitals 1.B.2
- 7.8 Electron Configuration 1.B.2
- 7.9 The Building-Up Principle 1.B.2 1.C.1

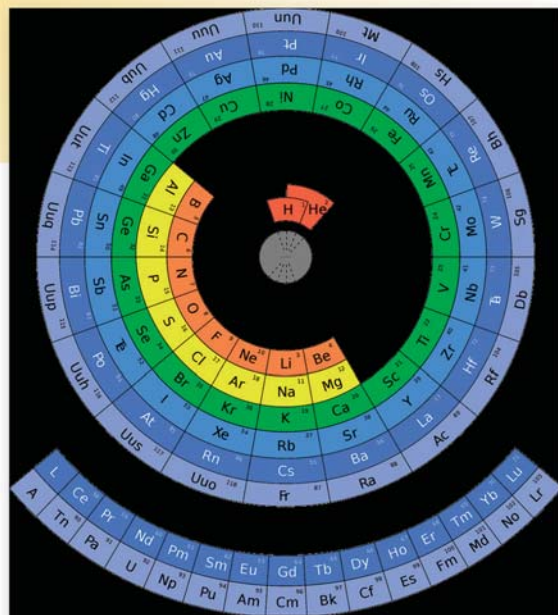
## Introduction

What happens inside an atom to release the “neon light” pictured above? This is a “black box” problem, where the inner structure cannot be viewed directly and can only be solved by observing what goes in and what comes out as we prod and poke the system. Light energy emitted by atoms in a high-energy state provides clues to solving the inner structure of the atom (EK.1.B.2). Other studies including the photoelectric effect (EK.1.B.1) and the behavior of atoms in a magnetic field provide evidence for the arrangement of electrons within the atom. Results of these experiments led to a “remodeling” of the atom during the 1900s, with each successive model building upon earlier models. Our current model, the Quantum Theory, not only explains the production of light in neon signs but also the periodic properties of the elements (EK.1.C.2), the bonds and shapes of molecules and ultimately, the behavior of all matter.

# Chapter 8

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## Periodic Relationships Among the Elements



The periodic table takes many different forms from the days of Mendeleev. This circular version shows that as one moves towards the center, atomic size decreases.

### Chapter Contents

- 8.1 Development of the Periodic Table
- 8.2 Periodic Classification of the Elements 1.C.1
- 8.3 Periodic Variation in Physical Properties 1.C.1
- 8.4 Ionization Energy 1.C.1 1.D.1
- 8.5 Electron Affinity 1.C.1 1.D.1
- 8.6 Variation in Chemical Properties of the Representative Elements 1.C.1

### Introduction

The periodic table is a scientific masterpiece, which summarizes much of what we know about the elements. On one level you will discover that while every element is unique, each element has properties similar to other elements, and when elements are arranged according to increasing atomic number, *periodic* patterns of properties become evident (EK.1.C.1). On a second level, you will see the genius in the table's presentation of these patterns. With the alkali metals on the left side, traveling through the transition and inner transition elements in the middle, and ending with the noble gases on the right side, the table's iconic arrangement permits an easy prediction of elemental properties. On a third level, you will find a relationship between the periodic table and the Quantum Theory of the atom (EK.1.B.2), allowing for explanations of periodic patterns.

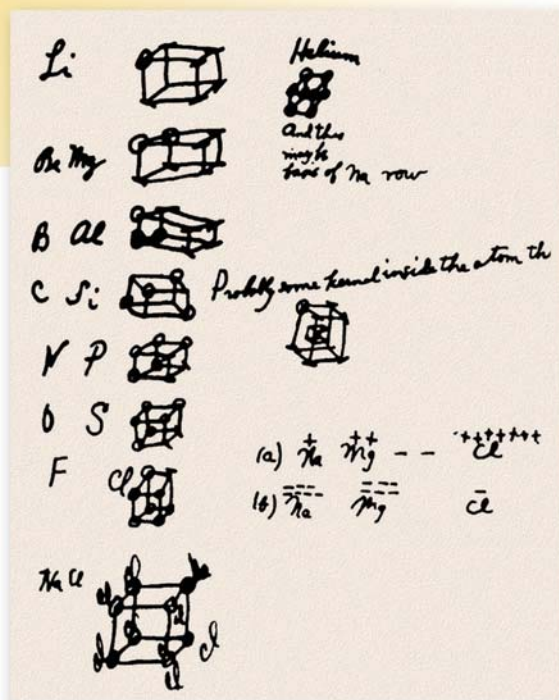
Sodium metal reacts vigorously with water. If potassium metal were added to water, would it also react? Why or why not? Both questions should be considered by the AP student. The first is easily answered through the periodic table. The second question is more challenging and requires an understanding and application of Quantum Theory in addition to Coulomb's Law (EK.1.B.2; EK.1.C.1).



# Chapter 9

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## Chemical Bonding I Basic Concepts



Lewis first sketched his idea about the octet rule on the back of an envelope.

### Chapter Contents

- 9.1 Lewis Dot Symbols [1.C.1](#) [2.C.1](#) [2.C.2](#)
- 9.2 The Ionic Bond [2.C.1](#) [2.C.2](#)
- 9.3 Lattice Energy of Ionic Compounds [1.B.1](#) [2.C.2](#) [2.D.1](#)
- 9.4 The Covalent Bond [2.C.1](#) [2.D.1](#) [2.D.4](#)
- 9.5 Electronegativity [1.C.1](#) [2.C.1](#)
- 9.6 Writing Lewis Structures [2.C.1](#) [2.C.4](#)
- 9.7 Formal Charge and Lewis Structure
- 9.8 The Concept of Resonance [2.C.4](#)
- 9.9 Exceptions to the Octet Rule [2.C.4](#)
- 9.10 Bond Enthalpy [3.C.2](#) [5.C.1](#) [5.C.2](#)

### Introduction

Sodium metal can be cut with a knife and reacts vigorously with water. Elemental chlorine is a smelly, toxic gas. When they are placed together, a highly exothermic reaction occurs producing sodium chloride, an extremely stable solid that is the common seasoning—table salt. What changes occurred in these elements that result in a substance that we sprinkle on our French fries? The answer lies in the interaction of their outer shell or valence electrons, which form chemical bonds. These bonds give sodium chloride its stability and a set of properties unlike those of the elements from which it was formed. Bonds between elements are at the heart of millions of unique compounds.

Chapter 9 focuses on representing bonds through Lewis electron dot diagrams. Following a few rules, you can easily create these diagrams to represent the valence electrons involved and the bonds they make ([EK.2.C.4](#)). Not all bonds are the same. On a continuum, bonds range from equal sharing to unequal sharing to an effective transfer of valence electrons from one atom to another. The bond type is predicted through the use of electronegativity, a periodic property of the atoms involved in the bond. ([EK.2.C.1](#); [EK.2.C.2](#)). As seen in Chapter 10, bond type along with the 3-D shape of a molecule determine its chemical and physical properties.

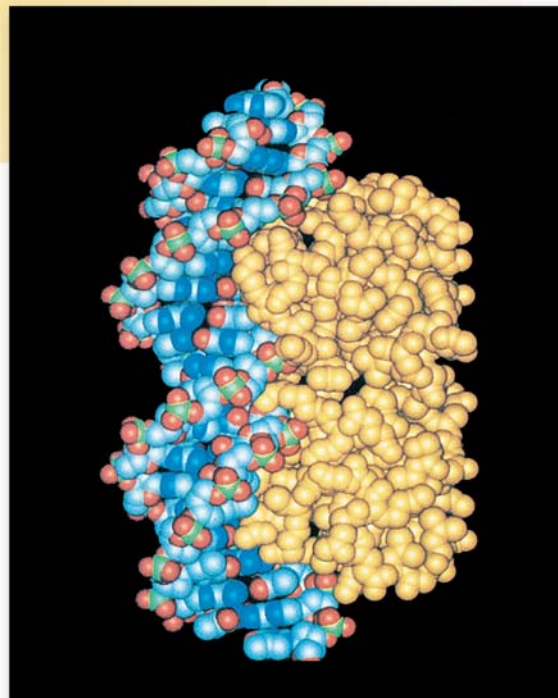
Although abstract in its nature, the chemical bond is real in its effect on us and our world. For example, read about the life saving molecule, nitric oxide, in “Just Say NO” on page 399. It is chemistry in action.

# Chapter 10

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## Chemical Bonding II

### Molecular Geometry and Hybridization of Atomic Orbitals



The shape of molecules plays an important role in complex biochemical reactions such as those between protein and DNA molecules.

### Chapter Contents

- 10.1** Molecular Geometry 2.C.4
- 10.2** Dipole Moments 2.C.1 2.C.4
- 10.3** Valence Bond Theory 2.C.4
- 10.4** Hybridization of Atomic Orbitals 2.C.4
- 10.5** Hybridization in Molecules Containing Double and Triple Bonds 2.C.4
- 10.6** Molecular Orbital Theory
- 10.7** Molecular Orbital Configurations
- 10.8** Delocalized Molecular Orbitals 2.C.4

### Introduction

Although molecular diagrams are most often represented as two-dimensional and appear flat, molecules themselves are not. Like us, molecules have three dimensions, filling space along their x, y and z axes with single electrons and electron pairs—either in bonds or alone. Electrons repel and, along with the atoms they are attached to, move into positions that minimize this repulsion while maintaining the bonds holding the molecule together. The result is a 3-D structure. Shapes of molecules will vary depending on the number of atoms and electrons involved and these shapes can be predicted using Valence Shell Electron Pair Repulsion (VSEPR) theory ([EK.2.C.4](#)). Knowing the shape of a molecule, a person can explain and predict the chemical and physical properties for that molecule. For example, most solids sink in their own liquids. Yet an ice cube floats on water. This unusual behavior of water can be understood by considering the bent shape of the water molecule and the polarity and intermolecular bonds that result.

To understand the behavior of a molecule one must consider not only its bonds but also the shape that it assumes. Chapter 10 will guide you in predicting molecular shapes and explain how these shapes impact the properties of molecular compounds.

# Chapter 11

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## Intermolecular Forces and Liquids and Solids



A person throwing boiling water into the air at  $-51^{\circ}\text{C}$ .

### Chapter Contents

- 11.1 The Kinetic Molecular Theory of Liquids and Solids 2.A.1
- 11.2 Intermolecular Forces 2.B.1 2.B.2 2.B.3 5.D.1
- 11.3 Properties of Liquids 2.A.1 2.A.3 2.B.2
- 11.4 Crystal Structure
- 11.5 X-Ray Diffraction by Crystals
- 11.6 Types of Crystals 2.C.2 2.C.3 2.D.1 2.D.2 2.D.3 2.D.4
- 11.7 Amorphous Solids
- 11.8 Phase Changes 2.A.1 2.A.2 5.A.1 5.B.3 5.D.1 5.D.2 6.A.1
- 11.9 Phase Diagrams

### Introduction

Carbon has three main crystalline forms: diamond, graphite, and buckyballs. Each form has its own unique characteristics. Diamond is the hardest naturally occurring substance, a poor conductor of electricity, and transparent in its pure form. Graphite is soft enough to be used in a pencil to make marks on paper, is an excellent conductor of electricity, and has a dull gray color (EK.2.D.3). Buckyballs are found as carbon atoms in a cage shape that can be joined to give long fibers called nanotubes. Nanotubes conduct electricity and are extremely strong. Applications of nanotubes are many and varied—from tennis rackets to human tissue engineering—and have given rise to a field of study called nanotechnology.

Differences in properties can be explained by considering the bonds that join the carbon atoms together in each crystal. Bonds linking one particle to another to form liquids and solids are called *intermolecular* bonds. They are related to but different from *intramolecular* bonds which cause atoms to be held together within a molecule—as described in Chapters 9 and 10. The challenge of Chapter 11 is to relate the physical properties of liquids and solids to the various types of intermolecular forces, which account for their attraction to each other (EK.2.A.1; EK.2.B.1–2.B.3; EK.2.C.2; EK.2.C.3; EK.2.D.1–2.D.4).

# Chapter 12

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## Physical Properties of Solutions



A sugar cube dissolving in water. The properties of a solution are markedly different from those of its solvent.

### Chapter Contents

- 12.1** Types of Solutions 2.A.3
- 12.2** A Molecular View of the Solution Process 2.A.3 2.B.3 5.E.1
- 12.3** Concentration Units 2.A.3
- 12.4** The Effect of Temperature on Solubility
- 12.5** The Effect of Pressure on the Solubility of Gases
- 12.6** Colligative Properties of Nonelectrolyte Solutions 2.A.3
- 12.7** Colligative Properties of Electrolyte Solutions 2.A.3
- 12.8** Colloids

### Introduction

Assume a solid “X” is added to a liquid “Y”. Will a solution form as with salt added to water? Or will X and Y remain separate as with sand added to water? The outcome can generally be predicted by knowing the types of intermolecular forces in each component and by applying the axiom “like dissolves like.” That is if the intermolecular forces of X and Y are “alike,” X and Y will form a solution. If the forces are not “alike,” X and Y will not form a solution, but rather will remain a heterogeneous mixture. Why this axiom can be used to predict solubility is understood by comparing the bonds involved and their relative strengths. Solubility is truly a “battle of the bonds!” If solvent-solute bonds between molecules are stronger than the sum of bonds holding solute molecules together and solvent molecules together, a solution will form. (See page 119 for a definition of solute and solvent.) If the relative strengths are close, a solution may still form, driven by the natural tendency for objects to mix and become disorganized—a concept known as “entropy.”

Chapter 12 focuses on the role of intermolecular forces in the dissolving process, the method by which we express the concentration of the solute dissolved and the physical properties of the resulting solution (EK.2.A.3).

# Chapter 13

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## Chemical Kinetics



The rates of chemical reactions vary greatly. The conversion of graphite to diamond in Earth's crust may take millions of years to complete. Explosive reactions such as those of dynamite and TNT, on the other hand, are over in a fraction of a second.

### Chapter Contents

- 13.1** The Rate of a Reaction 4.A.1 4.A.3
- 13.2** The Rate Law 4.A.1 4.A.2 4.A.3
- 13.3** The Relation Between Reactant Concentration and Time 4.A.1 4.A.2 4.A.3
- 13.4** Activation Energy and Temperature Dependence of Rate Constants 4.A.3 4.B.1 4.B.2 4.B.3
- 13.5** Reaction Mechanisms 4.B.3 4.C.1 4.C.2 4.C.3
- 13.6** Catalysis 4.B.3 4.D.1 4.D.2

### Introduction

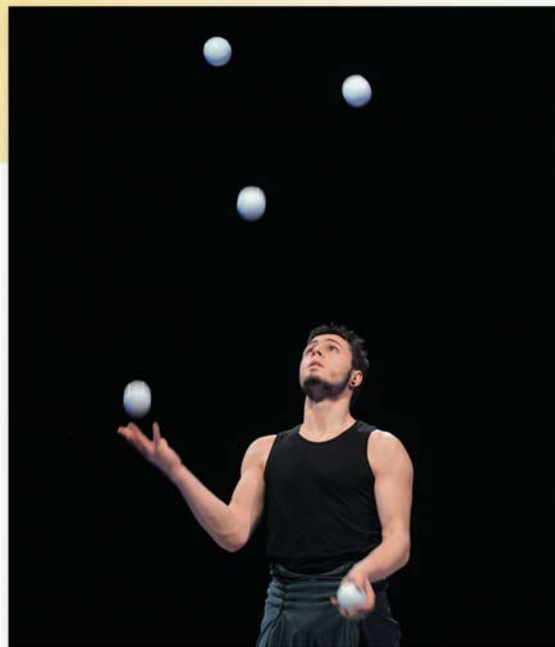
Chemical equations for burning wax and burning methane are similar. In both, oxygen is a reactant, and carbon dioxide, water, and energy are products. What is missing and is not shown in the equations is the vast difference in the speed at which these reactions occur. Normally, the burning of wax is measured over a span of hours, while methane burns much more rapidly. This difference in reaction speed or “rate” will result in one reaction occurring in a calm and peaceful manner, while the other will occur in an explosive and violent manner. Clearly, rates of reactions are important!

The focus of this chapter is to understand how and why reaction rates differ. We will establish ways to express and measure rates (EK.4.A.1)—in addition to examining factors that increase or decrease rates (EK.4.A.2,3). We will look at the role of catalysts in biological systems and in the recovery of our environment (EK.4.D.1,2). At the core of our understanding is the study of what occurs at the particle level as reactants collide, break old bonds, and form new bonds (EK.4.B.1-3; EK.4.C.1,2).

# Chapter 14

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## Chemical Equilibrium



Chemical equilibrium is an example of dynamic equilibrium, much like what the juggler is trying to establish here.

### Chapter Contents

- 14.1** The Concept of Equilibrium and the Equilibrium Constant 6.A.1 6.A.2 6.A.3 6.A.4
- 14.2** Writing Equilibrium Constant Expressions 6.A.1 6.A.2 6.A.3 6.A.4
- 14.3** The Relationship Between Chemical Kinetics and Chemical Equilibrium 6.A.3
- 14.4** What Does the Equilibrium Constant Tell Us? 6.A.1 6.A.2 6.A.3 6.A.4
- 14.5** Factors That Affect Chemical Equilibrium 6.A.2 6.A.3 6.B.1 6.B.2

### Introduction

When sugar is mixed with water, it will initially dissolve and sweeten the water. But if too much sugar is added, a layer of sugar collects on the bottom of the container. Stirring the mixture helps to dissolve the sugar but eventually, no additional changes will be observed in either the sweetness of the water or in the amount of undissolved sugar at the bottom. At this point, all outward signs of activity between the sugar and the water have disappeared. However, if you had the ability to view the system at the molecular level, you would see plenty of action. Solid sugar is still dissolving. Yet the water is no sweeter because the reverse process is also occurring; the dissolved sugar is reforming into the solid—and at the same rate! The two ongoing, and opposing reactions are balanced, causing the taste of the water and the amount of undissolved sugar at the bottom to remain constant. This balanced state is called “chemical equilibrium.”

When reactants and products are within a closed system—such as an aqueous solution—balance will be achieved. Balanced reactions are common (EK.6.A.1) and the nature of equilibrium can be described both qualitatively and quantitatively (EK.6.A.2-4). When balance or “equilibrium” is disturbed as a result of changes in concentration, temperature, or pressure, the equilibrium will change. The result of this imbalance can be predicted using Le Châtelier’s principle (EK.2.B.1-2).

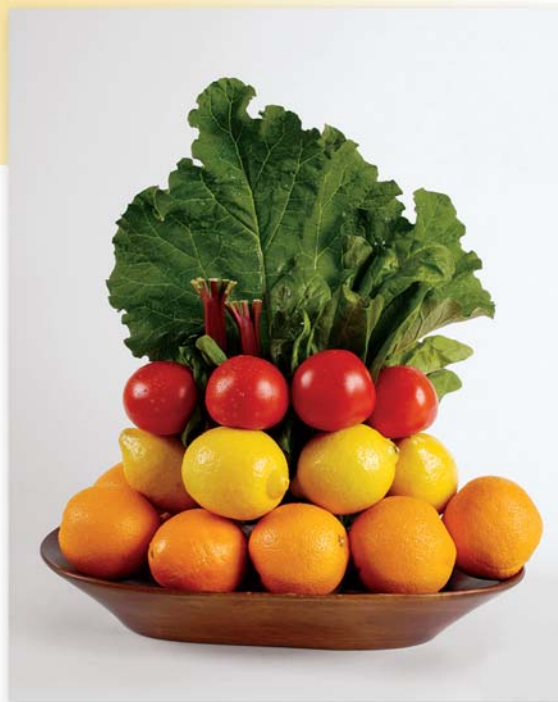
# Chapter 15

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## Acids and Bases

### Chapter Contents

- 15.1** Brønsted Acids and Bases 6.C.1
- 15.2** The Acid-Base Properties of Water 6.C.1 6.C.2
- 15.3** pH—A Measure of Acidity 6.C.1 6.C.2
- 15.4** Strength of Acids and Bases 6.C.1 6.C.2
- 15.5** Weak Acids and Acid Ionization Constants 6.C.1 6.C.2
- 15.6** Weak Bases and Base Ionization Constants 6.C.1 6.C.2
- 15.7** The Relationship Between the Ionization Constants of Acids and Their Conjugate Bases 6.C.1 6.C.2 6.A.4
- 15.8** Diprotic and Polyprotic Acids 6.C.1 6.C.2 6.A.4
- 15.9** Molecular Structure and the Strength of Acids 6.C.1
- 15.10** Acid-Base Properties of Salts 6.C.1 6.C.2
- 15.11** Acid-Base Properties of Oxides and Hydroxides 6.C.1 6.C.2
- 15.12** Lewis Acids and Bases



Many organic acids occur in the vegetable kingdom. Lemons, oranges, and tomatoes contain ascorbic acid, also known as vitamin C ( $C_6H_8O_6$ ), and citric acid ( $C_6H_8O_7$ ), and rhubarb and spinach contain oxalic acid ( $H_2C_2O_4$ ).

### Introduction

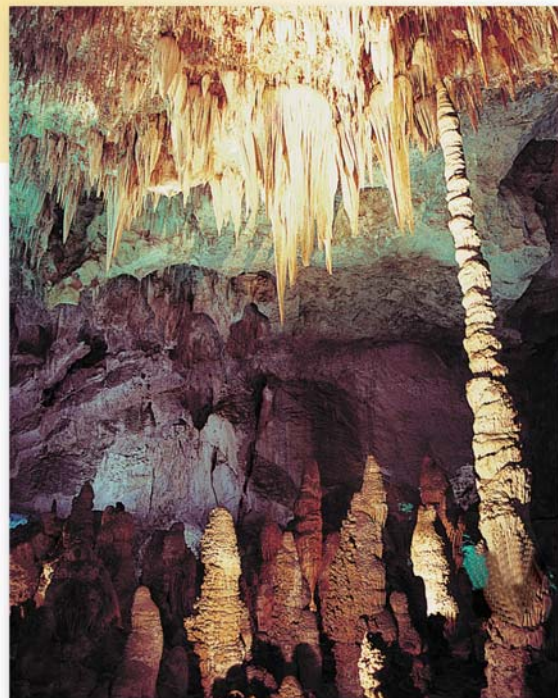
Acids and bases are common substances. Consider a simple breakfast. You might start with a glass of orange juice containing a weak but tangy acid called citric acid, followed by a helping of pancakes, made fluffy due to baking soda, a mild base. Or perhaps you ate a cup of yogurt containing a weak, slightly sour acid called lactic acid. All of these foods will be digested in your stomach with the help of a strong acid, hydrochloric acid, secreted naturally by your stomach lining. Stress of an impending chemistry test may lead to an over secretion of this acid causing stomach upset, prompting you to take an antacid tablet. The active ingredient in this tablet is a weak base, which neutralizes the acid and brings relief to your discomfort. (See *Chemistry in Action*, pp. 708–709). Many more examples of acids and bases can and will be given in this chapter. They are all around us. They are also within us.

The properties of acids and bases depend on the concentration of two ions: the hydrogen ion (also known as the hydronium ion), a characteristic of all acids, and the hydroxide ion, produced by the reaction of a base with water. The concentrations of these ions are determined by the strength and concentration of both the acid and the base involved, as well as by the interaction of these ions with water (EK.6.C.1). The level of acidity in any system, whether high or low, is indicated by its pH (i.e., the negative logarithm of the hydronium ion concentration) (EK.6.C.1). Chapter 15 takes you, step-by-step, through the process of finding the pH of any system, be it yogurt, orange juice or pancake batter.

# Chapter 16

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## Acid-Base Equilibria and Solubility Equilibria



Downward-growing icicle-like stalactites and upward-growing, columnar stalagmites. It may take thousands of years for these structures, which are mostly calcium carbonate, to form.

### Chapter Contents

- 16.1** Homogeneous versus Heterogeneous Solution Equilibria 6.C.1 6.C.3
- 16.2** The Common Ion Effect 6.B.1 6.B.2 6.B.4
- 16.3** Buffer Solutions 6.C.1 6.C.2
- 16.4** Acid-Base Titrations 1.E.2 3.A.2 3.B.2 6.C.1
- 16.5** Acid-Base Indicators 6.C.1 3.B.2
- 16.6** Solubility Equilibria 6.C.3
- 16.7** Separation of Ions by Fractional Precipitation 6.B.1 6.B.2 6.C.3
- 16.8** The Common Ion Effect and Solubility 6.B.1 6.B.2 6.C.3
- 16.9** pH and Solubility 6.B.1 6.B.2 6.C.3
- 16.10** Complex Ion Equilibria and Solubility 6.B.1 6.B.2 6.C.3
- 16.11** Application of the Solubility Product Principle to Qualitative Analysis 6.B.1 6.B.2 6.C.3

### Introduction

Acid meets base in Chapter 16—and when they meet, the acid will neutralize the base, producing water. By noting the changes in pH during this neutralization reaction, the strengths of acids and bases may be determined. (EK.6.C.1). In another reaction, a weak acid or base may be matched with an equal volume and concentration of a soluble salt of its anion forming a solution that maintains a constant pH (EK.6.C.2). This solution is called a “buffer” solution. Buffers are vital in maintaining the narrow pH range needed for blood to efficiently deliver oxygen and remove carbon dioxide from cells. Buffer solutions can be understood through equilibrium principles presented earlier in the course.

The last half of Chapter 16 is the study of slightly soluble compounds in equilibrium with their aqueous ions. Commonly known as saturated solutions, these systems show a balance between the dissolving solid and the precipitating ions. This balance is described through an equilibrium expression called the solubility product or  $K_{sp}$  (EK.6.C.3). There are conditions under which the solubility of a compound may be altered. An example of this is the mineral clogging a coffee maker causing it to overheat. This compound is the slightly soluble mineral, calcium carbonate. Le Châtelier’s principle, which was introduced in Chapter 14, applies to the solubility of the calcium carbonate. A mild acid such as vinegar will shift the equilibrium, increase the solubility of the calcium carbonate, and help the coffee maker run more efficiently. Knowing chemistry makes life more understandable!



# Chapter 17

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## Entropy, Free Energy, and Equilibrium



The laws of thermodynamics set an upper limit on how much heat can be converted to work, as in the case of an internal combustion automobile engine.

### Chapter Contents

- 17.1** The Three Laws of Thermodynamics 5.B.2
- 17.2** Spontaneous Processes 5.E.2 5.E.5
- 17.3** Entropy 5.E.1
- 17.4** The Second Law of Thermodynamics 5.E.1 5.E.2
- 17.5** Gibbs Free Energy 5.E.2 5.E.3 5.E.4 5.E.5
- 17.6** Free Energy and Chemical Equilibrium 6.D.1
- 17.7** Thermodynamics in Living Systems 5.E.4

### Introduction

In the natural world, two fundamental forces are responsible for chemical and physical changes. One is the tendency for a system to become reduced to its lowest internal energy level. Consider what happens when a ball rolls down an incline. As the ball rolls downward, its potential energy is reduced, being converted into kinetic and heat energy. A second force, perhaps more subtle, is the tendency for matter to become more disorganized. If solid sugar is added to warm water, the sugar will dissolve. In this case, the solid sugar and warm water, before they are mixed, are more organized forms of matter. The dissolved sugar water is a more disorganized form of matter.

These same forces are also at work in chemical systems, determining whether a reaction will occur, in what direction, and to what extent. The tendency for a system to reduce its internal energy by the release of heat is called enthalpy change, previously introduced in Chapter 6. The tendency for a system to increase its randomness, or disorder, is called entropy change (EK.5.E.1). Along with temperature, these forces are part of a function called Gibbs free energy, which can be used to determine the spontaneity of a reaction (EK.5.E.2-EK.5.E.4). The study of enthalpy, entropy, and free energy is included in a subject called thermodynamics. While enthalpy, entropy, and temperature may be used to predict the spontaneity of a reaction, they do not provide information on activation energy or rate of a reaction (EK.5.E.5). For example, in the case of a burning candle, enthalpy, entropy, and temperature predict that the reaction will be spontaneous and strongly favor products. However these factors do not predict activation energy—in this case in the form of a lighted match—necessary to initiate the reaction so that the candle can burn. The study of those factors necessary for a reaction to occur spontaneously will be included in this chapter.

# Chapter 18

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## Electrochemistry



Michael Faraday at work in his laboratory. Faraday is regarded by many as the greatest experimental scientist of the nineteenth century.

### Chapter Contents

- 18.1** Redox Reactions 3.B.3
- 18.2** Galvanic Cells 3.B.3 3.C.3
- 18.3** Standard Reduction Potentials 3.B.3 3.C.3
- 18.4** Thermodynamics of Redox Reactions 3.C.3
- 18.5** The Effect of Concentration of Cell Emf
- 18.6** Batteries 3.B.3 3.C.3
- 18.7** Corrosion 3.B.3
- 18.8** Electrolysis 3.B.3 3.C.3 5.E.4

### Introduction

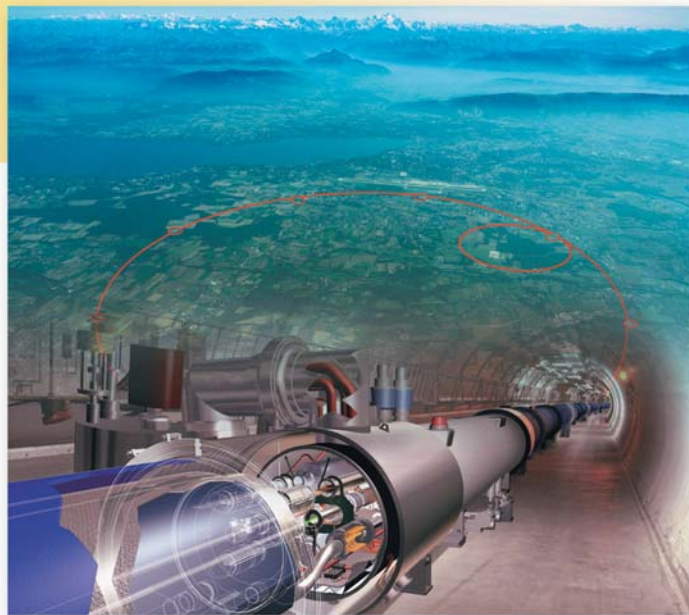
To say our world runs on batteries is an understatement. In one day, you might rely on battery-powered devices to communicate with your friends and family, to play video games, to compose a lab report, and to drive or ride to school. The energy powering these devices is taken for granted until a variety of signals reminds you that the stored energy in the battery is diminishing. Thus, one of our daily chores may be to recharge those batteries by plugging our phone, computer, or perhaps hybrid powered car into an electrical outlet. All batteries undergo chemical reactions that generate electrical energy. If a battery is rechargeable, the reactions must be reversible to allow reactants to be separated so that electrons exchanged will flow through an external circuit and perform useful work. All batteries are variations of a Galvanic cell, where chemical energy is converted into electrical energy (EK.3.C.3).

For some oxidation-reduction reactions to occur, the reverse of a Galvanic cell or battery is required. In this case, electrical energy must be converted into chemical energy. The electrical energy may be provided by a battery, which delivers electrons to the system, causing a nonspontaneous reaction to occur. This process is called electrolysis and occurs in an electrolytic cell. Electrolysis is commonly used to plate one metal onto another—as in chrome plated car parts and plumbing fixtures. The thickness of metal plating may be determined by considering Faraday's Law—which relates electrical current and time to moles of electrons (EK.3.C.3). The interconversion of electrical and chemical energy is the focus of this chapter.

# Chapter 19

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## Nuclear Chemistry



The Large Hadron Collider (LHC) is the largest particle accelerator in the world. By colliding protons moving at nearly the speed of light, scientists hope to create conditions that existed right after the Big Bang.

### Chapter Contents

- 19.1 The Nature of Nuclear Reactions
- 19.2 Nuclear Stability
- 19.3 Natural Radioactivity 4.A.3
- 19.4 Nuclear Transmutation
- 19.5 Nuclear Fission
- 19.6 Nuclear Fusion
- 19.7 Uses of Isotopes
- 19.8 Biological Effects of Radiation

### Introduction

With the exception of radioactive decay covered in section 19.3 of this chapter, topics of nuclear chemistry are beyond the scope of the redesigned AP Chemistry curriculum and the new AP Chemistry exam.

Radioactive decay of isotopes is an example of first-order reaction kinetics. As with all first order reactions, the half-life of an isotope is independent of its concentration or starting amount and constant in value. Half-life of an isotope is an intrinsic property. For example, the half-life of the nuclear fuel, Plutonium-239 is 24,400 years, indicating that any amount of this dangerous material would decay at an extremely slow rate. Knowledge of isotopic decay provides a context for understanding the safety concerns surrounding the use and disposal of radioactive material (EK.4.A.3.e).

Other topics of nuclear chemistry while not covered in the AP curriculum are still relevant in today's world. These topics include carbon dating, diagnostic imaging in medical treatments, irradiation of food, and the production of atomic power. Such topics could be covered during the time following the AP exam.

# Chapter 20

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## Chemistry in the Atmosphere



Lightning causes atmospheric nitrogen and oxygen to form nitric oxide, which is eventually converted to nitrates.

### Chapter Contents

- 20.1** Earth's Atmosphere
- 20.2** Phenomena in the Outer Layers of the Atmosphere
- 20.3** Depletion of Ozone in the Stratosphere
- 20.4** Volcanoes
- 20.5** The Greenhouse Effect
- 20.6** Acid Rain
- 20.7** Photochemical Smog
- 20.8** Indoor Pollution

### Introduction

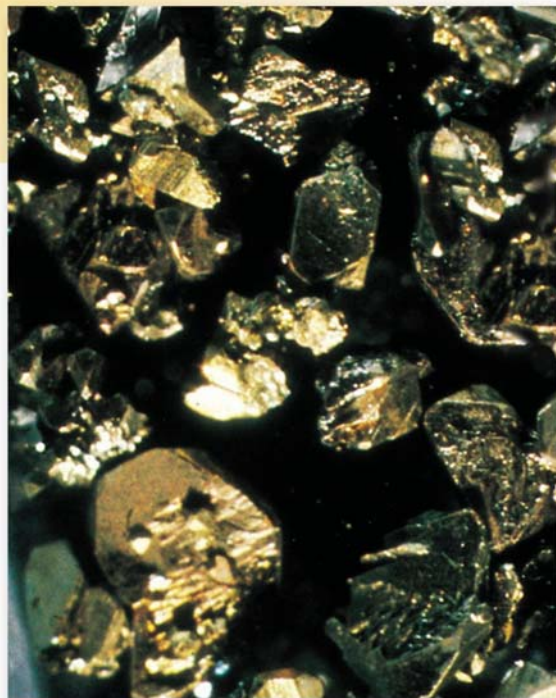
In a direct sense, Chapter 20 is beyond the scope of the redesigned AP Chemistry curriculum and the new AP Chemistry exam.

However, several topics within the chapter may be used as practical applications for AP exam questions focusing on the six Big Ideas of the new Curriculum Framework. For example, in section 20.3, reaction mechanisms are used to describe the processes behind ozone depletion in the stratosphere—relating back to Chapter 13 “Chemical Kinetics.” In sections 20.1 and 20.5, conservation of atoms is illustrated by the global cycles for nitrogen, oxygen and carbon—relating back to Chapter 3 “Mass Relationships in Chemical Reactions.” Acid/base interactions—introduced in Chapter 15 “Acids and Bases”—are used in section 20.6 to explain the formation and effects of acid rain. These provide important illustrations of earlier concepts—all of which are part of the new AP Chemistry redesigned curriculum.

# Chapter 21

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## Metallurgy and the Chemistry of Metals



Crystals of salt composed of sodium anion and a complex sodium cation with an organic compound called crown ether.

### Chapter Contents

- 21.1** Occurrence of Metals
- 21.2** Metallurgical Processes
- 21.3** Band Theory of Electrical Conductivity 2.D.2 2.D.3
- 21.4** Periodic Trends in Metallic Properties 1.C.1
- 21.5** The Alkali Metals 1.C.1
- 21.6** The Alkaline Earth Metals 1.C.1
- 21.7** Aluminum 3.C.3

### Introduction

One of the strengths of chemistry is the ability to make connections between what is observed on the visible level with what occurs on the invisible or particle level. Making these connections is a point of emphasis in the redesigned AP Chemistry curriculum. An example of this is the ability of silicon to conduct electricity. This characteristic of silicon can be used to make computer “chips” with miniature circuits that can operate handheld devices such as calculators and electronic tablets. How can silicon, with bonding similar to diamond, which is non-conducting, have this ability? Section 21.3 will provide an answer to this question and explain why metals conduct electricity while other materials, such as wood and glass, do not ([EK.2.D.2,3](#)).

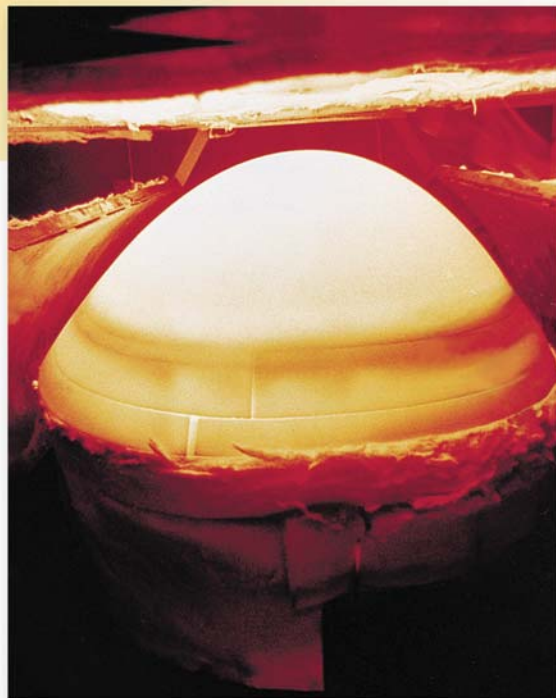
A summary of trends in metallic properties is provided in section 21.4 ([EK.1.C.1](#)). The challenge for students of AP Chemistry is to connect such properties to the arrangement and interaction of particles in the substance.

All other sections in Chapter 21 are beyond the scope of the redesigned AP Chemistry course and the new AP Chemistry exam.

# Chapter 22

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## Nonmetallic Elements and Their Compounds



The nose cone of the space shuttle is made of graphite and silicon carbide and can withstand the tremendous heat generated when the vehicle enters Earth's atmosphere.

### Chapter Contents

- 22.1 General Properties of Nonmetals 1.C.1
- 22.2 Hydrogen
- 22.3 Carbon
- 22.4 Nitrogen and Phosphorus
- 22.5 Oxygen and Sulfur
- 22.6 The Halogens 1.C.1

### Introduction

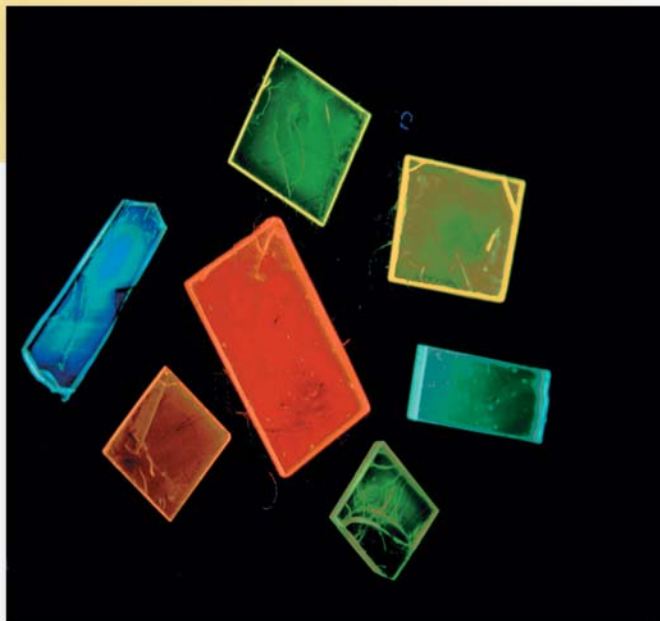
Chapter 22 is a comprehensive study of nonmetals and their compounds. The elements, hydrogen, carbon, nitrogen, oxygen, phosphorus, sulfur, and the halogens are examined. While the content of this chapter is generally beyond the scope of the redesigned AP Chemistry course, there are sections that include periodic nonmetal properties, which provide specific examples and enrich earlier chapters—particularly Chapter 8. Section 22.1 describes the general properties of nonmetals while section 22.6 and table 22.4 present the properties of the halogens (EK.1.C.1).

Memorization of these properties is not required for the redesigned AP Chemistry course. What is expected is that you are able to link properties to the particles involved and to the bonds that hold them together. As an example: the boiling points of the elemental halogens increase when moving down group 7 from fluorine to iodine. This can be explained considering intermolecular bonds and diatomic molecules—topics found in Chapter 11. Moving from the visible to invisible, from property to particle, is an important intellectual transition you must make when studying AP Chemistry!

# Chapter 23

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## Transition Metals Chemistry and Coordination Compounds



Copper ions implanted in  $\text{Al}_2\text{O}_3$  emit visible radiation when excited by UV light. The color of light can be changed by adding other elements in small amounts.

### Chapter Contents

- 23.1** Properties of the Transition Metals 1.C.1
- 23.2** Chemistry of Iron and Copper
- 23.3** Coordination Compounds
- 23.4** Structure of Coordination Compounds
- 23.5** Bonding in Coordination Compounds: Crystal Field Theory
- 23.6** Reactions of Coordination Compounds
- 23.7** Applications of Coordination Compounds

### Introduction

The properties of transition metals, presented in section 23.1, provide more examples of the principle of periodicity—originally introduced in Chapter 8 (EK.1.C.1). A more detailed study of the coordination compounds and complex ions formed by transition metals is beyond the scope of the redesigned AP Chemistry course.

As enrichment, this chapter offers explanations for the variety of colors displayed by coordination compounds. Some colors for first-row aqueous transition metal ions are shown in figure 23.20. In this picture, you may recognize the sky blue color of the aqueous copper ions or the pink of aqueous cobalt ions. Section 23.5 will explain why these ions display colors while other aqueous ions are colorless.

# Chapter 24

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## Organic Chemistry



A chemical plant. Many small organic compounds such as acetic acid, benzene, ethylene, formaldehyde, and methanol form the basis of multi-billion-dollar pharmaceutical and polymer industries.

### Chapter Contents

- 24.1 Classes of Organic Compounds
- 24.2 Aliphatic Hydrocarbons
- 24.3 Aromatic Hydrocarbons
- 24.4 Chemistry of the Functional Groups

### Introduction

The specific content of organic chemistry covered in this chapter is beyond the scope of the redesigned AP Chemistry course. However, carbon compounds can be used in AP Chemistry exam questions that are testing *other* concepts. As an example, for a question testing knowledge of intermolecular bonding, the question may provide structural diagrams for two organic molecules—methanol and propanol. Considering intermolecular bonding, you may be asked to explain why methanol boils at a higher temperature than propanol. Another bonding question may provide a structural diagram of a compound containing two carbons, doubly bonded together. You may be asked to identify the orbitals used in the double bond.

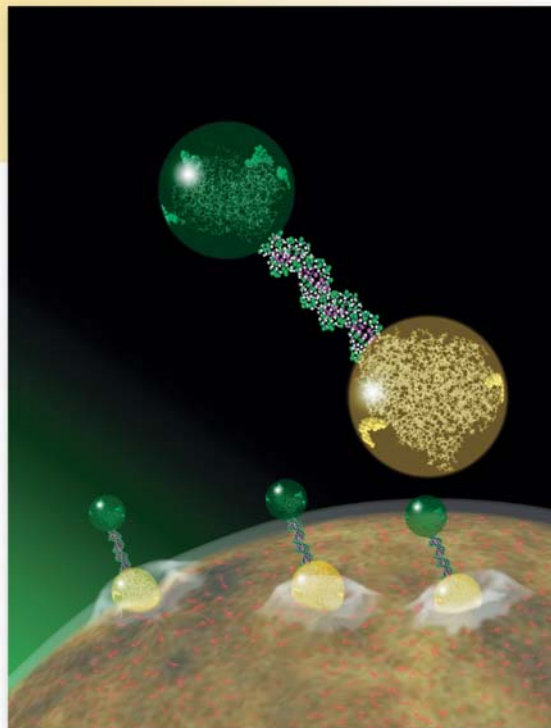
Both of these examples call for an application of principles learned earlier in the course with the carbon compounds serving as “case studies.” Consider Chapter 24 as a source of information for the organic or carbon based compounds you encounter during the course.



# Chapter 25

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## Synthetic and Natural Organic Polymers



University of Michigan researchers have developed a faster, more efficient way to produce nanoparticle drug delivery systems, using DNA molecules to bind the particles together.

### Chapter Contents

- 25.1 Properties of Polymers
- 25.2 Synthetic Organic Polymers
- 25.3 Proteins 2.B.2 5.D.3
- 25.4 Nucleic Acids 2.B.2 5.D.3

### Introduction

Earlier chapters identified the bond as an attraction between positive charges and negative charges. The positive charge may be from a nucleus, a cation, or an atom that has had its electron density shifted and is now fractionally positive. The negative charge may be from electrons, an anion, or an atom that has had its electron density shifted and is now fractionally negative. Regardless of the origin of positive and negative charges, once they are near each other, they will bring particles together through mutual attraction.

In Chapter 25, the particles involved are proteins and nucleic acids—both of which are long chained structures called polymers. These particles are large, stable structures and can have complex shapes. Bonds can provide connections to make the polymers as well as lead to unique shapes—such as the double helix of DNA which provides stability for the structure. The attraction of negative for positive, which creates these bonds and is crucial to life on Earth, is illustrated in this chapter (EK.5.D.3).