

RESOURCE GUIDE



An Introduction To Chemistry



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Academic Decathlon®

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Introduction

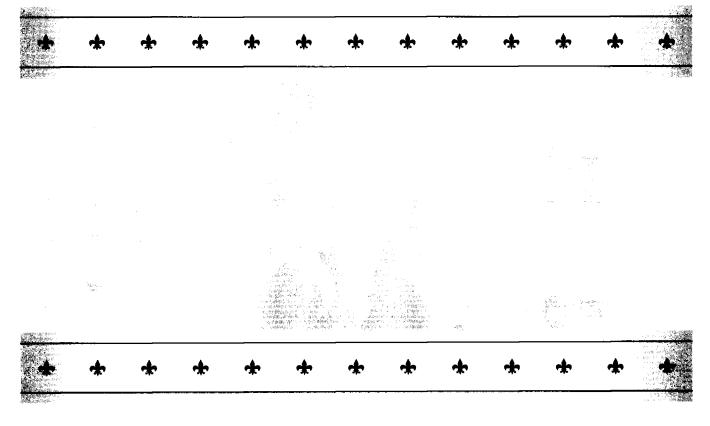
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his Science Resource Guide provides an overview of important concepts in the discipline of chemistry. In Section I, you will find a very brief overview of the history of the development of chemistry. This overview includes a particular focus on the work of Antoine Lavoisier (1743–94), who is sometimes referred to as the "father of modern chemistry." This focus is in keeping with the overall curricular theme, as Lavoisier's life was greatly impacted by the French Revolution. At the time of the French Revolution, the field of science was changing rapidly as an increasing number of people began performing scientific experiments and using what later became known as "the scientific method" to think about nature and the results of those experiments.

In Section II of the resource guide, you will learn about chemists' theories for the way matter is arranged and held together by the forces of nature. Only a relatively few simple particle types are needed to build the enormous variety of substances found in nature and manufactured by chemists. Section III of the resource guide explores how substances exist in different phases as gases, liquids, or solids depending on variables, such as the temperature and pressure. You will also learn about the laws and theories that govern how substances change from one phase to another. And, you will learn that the basis for our understanding of chemistry is a model that has small particles moving, vibrating, or rotating. You will also learn about simple ways that substances can mix together in solution.

Chemistry is a science of change. Section IV deals with how chemicals interact with one another to make new structures and substances. Section IV will review the laws and principles that govern chemical reactions, and you will learn how to use these laws and principles to predict what will happen in certain reactions and how fast these reactions will occur.

Note: Throughout the resource guide you will notice that some terms have been boldfaced. These boldfaced terms can be found in the glossary at the end of the guide.





Section I:

A Brief History of Chemistry

Introduction

o one really can say precisely how and when the science of chemistry was first practiced. However, it is safe to say that the first time someone started a fire and observed what happened to materials such as meat put into the fire, they were observing chemistry. Furthermore, when they observed that the same thing always happened to a type of object placed into the fire, they were discovering one of the basic principles of scientific experimentation—that experiments must be able to be repeated exactly. Metals were found to form when certain rocks were heated, and colored solutions were found when plants were soaked in water. These colored solutions could be used to dye fibers in clothing or human skin, as with tribal tattoos. In each of these cases, early cultures were discovering and using chemistry. As early people took this information and applied it to new situations—such as experimenting with the effects of heat on new substances or testing new plants for food or dyes—they were beginning to develop and exchange the facts and concepts of chemistry.

Occurring independently in several different places on Earth, these discoveries and inventions eventually developed into a remarkably large set of practical skills based, essentially, on a growing body of knowledge about the properties of materials and the ways these materials could be manipulated. The study of the properties of and changes in materials is central to the study of chemistry. Over a long history, humans have applied chemistry for peaceful purposes as well as for the purpose of making war.

Eventually, a relatively small group of people decided that there was indeed order in the seemingly vast array of behaviors of materials, and early philosophers began to develop the theories of chemistry and science in general as they asked and sought to answer the following question: How does nature work? These early ideas were based on the best information available at the time—before any scientific instruments or methods had been properly formulated.

One significant factor that prohibited the advancement of science was the lack of good means to communicate with others regarding discoveries. A result of this lack of communication was the fact that vastly different approaches to chemistry were formulated in different parts of the world. Eventually, means of storing and communicating

information, such as books, conferences, scientific and philosophical societies, and so on, were developed and greatly aided the advancement of chemistry and science in general. Nowadays, the Internet is the newest and fastest way of sending and retrieving information about chemistry.

So much has happened in the history of chemistry—starting slowly over thousands of years and then escalating in spurts, with the most recent such "spurt" starting around 1600 A.D. and continuing to the present—that a lifetime could be spent exploring the development of chemical knowledge and understanding. To complicate matters, historians rightly argue over whether independent discoveries were made or whether ideas exchanged and brought from "abroad" started the spread of new technologies and improved humans' understanding of nature. Given the vastness of chemical discoveries, this resource guide will highlight just a few brief episodes in the history of chemistry.

Chemistry Prior to the Scientific Revolution

The first known metal was probably gold. Ornaments made of gold have even been found with stone tools from the **Neolithic age**. Copper may have been the next metal known and utilized by humans, and it may have been the earliest to be smelted from ores in Egypt and **Mesopotamia**. Tin and **bronzes** (a mixture of tin and copper) soon followed and were used for adornment, displays of wealth, and the manufacturing of weapons. Other elements that were sometimes incorporated were silver, antimony, and lead. Descriptions and pictorial representations of metalworking are common in remains from the early dynasties in Egypt. Zinc and **brass** (a mixture of zinc and copper) were known in pre-Roman Empire times. From the earliest times, both raw materials and metal artifacts formed the basis for the beginnings of trade and commerce.



Steel is a metal alloy with iron as its major component.

PHOTOGRAPH BY JOHANNES HEMMERLEIN.



MIXING METALS TO MAKE BRONZE

Bronze was one of the earliest alloys to be made and used. Which elements have been used to make bronzes? What different methods have been used for the manufacturing of bronze artifacts by different cultures over the years?



Very little iron occurs alone as an element in nature, and so iron had to be made by heating ore. The first iron to be used may have been from meteorites. Around 1700 B.C., iron was mixed with carbon and other impurities to form an **alloy** called **steel**. Steel made especially effective weapons and therefore played an important role in major political upheavals.



THE USE OF DYES AND PRESERVATIVES

People are fascinated by color. In your independent research, you should find out how colors were extracted and used in the form of dyes, such as "Tyrian purple," and preservatives. In what ways were dyes and preservatives used to assist mummification, reform textiles, and for artwork?



Glass and pottery have been made for thousands of years by heating mixtures of sand or clay. Remains of glass and pottery "factories" have been dated to thousands of years B.C., and research indicates that pottery and glass were widely traded and transported. Ceramics and glasses had both utilitarian and decorative uses. Ceramic containers were important for storing and shipping oils and wines in many cultures in both the old and new worlds.



ALCHEMY AND THE "PHILOSOPHER'S STONE"

In your independent research, you should discover why chemistry entered a period that focused on a discipline known as "alchemy." A popular substance sought by alchemists was the "philosopher's stone." What was the "philosopher's stone"?



In your independent research, you should find out where gunpowder and fireworks were invented. For what purposes were gunpowder and fireworks used when these inventions became more widely known?





EARLY THINKERS ON THE NATURE OF MATTER

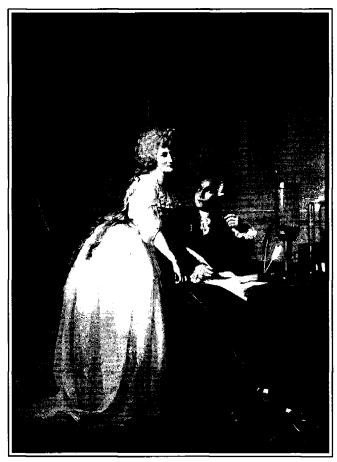
The famous Greek philosophers Aristotle and Democritus thought deeply about the world in which they lived. How did these philosophers explain the nature of matter?



Antoine Savoisier and the Birth of Modern Chemistry

Modern chemistry developed when individuals began to make careful observations of matter and began to develop instruments to help keep track of all matter, including solids, liquids, and gases. Gases were the easiest to understand but the hardest to trap and measure. The importance of chemistry's applications to the science of medicine provided a crucial impetus for the development of modern chemistry, Johann Baptista van Helmont (1579–1644) was a key Belgian noble who introduced the term "gas," developed an effective balance, was probably the first to identify and demonstrate the law of conservation of mass, and provided several refutations of the notion of alchemy and element transmutation.12 Van Helmont's ideas were at first challenged by leading physicians, but his ideas were supported and followed closely by an expanding group of non-medical scientists working in France and Britain, notably Robert Boyle (1627–91), Joseph Black (1728–99), Henry Cavendish (1731-1810), Joseph Priestly (1733-1804), and Antoine Lavoisier (1743-94).

These scientists, and many others, were beginning the modern practice of operating research laboratories, producing more frequent publications, and encouraging the exchange of ideas. Several even visited one another to observe each other's laboratory experiments. Much of their practice and discussion focused on reactions in air, as they tried to explain the results of reactions of metals and carbon compounds heated in air—reactions that are now known as "oxidation." Another fascination of these scientists had to do with how a tree acquires its mass. The false conclusions they drew on this topic continue to be commonly held misconceptions to the present day.



Portrait of Antoine Lavoisier, widely considered the father of modern chemistry, and his wife, chemist Marie-Anne Pierrette Paulze.

PORTRAIT BY JACQUES-LOUIS DAVID.





The classic "living tree" experiment was performed by Jan Baptiste van Helmont, and others repeated his experiments. Your research should find out what the experiments involved, how early scientists interpreted their results, and what we now believe the results mean.

* * * * * * * * * *

Antoine Lavoisier was able to pull the work of many scientists together to create a more coherent picture of what the "rules" of chemistry are. Lavoisier made numerous influential findings—he discovered that air is composed primarily of oxygen and nitrogen; that oxidation involves a reaction with oxygen; that water cannot be converted into Earth; and that matter is truly conserved in chemical and

physical interactions. Lavoisier also began (with others) to systematize the naming of chemical compounds. Lavoisier's great findings have held true over the course of time.

By the early eighteenth century, the French scientific community was at a peak of excellence, especially in chemistry and mathematics. As new ideas were coming to the fore in the world of science, so too were new ideas gaining ground in the political and social realms. In England, these new ideas led to institutional reform. In other parts of Europe and in America, they led to revolution. As revolution swept France in 1789, French chemists rallied to assist the new government with better tools for warfare. However, political change in France and an obsession with nature as a moral model led the Jacobins—a radical revolutionary group---to attempt to suppress science (other than biology) and scientists as a "stubborn bastion of aristocracy." The Jacobins accused scientists of "drawing a veil of obscurity between nature and the people." In August 1793, the National Convention, dominated by the Jacobins, abolished the Academy of Sciences, charging that the institution was incompatible with a republic. Lavoisier himself became a key target.

ANTOINE LAVOISIER AND HIS ROLE AND FATE IN THE FRENCH REVOLUTION

Along with his scientific efforts, what was Antoine Lavoisier's role and fate in the French Revolution?

MADAME LAVOISIER

How was Madame Lavoisier involved in her husband's chemistry work? What was her fate during and after the French Revolution?

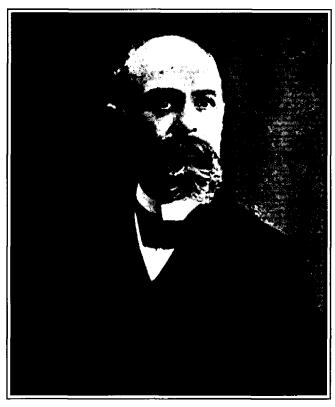
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Chemistry after Lavoisier

A large amount of progress concerning the chemical and physical understanding of matter had been accomplished soon after the time of Antoine Lavoisier. By 1900, enormous numbers of chemical elements and compounds had been isolated or manufactured, and their properties had been examined and documented. One major part of chemistry that remained to be discovered was atomic and nuclear structure. An understanding of atomic and nuclear structure began to come to light with Henri Becquerel's (1852–1908) discovery in 1896 that uranium compounds were **radioactive**.⁴

A rapid period of both experimental and theoretical developments followed. These developments included the explanation of **radioactive decay** by Pierre (1859–1906) and Marie (1867–1934) Curie, the discovery of the **electron** by J. J. Thomson (1856–1940), the discovery of the atomic **nucleus** by Ernest Rutherford (1871–1937), the discovery of the **neutron** by James Chadwick (1891–1974), the discovery of electron **orbitals** by Niels Bohr (1885–1962), and the discovery of **isotopes** by Frederick Soddy (1877–1956).

Following his successful explanation of how **photons** are able to release electrons from a metal surface in the **photoelectric effect**, Albert Einstein (1879–1955) proceeded to postulate that all matter and energy are interchangeable and can be quantified by a single equation $E = \text{mc}^2$, where E is energy, m is mass, and c is the speed of light. Further wave-particle ideas and discrete **quantum** energy relationships were introduced by Louis de Broglie (1892–1987) and Erwin Shroedinger (1887–1961). These ideas were closely tied to mathematical models that could be used to compute the energies of atomic states (and eventually nuclear states).



Picture of Henri Becquerel, the French physicist who discovered radioactivity.

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RADIOACTIVITY AND NUCLEAR STRUCTURE

How did experiments by Ernest Rutherford, James Chadwick, Antoine Henri Becquerel, and Pierre and Marie Curie help to establish correct ideas about radioactivity and nuclear structure?



Despite the tremendous progress that had been made, a great deal of effort was still needed to provide the theoretical and predictive underpinnings to the enormous amount of information collected about chemicals. One such organizer was the **periodic table**. Another significant advancement was an improved understanding of the forces between atoms joined in chemical compounds (known as **intramolecular forces**) and the forces between such compounds (**intermolecular forces**) that appear to govern the physical properties of substances, such as boiling and melting temperatures.

Gilbert N. Lewis (1875–1946) and Linus Pauling (1901–94) formulated ideas about how atoms formed bonds as a result of two electrons pairing to hold the adjacent nuclei together by means of **electrostatic forces** between opposite charges. When a positive and a negative charge are located a small distance apart, such as at the ends of a chemical bond, this creates an electric **dipole**. On two neighboring molecules, these dipoles can line up with opposite charges adjacent, thus creating a force of attraction. Intermolecular forces result from permanent and/or **temporary dipoles** arising from positive and negative charge separation in molecules.



THE PERIODIC TABLE AND ASSOCIATED PERIODICITY

Dmitri Mendeleev was a Russian scientist who decided that all of the identified chemical elements could be organized in a logical way. What characteristics did Mendeleev use to organize the chemical elements into a table? What did he do about elements that didn't seem to fit into the table?



The recognition of the interrelationship of heat and mechanical motion, or work, as well as the concept of temperature was a development that greatly improved scientists' understanding of energy relationships in chemistry. This development came from industry, weapons development, and physics. Benjamin Thompson (1753–1814) (later titled Count Rumford), an American scientist working in Europe, had measured the heat from friction of boring iron to make a cannon. Through several experiments James Joule (1818– 89) established the exact relationship between mechanical work and heat by using measurements of mechanical work and determining how it changed the temperature of water by precise amounts. Many scientists applied this concept to energy changes in chemical reactions. Another American scientist, Josiah Willard Gibbs (1839-1903), formulated the precise relationships that enable chemists to predict the probability that a chemical reaction will or will not occur. All that was needed to determine this was a knowledge of the values of fundamental measurable parameters of the reactants and products.

The field of **chemical kinetics**—which studies factors that affect the rate at which a chemical reaction occurs—has seen developments in both experimental and theoretical work, mostly in the last century. Mathematical models can now be used that perceive a chemical reaction as a gradual motion of atoms on a chemical potential energy graph extending from reactants to products. The collection of atoms must overcome an energy barrier on its travel, and the height of that barrier essentially establishes the maximum rate at which the reaction can occur.

Section 1 Summary

- Chemistry began with early practical applications of humankind's ability to manipulate materials found in nature.
- Chemistry progressed to a period where careful experiments, involving gases and the measurement of mass and energy, identified key principles and laws according to which materials changed. This period constituted a Scientific Revolution.
- More recently, mathematical models have helped chemists understand and explain the nature of matter, its structure, and its ability to change. These models are now used by thousands of chemists worldwide to predict the results of chemical and physical changes.
- Throughout history, key individuals have contributed special insights that have enabled progress to be made in understanding the nature of matter.



Section II:

The Structure of Matter

Introduction

n this section of the resource guide, we will review how matter is structured and holds together. All matter is made up of a few fundamental particles, and tremendous variety occurs in nature because of the many different ways in which these particles can be arranged. After we discuss these fundamental particles and learn about how their properties differ, we will examine the forces that can hold them together in the different ways. We also will begin to explore how just a few basic bonding principles can explain how almost every kind of matter is held together. Understanding these principles can enable us to predict both structures and properties of materials. We will also see

DETECTION
Faraday collectors

Mase 45

how chemists have learned to organize enormous amounts of information into tables that summarize properties by emphasizing trends rather than specific information.

Atomic Theory and Atomic Structure

INTRODUCTION

Atoms are the smallest distinctive particles that make up matter. Each of the approximately one hundred chemical **elements** is made up of atoms that have a distinctive structure and mass. The internal structure of the atom determines the element's chemical behavior. This internal atomic structure is made up of positive, negative, and neutral particles. Atoms of identical structure all behave the same. Usually atoms have a neutral charge; however, atoms can lose or gain one or more electrons to form, respectively, positive or negative **ions**.

DETERMINATION OF ATOMIC MASSES

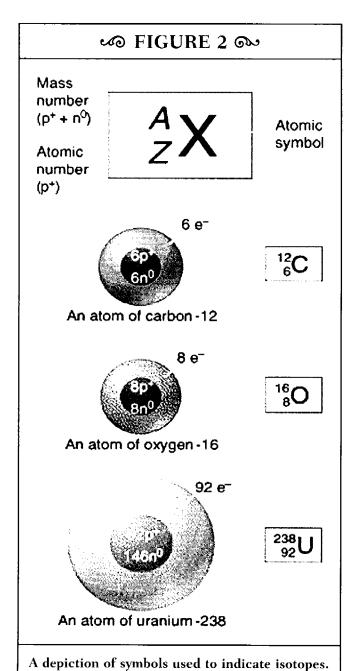
Generally, a single atom cannot be weighed. However, by measuring a large number of identical atoms simultaneously, chemists can calculate the mass of one atom. Chemists determined these calculations by carefully measuring the relative masses of atoms that combine to form compounds. For example, the mass of hydrogen that combines with a mass of fluorine to form the compound HF was found to be in the ratio 1:19 (1 gram of hydrogen always combines with 19 grams of fluorine). By measuring many ratios, chemists were able to establish a table of relative **atomic masses**. Then, by fixing a standard mass, they were able to construct a complete mass table for all of the chemical elements.

In modern laboratories, using instruments called **mass spectrometers**, chemists can determine the relative masses of single atoms quite accurately. Mass spectrometers can also measure the masses of atoms bonded in compounds. Mass spectrometers are even used at airports to check that passengers are not carrying compounds that are typical explosives. The schematics of a simple mass spectrometer are shown in Figure 1.



One way that chemists identify atoms and compounds is by measuring their mass. In modern laboratories this is done by using an instrument called a mass spectrometer. In your independent research, you should discover how mass spectrometers measure the masses of atoms and compounds.





ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES

It is important for chemists to be able to identify and communicate the exact composition of an atom. Since each chemical element has different but unique types of atoms, the use of a standardized identification symbol is important. Atoms consist of a positively charged central **nucleus** and surrounding negative **electrons**, symbol e^- . The nucleus always contains at least one positive **proton**, symbol p^+ , and, except for hydrogen, one or more **neutrons**, symbol n^0 .

The Use of Symbols for Isotopes

The symbol format used by chemists is ${}_{Z}^{A}X$, where X is the chemical symbol, Z is the atomic number or number of protons in the nucleus of the atom, and A is the mass **number**, which is equal to the number of neutrons plus the number of protons. Although the proton number can't change for an atom, atoms of a given element can have different numbers of neutrons in their nucleus. The same type of atom with a different mass number is called an isotope of the element. Any arrangement of protons and/or neutrons forming a nucleus is called a nuclide. The mass number minus the number of protons will indicate the number of neutrons. For example, carbon has three isotopes that are found in nature: ${}_{6}^{12}$ C, which has six (12 - 6 = 6) neutrons in its nucleus, ${}_{6}^{13}$ C, which has seven (13 - 6 = 7) neutrons, and ${}_{6}^{14}$ C, which has eight (14 - 6 = 8) neutrons. ${}_{6}^{14}$ C atoms are unstable and decay over time and are therefore classified as **radioactive**. Figure 2 shows the use of symbols to indicate isotopes.

Standard for Atomic Mass: 12C = 12.00000

Originally, hydrogen was established as having a mass of 1, but then it was found more convenient to establish the mass of a carbon atom as equal to 12 mass units (named **atomic mass units**, or amu). This only made a slight change to the relative scale, but solid carbon was easier to use for comparison purposes than hydrogen gas. Hydrogen, with a mass of 1, is 1/12 the mass of a carbon. Helium, with a mass of 4, is 1/3 the mass of carbon.



PROPERTIES AND IMPORTANCE OF COMMONLY RECOGNIZED ISOTOPES-2H, 14C, 60Co

Isotopes are important in chemistry, life sciences, and earth sciences. In your independent research, you should discover where ²₁H, an isotope of hydrogen, ¹⁴₆C, an isotope of carbon, and ⁶⁰₂₇Co, an isotope of cobalt, are found as well as how they

are made. How are they used? What properties of these isotopes have made them useful?



Calculating Weighted Averages

When the mass of atoms found in nature is determined by measuring the masses of elements combining into compounds, enormous numbers of atoms have to be used to make the measurement. As a result, when several isotopes of the element exist, the mass determined is a weighted average of the masses of the isotopes. Calculating the weighted average atomic masses for a simple binary-isotopic element such as chlorine from masses and abundances of ¹⁵Cl and ¹⁷Cl involves several steps.

How this works for an element with two isotopes, such as natural chlorine, which has 75 percent $^{35}_{17}\text{Cl}$ atoms and 25 percent $^{37}_{17}\text{Cl}$ atoms, is shown in the following formula:

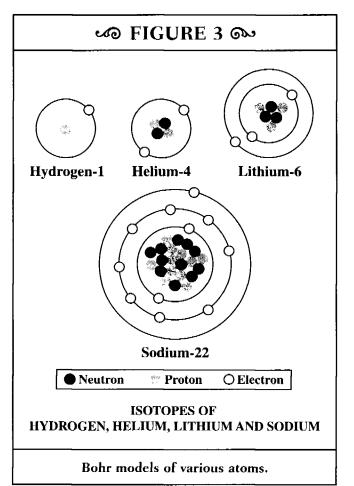
weighted average mass of a chlorine atom = $(0.75 \times \text{mass} \frac{35}{17}\text{Cl atom}) + (0.25 \times \text{mass} \frac{35}{17}\text{Cl atom})$

Since 75 percent of the atoms have a mass of 35, we know that the majority of the atoms have a mass of 35, so the average mass will be closer to 35 than to a mass of 37. The accepted value of the weighted average atomic mass for chlorine found in nature is 35,453 amu.

ELECTRON ENERGY LEVELS AND EMISSION AND ABSORPTION ATOMIC SPECTRA

Electrons surround the nucleus of every atom. The total number of electrons equals the number of protons (atomic number, Z), but electrons must occupy specific energy levels in the atom. To determine these energy levels, chemists excite electrons. When atoms are heated strongly, electrons can be excited, or raised to a higher energy level, by the absorption of a **quantum** of light energy (**photon**). When the atom returns to a lower energy state from the excited state, a photon can be emitted. When the absorbed or emitted photons are examined for their energy, the result is known as an **absorption or emission spectrum**.

By measuring the exact energies of the photons, the differences between the energy levels of the electrons can be determined. In this way, the scientist Niels Bohr and others established a detailed model of the electronic structure of atoms that matched the observed spectra. This model is known as the **Bohr Model of the atom**. (See Figure 3.) In the Bohr atom, the negatively charged electrons occupy fixed (meaning apparently unchanging) orbits around the central positively charged nucleus. Each orbit is at a certain distance from the center of the atom and provides the electron-nucleus charged pair with a certain potential energy. No more than two electrons can occupy the same exact orbit, and even these two would have very slightly different properties. If an electron can move to an orbit closer to the nucleus, the emission of light will occur at an energy corresponding to the potential energy difference in the atom. The absorption of light results in the electron moving further from the nucleus to a higher energy condition.



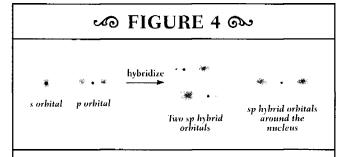
Scientists also came to recognize that all matter appeared to have wave properties as well as mass properties. This was especially noticeable for extremely light particles such as the electron. This idea eventually led to a model of the atom in which electrons exhibited these wave properties. In this model, the electrons are not at exact distances or fixed locations and do not rotate in orbits. This model of the atom is called the **Quantum Mechanical Model**.



WAVE AND PARTICLE NATURE OF THE ELECTRON AND PHOTON

All matter is now known to exhibit both particle and wave properties. In your independent research, you should discover a good explanation of the "wave-particle duality" of matter. You should also be able to identify and explain at least one experiment that shows that waves of light also have particle properties, and one experiment that shows that electrons also have wave properties.





Two simple electron orbitals, s and p.

IMAGE COURTESY OF DR. MICHAEL BLABER, DEPARTMENT OF BIOMEDICAL SCIENCES, FLORIDA STATE UNIVERSITY.

ELECTRON ORBITALS

Under the Quantum Mechanical Model, each electron in an atom creates an **electron cloud** that fits exactly at a specific distance and energy around the nucleus. This location is called an **orbital**. FIGURE 4 shows two simple atomic electron orbitals, which are called s and p to differentiate their shapes.

The number and shapes of electron orbitals in a specific atom determine the kind of chemical bonding that the atom can undergo. When the atom combines with other atoms to form a chemical compound, the structures and shapes of the molecules that are formed depend on the electron orbitals present. The limitations imposed by how the electron waves fit together underlie the fact that as protons and electrons are added to build larger atoms, certain patterns of relationships recur. When all known stable atoms are placed in a chart according to their electronic orbital structures, this is found to agree precisely with a pattern originally based on physical and chemical properties. This chart is known as the periodic table.

PERIODIC RELATIONSHIPS—DEFINITIONS, GENERAL TRENDS, AND CAUSES

It can be seen that variations in orbital structure explain the position of elements in the periodic table. As each element is added to the table, one proton is added, increasing the atomic number by one. Elements are added in a **period** (or row). When the type of orbital is repeated, the properties appear to repeat, so a new period is formed, and the next element is placed in the new period under the position furthest to the left. Filling the period then continues. In the first period, only one orbital exists, and it can hold two electrons. Thus, the second period begins at element three. For this period, two orbitals exist, which between them can hold eight electrons, and so the next period begins at element eleven. The trends in properties across the periods and down the columns of the periodic table are readily explained as a result of the changing trends in the electronic orbital structures of the atoms and the attractive forces between the nucleus and the electrons.

TRENDS ACROSS A PERIOD OF THE PERIODIC TABLE

Atomic radius (a measure of atomic size) decreases from left to right across a period in the periodic table. As one proton and one electron are added to the atom, there is one more proton to attract the electrons. If the added electrons are approximately the same distance from the nucleus, then the increased attraction will shrink the radius and size of the atom. For example, a carbon atom has six protons and six electrons. The next atom across the period, nitrogen, has seven protons and seven electrons. There is room for the seventh electron in nitrogen to be about the same distance from the nucleus as the sixth, so nitrogen atoms are smaller than carbon atoms due to the increased nuclear charge pulling them in closer.

Ionization energy increases from left to right across a period in the periodic table. The ionization energy is the energy required to entirely remove an electron from an atom. Thus, the ionization energy measures exactly the force of attraction holding the electron to the nucleus. As the number of protons increases, so does the average force pulling on the atoms' electrons. Thus, it will take more energy to remove an electron from a nitrogen atom than from a carbon atom. Nitrogen has a higher ionization energy than carbon.

The **electron affinity** of an atom is the amount of energy change when an electron is added to a neutral atom to form a negative ion. For example, $F + e^- \rightarrow F^-$. This property follows the same trend as those described above, generally increasing from left to right across a period of the periodic table as a result of the increased effective positive charge. So, fluorine will have a higher electron affinity than oxygen. Unfortunately, there are some exceptions (such as N) caused by details of the atoms' electronic orbital structure.

TRENDS DOWN A COLUMN OF THE PERIODIC TABLE

To consider trends in properties down a column of the periodic table starting at the top, a new concept must be introduced. Once a period across the table has been completed, there is no more room for electron orbitals at about the same distance. The next orbital allowed is at a significantly greater distance from the nucleus. This increase is often referred to as starting a new "shell." Once a shell begins to fill, it can continue until all the orbitals at that distance from the nucleus are filled. Once this happens, a new shell begins to fill a further distance from the nucleus, corresponding to the next period down the table. The elements in a vertical column of the periodic table were found to have such similar properties that they are referred to as a **group** or family.

Atomic radius increases down a column of the periodic table. This is a result of the increase in the distance of the outer electrons from the nucleus due to the added electron shell. Since the outer electrons are farther away, they are shielded from the attractive pull of the nucleus.

lonization energy decreases down a column of the periodic table. This is a result of the extra distance from

the electrons to the nucleus and thus a reduced attractive force.

Electron affinity decreases down a group in the periodic table. This is a result of the extra distance from the electrons to the nucleus and thus a reduced residual attractive force.



In your independent research, you should learn the definition of electronegativity and gain a general understanding of the concept of electronegativity. What are electronegativity's trends across periods and down columns of the periodic table? How is electronegativity useful for predicting the kind of chemical bonding that an atom will form?



Chemical Bonding and Intermolecular Forces

INTRODUCTION

Atoms of elements cluster together in arrangements that only break apart if energy is applied. Such a "cluster" is a chemical compound, and the specific arrangement of forces between two atoms is called a **chemical bond**. A group of atoms held together in a constant ratio by strong bonds is often called a **molecule**. For example, carbon dioxide, with the formula CO₂, is a molecule. Some strongly bonded compounds exist in the form of a network that extends over a very large number of atoms, for example silica (similar to sand), with the formula SiO₂, or table salt, with the formula NaCl. The **chemical formula** for compounds represents the ratio of atoms that exist in the large lattice, not the total number of atoms in the lattice.

There are forces of attraction between most molecules and neighboring molecules. These forces are called **intermolecular forces**. Intermolecular forces are responsible for the physical properties of substances. For instance, strong intermolecular forces between water molecules (H₂O) are the reason that water has a relatively high **boiling point** (100°C) whereas relatively weak intermolecular forces between methane (CH₄) molecules result in a very low boiling point for liquid methane (-162°C).

BINDING FORCES

lonic bonds, covalent bonds, metallic bonds. Van der Waals forces, and hydrogen bonds are five types of attractive forces that are recognized as being associated with atoms and molecules. All of these forces originate in just one type of fundamental force of attraction—that between positive

and negative charges, which is known as the electrostatic force.

Chemists further distinguish intramolecular forces as those that occur mostly between two atoms in such close association that a chemical bond is formed. These bonds are classified as ionic, covalent, or metallic bonds. Chemical bonds need substantial energy to break apart and generally are only broken in chemical reactions.

- Ionic bonds result when one or more electrons transfer from one atom to another, forming a positive and a negative ion. These ions are strongly attracted to one another and form an ionic bond. A common example is sodium chloride, which forms the ions Na⁺ and Cl⁻.
- Covalent bonds are formed when two atoms share a pair of electrons. The pair of electrons attracts the nuclei of both atoms, creating the force that holds the nuclei together and forming the covalent bond.
- Metallic bonds exist when atoms, generally of the same element, are bonded in such a way as to allow many electrons to move more freely among the atoms. This is called an electron "sea." These moving electrons make up electricity; hence, metals have good electrical conductivity. Examples of metals are lithium (Li) and sodium (Na). When more than one element is present, then the metal is called an alloy. Brass (made from copper and zinc) is a metal alloy and is a good conductor of electricity.

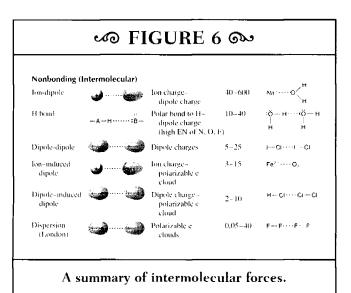
Intramolecular forces are summarized in FIGURE 5, which also shows typical energies and examples of substances containing each type of force.

Intermolecular forces are weaker forces of attraction than intramolecular forces. Intermolecular forces are responsible for causing a clustering affect in liquids and solids. Intermolecular forces are readily broken apart and/or formed when substances undergo physical changes, such

Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Bonding				
Ionic	#3	Cation-anion	400-4000	NaCl
Covalent	•	Nuclei-shared e pair	150-1100	НН
Metallic		Cations - delocalized electrons	75~1000	Fe

as when substances interchange from gases to liquids or to solids

- Van der Waals force is the name given to an intermolecular force first recognized by Johannes van der Waals (1837–1923). This force results from the partial charges existing either temporarily or permanently on molecules as a result of a lack of complete symmetry in the distribution of positive and negative charges. Molecules are said to be **polar** if a permanent distance can be measured between the center of positive charge and the center of negative charge. Such a molecule is also said to have a dipole moment, meaning that a rotational force will exist if the molecule is placed in an electric field. Molecules can develop induced dipoles as a result of the motion of the electron orbitals on atoms in a molecule. Even though these dipoles are temporary, they are still a source of attractive intermolecular forces. Such forces were delineated by Fritz London (1900–54) and are also known as **London** dispersion forces.
- Hydrogen bonds are formed when a strongly electronegative atom is bonded with hydrogen on another molecule. The most common example is water, H₂O. In liquid and solid water, the H atom on one water molecule forms a weak bond to the strongly electronegative oxygen atom on the next water molecule. This is an example of a strong dipole-dipole interaction. In FIGURE 6, partial charges are indicated by adding a δ sign (i.e., δ + or δ -). Because these hydrogen bonds must break in order for water molecules to move around, they are the reason that water has to be strongly heated to melt or boil. The rigidity conferred by these bonds is also responsible for the rather open structure found in solid water, commonly known as ice. This open structure is responsible for giving ice a lower density than liquid water (at a temperature just above its melting point), and so ice floats on water.



Intermolecular forces are summarized in Figure 6, which also shows typical energies and examples of substances containing each type of force.



THE IMPORTANCE OF HYDROGEN BONDING IN LIVING SYSTEMS

In your independent research, you should learn to identify and explain the key role that hydrogen bonding plays in determining the DNA structure of living cells.



THE EFFECT OF BOND TYPES ON THE STRUCTURES OF SOLID COMPOUNDS

Ionic Lattice

The electrostatic forces created by the positive and negative charges in an ionic bond extend in all directions, so multiple neighboring ions of opposite charge are attracted. Therefore, compounds with ionic bonds exist as lattice networks extending in three dimensions. The most common example is sodium chloride, which consists of a lattice of alternating Na⁺ and Cl ions repeated over and over again until the edge of a crystal of salt is reached.

Covalent Network or Molecular Solid

The force creating a covalent bond is focused on just the two atoms involved. However, depending on their electronic structure, many atoms can repeat the sharing process several times, bonding multiple atoms. For instance, while hydrogen only forms a single covalent bond to itself in the compound hydrogen (H₂), carbon can form up to four covalent bonds with other atoms, as in methane (CH₄). Covalently bonded molecules can exist as small discrete units, such as methane (CH₄) and water (H₂O), or as a **covalent network**, such as with a diamond, which consists of carbon atoms, each bonded covalently to four other carbon atoms in a lattice of such bonds—this forms one of the strongest known substances.

Metals with metallic bonds are highly ordered solids with atoms packed quite closely together. As a result, metals are strong, but can be bent, hammered into thin sheets, or pulled into wires, as layers of the atoms can slide over one another relatively easily.

HOW INTERMOLECULAR FORCES DETERMINE STATES OF MATTER (SOLID, LIQUID, GAS) AT ROOM TEMPERATURE

In a solid, molecules are not moving over long distances, a type of motion called **translational motion**, to distinguish it from **vibrational motion** and **rotational motion**,

which do occur in solids. In a liquid, limited translational motion occurs, but the molecules stay in close contact. In a gas, the molecules are free to move throughout the container in which the gas is placed. When substances change from solid to liquid or from liquid to gas, energy must be added to overcome any intermolecular forces that are preventing the molecules from moving around more freely.

The temperature of a substance is a direct measure of molecular motion. So, at any given temperature, the state of matter (solid, liquid, or gas) is determined by the strength of the intermolecular forces that can inhibit molecular motion. As a result, we can predict the state that a substance will have at room temperature if we know the types of intermolecular forces that it contains.

Solids are formed if strong forces exist in all three dimensions among the molecules and/or atoms. Thus, ionically bonded substances, such as sodium chloride (NaCl), and those forming covalent networks, such as quartz (SiO₂) or diamond (C), will be solids. Compounds with intermediate strength forces, such as hydrogen bonds, are liquids such as water (H₂O). Very few naturally occurring compounds are liquids at room temperature. Compounds such as methane (CH₄) and carbon dioxide (CO₂) have weak intermolecular forces that are called **dispersion forces** since there are no permanent dipoles in these molecules. Since these molecules are easy to separate from each other, they are gases at room temperature.

It is important to be clear about the difference between the types of bonds and the forces in them and intermolecular forces. Some compounds, such as liquid water, have both types of bonds. With liquid water, each individual molecule has two strong covalent H-O **bonds** and then up to two hydrogen bonds that "join" it to one or more adjacent water molecules.

POLARITY OF BONDS

The individual covalent bonds in molecules have an electron pair holding the two nuclei together. When the two atoms are identical, such as in H-H, or H₂, molecules, then the electrons are symmetrically distributed between the two nuclei. Because of the symmetry, there is no net charge separation, so the molecule is non-polar. However, in most molecules, the bond is not between two identical atoms. In this case, the electrons are not symmetrically distributed, but rather are more attracted toward the atom that has the greatest electronegativity value.

In your independent research on electronegativity, you should have found a table of electronegativity values. Since the electrons come from atoms that were originally neutral in charge, one atom is now getting "more than its fair share" of the electronic charge and so will be somewhat negatively charged, leaving the other atom with an equal but positive charge. Such a situation creates a polar bond. Molecules with such bonds will have permanent van der Waals intermolecular forces when they come close to other similar or identical molecules. An example of this is water.

∽ FIGURE 7 യം

Representative Lewis structures.

Molecular Models

INTRODUCTION

Chemists build real, mental, and mathematical models to help them understand the properties of matter and how atoms and compounds react to form new compounds. These models also help chemists to predict future reactions and to predict the properties of new materials. In the laboratory, chemists perform experiments to test their predictions. Chemists continually seek to develop new models or refine old ones to make predictions that better agree with experiments. However, even models that are known to incorrectly represent the real situation are kept if they assist in making useful predictions.

LEWIS STRUCTURES FOR ATOMS AND COVALENT AND IONIC COMPOUNDS

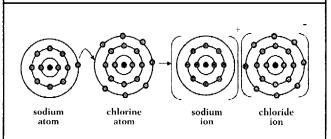
G. N. Lewis (1875–1946) at the University of California developed an idea that helps to account for where electrons are in atoms and compounds by using a dot to represent each electron. While the modern view of electrons in atoms does not visualize them as dots, Lewis's model helps account for all the electrons and the common bonding observed for each atom. The electrons in the outermost shell of an atom are known as **valence electrons** and are the ones used in forming bonds.

Lewis structures can be drawn for individual atoms, ions, and atoms in compounds. Complete Lewis structures sometimes show all of the electrons, though usually only valence electrons are shown. Lewis also decided that covalent bonds should each have a pair of electrons. The structures shown in Figure 7 are representative Lewis structures.

Notice that the structures for F_2 , H_2O , and NH_3 have electron pairs in bonds, but also electron pairs that are not involved in bonding. These are known as non-bonding electron pairs, or simply "lone pairs." The total electrons in each neutral structure must add up to the total number of protons in the nuclei of the atoms. In ions, this number can differ by the amount of the value of the ion's charge.

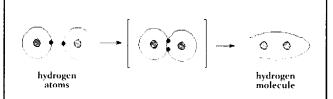
FIGURE 8 shows how a complete Lewis structure can be used to show the change as a sodium atom gives up its one outer valence electron to the nearly full valence shell of the chlorine atom to form ions of opposite charge that then attract one another in an ionic bond.

∽ FIGURE 8 യെ



Lewis structure depicting an ionic bond between sodium and chlorine.

∽ FIGURE 9 യെ



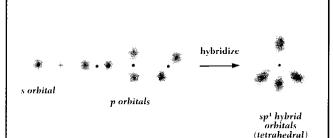
A depiction of a single bond.

IMAGE COURTESY OF DR. GARY TRAMMELL, UNIVERSITY OF ILLINOIS AT SPRINGFIELD.

ൃ FIGURE 10 യ

A depiction of double and triple bonds.

ൃ FIGURE 11 യ



Depiction of the hybridization of atomic orbitals.

VALENCE BOND CONCEPT

As chemists developed the idea that electrons in atoms were in orbitals, they needed to fit the successful Lewis electron pair bond that explained compound formation and formulas into this picture. This was done by supposing that an electron in an orbital on atom A interacted with an electron in an orbital on atom B with which it forms the bond.

FIGURE 9 illustrates the idea that the electron orbital shapes, which represent the map of electron density, overlap between the atoms to provide the attractive force in the bond. Such a bond is called a "single" bond.

In some cases, atoms can overlap more than one orbital at the same time to form "double" or even "triple" bonds as is shown in Figure 10. Carbon and oxygen share two pairs of electrons in forming two double bonds in carbon dioxide, while nitrogen atoms share three pairs of electrons in forming a triple bond in nitrogen gas.

HYBRIDIZATION OF ATOMIC ORBITALS AS AN EXPLANATION FOR MOLECULAR SHAPES

Sometimes the interaction of the two atomic orbitals doesn't seem to predict the correct shapes of the molecules. In response to this concern, chemists developed the idea that two (or more) electron orbitals in an atom could combine to form new orbitals that had a different shape. This process is known as **hybridization** and can be used to rationalize the symmetric shapes of many molecules. Figure 11 shows how 1 s and 3 p type orbitals can hybridize to form 4 sp³ orbitals.

THE FORMATION OF MOLECULAR ORBITALS

In your independent research, you should learn the answers to the following questions:

- 1. What are electron molecular orbitals?
- 2. What is the difference between sigma and pi molecular orbitals?
- 3. How are these orbitals thought to be arranged in N₂, O₂, F₂, and CO₂?



VSEPR CONCEPT

Lewis structures only tell part of the story about chemical molecules because they are two dimensional, whereas all real molecules are three dimensional. Additional models are needed to understand and predict the geometric shapes of molecules. One of the most successful such models is the **Valence Shell Electron Pair Repulsion Model**, abbreviated VSEPR. This model takes the electron pair bond one

step further by proposing that electrons form pairs, but then the pairs, being all negatively charged, will mutually repel one another in three dimensions.



THE RESONANCE CONCEPT MODEL

In your independent research, you should discover what the resonance concept model is when applied to molecule structures and should be able to explain how resonance is used to rationalize the equal nature of bonds in the compounds O_3 and SO_3 .



OXIDATION STATES

Oxidation states are numbers that can be assigned an atoms that represent if the atom has lost or gained electrons to form an ion or bonds to other atoms. Oxidation states are computed by starting with the actual number of valence electrons the atom had before bonding or ionizing, and then determining whether the atom has lost or gained one or more electrons as it bonded or ionized.

For example, in Li⁺, the lithium atom had three electrons and has lost one, so its oxidation state is ± 1 . In H_2O , the hydrogen atom has effectively given up its electron to form the bond, so it has an oxidation state of ± 1 . The oxygen has gained two electrons (one from each hydrogen atom) in forming the bonds and so has an oxidation state of ± 2 . Note that the total oxidation state for all atoms in a neutral molecule must sum up to zero. For a charged species such as NH_4^+ , the total sum of the **oxidation numbers** will be equal to the charge, namely ± 1 . So, in NH_4^+ , if each H is ± 1 , N must be ± 3 .

GEOMETRY OF MOLECULES AND IONS

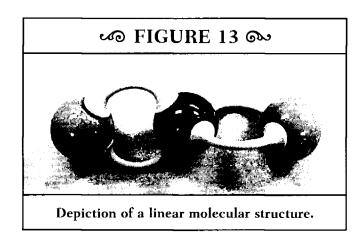
VSEPR makes predictions about the geometry of simple molecules by focusing on the central atom; for example, if there were four electron pairs forming bonds around a central atom, these would point toward the corners of a tetrahedron, as is shown Figure 12, which shows a model for a methane molecule.

DIPOLE MOMENTS OF MOLECULES

Earlier, we described dipoles and polar bonds. One of the clues to the geometry of a molecule that has polar bonds is whether the entire molecule has a dipole moment. If it does not, then symmetry must have allowed for the sum of dipoles pointing one way to cancel the sum of dipoles pointing the other way. (If you have learned about vectors in math, this means that the addition of vectors equals zero.) If the molecule does have a significant dipole moment, then

FIGURE 12 ®>

A VSEPR model for a methane molecule.



there must be a geometry in which the bond dipoles do not cancel.

For example, in CO_2 , carbon dioxide, the C=O bonds are polar (i.e., they have dipoles, or charge separation), but the molecule has zero dipole moment. So, we conclude that one C=O bond must exactly oppose the other. Thus, the structure would be linear as is shown in the model pictured in Figure 13.

THE RELATIONSHIP BETWEEN MOLECULAR PROPERTIES AND MOLECULAR STRUCTURE

The polarity and shapes of molecules affect the way in which intermolecular forces are distributed. Highly **polar molecules** will exhibit strong van der Waals forces, making the molecules more difficult to separate. Such molecules will therefore have higher melting points and boiling points than molecules with lower polarity. Molecules that are non-polar will have the lowest melting and boiling points.

Anchear Chemistry

NUCLEAR EQUATIONS, HALF-LIVES, AND RADIOACTIVITY

Nomenclature for Isotopes (Nuclides)

Recall from our discussion earlier that the nuclear structure of an atom is defined by the symbol A_ZX , where X is the chemical symbol, Z is the atomic number or number of protons, and A is the mass number, which is equal to the number of neutrons (N) plus the number of protons. So, A = Z + N.

Many stable nuclei exist, but varying the ratio of neutrons to protons (N/Z) can lead to unstable nuclei, making the atom radioactive. **Radioactive atoms** can regain stability via a number of pathways. **Alpha decay** involves the loss of a helium nucleus. With **beta decay**, a neutron transforms into a proton with the emission of a beta minus particle (electron), or a proton transforms into a neutron in beta plus or **positron decay**.

These pathways are shown in the nuclear decay equations in Figure 14.



DECAY EQUATIONS AND PREDICTING PRODUCTS OF DECAYS

Radioactive decay occurs by well known and predictable pathways. Your independent research should enable you to write typical decay equations for alpha, beta minus, and positron decay of an unstable atom. It should also enable you to predict the products of these decays if you are given the identity of the parent nuclide and the decay type.



ശ FIGURE 14 യ

- Alpha decay
- $^{238}U \rightarrow ^{234}Th + ^{4}He (^{234}Th + 2e^{-} + ^{4}He^{2+})$
- Beta decay
- ${}^{14}C \rightarrow {}^{14}N^+ + \beta^- + \overline{\nu}$
- 22 Na $\rightarrow ^{22}$ Ne + e⁻ + β ⁺ + ν + Q 1022 keV

 β = **beta particle** ν = neutrino

Types of decay.



ALPHA BOMBARDMENT REACTIONS

Nuclear transformations can also be initiated by bombarding one type of atom with another atom or with a fundamental particle, such as an alpha particle or a neutron. The first nuclear transformation performed was carried out using alpha particle bombardment. Your research should enable you to write a balanced equation for an alpha bombardment reaction.





FISSION AND FUSION REACTIONS

Uranium fission and the fusion reaction of deuterium (2H) with tritium (3H) are both possible sources of energy for human use. Your independent research should aim to answer the following questions:

- 1. What happens in fission of the uranium isotope 235U?
- 2. What happens when fusion of deuterium with tritium occurs?

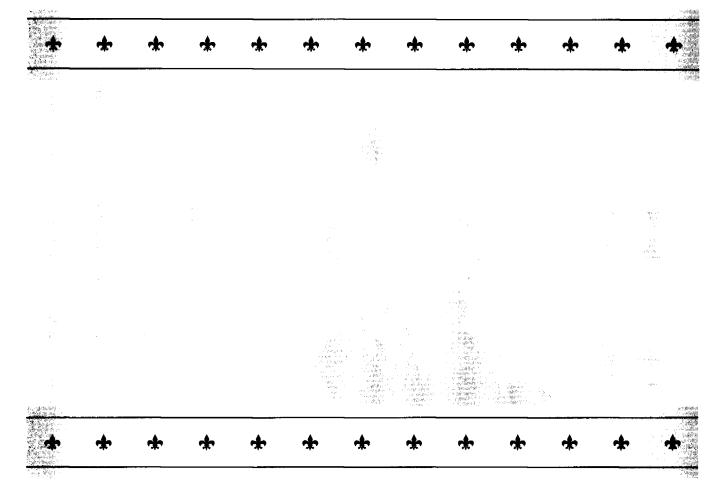


Section 11 Summary

- Matter consists of atoms, neutral particles made of protons (+1 charge), neutrons (0 charge), and electrons (-1 charge).
- An element has one type of atom distinguished by its number of protons. Isotopes of an element can have a different number of neutrons.
- Mass is a characteristic that helps to define an atom and may be used for its identification. Masses are compared to a carbon isotope defined as having a mass of exactly twelve atomic mass units.
- A definitive symbol system has been agreed upon for atoms. This system shows the number of protons, the chemical symbol, and the mass number.

- Atomic masses of elements are weighted averages of their isotopic masses.
- "cloud" wave shape called an orbital. The early Bohr model of the atom had the electrons in circular orbits around the atomic nucleus.
- Elements placed in order of atomic number show periodically varying properties and can be arranged in a periodic table.
- Properties of elements placed correctly in the periodic table exhibit trends across and down the table that can be explained by the electronic orbital structure of the atoms of the elements.
- These properties include atomic radius, ionization energy, electron affinity, and electronegativity.
- Various forces of attraction have been identified that hold atoms together in clusters called molecules or that attract one molecule to another.
- Intramolecular forces, which hold atoms tightly together in molecules or networks, are those found in chemical bonds. Ionic, covalent, or metallic bonds are the three types of bonds that explain all molecular and network structures.
- Intermolecular forces are weaker than intramolecular forces. Intermolecular forces account for the differ-

- ences in physical properties of substances, including whether a substance is a solid, a liquid, or a gas at room temperature.
- Lewis dot structures show electron distributions in atoms and molecules. Lewis structures are useful for illustrating the formation of covalent or ionic bonds according to Lewis's idea that two electrons are used in covalent bond formation.
- The Valence Shell Electron Pair Repulsion Model (VSEPR) is used to explain the three-dimensional shapes of simple molecules. When multiple covalent bonds are formed by one atom, hybrid atomic orbitals are sometimes needed in order to explain the shape that is found by experiment.
- Knowing the shapes of molecules enables a chemist to predict polarity. In turn, polarity is useful in predicting physical properties. For example, polar molecules have higher boiling points than non-polar molecules.
- Changes in the nucleus of an atom follow patterns similar to electronic changes. Radioactivity involves changes in nuclear structure to gain stability.
- Balanced equations can be written for alpha and beta decay pathways for all radioactive atoms and for nuclear fusion and fission transformations.





Section III:

States of Matter



Introduction

Tumans live their lives amidst a sea of chemical substances that are in different physical states, Deither gas, liquid, or solid. Most are not single substances, but mixtures of substances. If a mixture looks as if it is a single substance, it is called a solution. There are several different kinds of solutions, although most people think of solutions as liquids dissolved in liquids, or solids dissolved in liquids. Examples of these types of solutions include lemon juice and sugar in water, which make a solution called lemonade. Chemists also recognize gas in gas solutions and solid in solid solutions. For example, you breathe the gases that are found in the gas solution called air; you drink gas, liquid, or solid in liquid water solutions when you enjoy beverages such as lemonade, soda-pop, tea, or coffee; and you may ride a bicycle made of high-strength solid solutions called steels. Only gas in gas solutions and solid in liquid solutions will be reviewed in this section of the resource guide.

This section of the resource guide will explore the properties of each state of matter and how they vary with composition, temperature, and pressure. We will begin our discussion with pure gases and then proceed to pure liquids. You will learn that chemists discovered that all substances (except at **absolute zero temperature**) can be considered as an ensemble of rapidly moving particles. The dynamic motion and energy of these particles form the basis of the **Kinetic Molecular Theory** of matter. This model can be used with appropriate mathematics to explain the laws that govern the behavior of gases, which were discovered much earlier by experiment.

You will also learn about the major differences between gases, liquids, and solids and about **phase diagrams**, which show how all three states for a pure substance can exist under different conditions, and how they can interchange in a process of molecular dynamic equilibrium. Then we will examine the properties of solutions in water and in **organic** solvents. You will learn that these properties can be predicted if the molecular properties and the relative concentrations of the components of the solutions are known.

Gases

LAWS OF IDEAL GASES

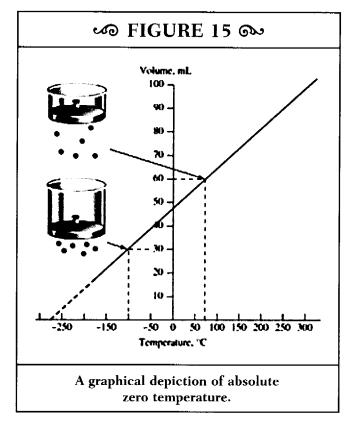
At the start of the Scientific Revolution, scientists were fascinated by the behavior of gases, which they had learned to isolate from one another or from air. These scientists observed that gases exhibited a high degree of regularity of behavior.

If a certain mass of gas was trapped in a flexible container, such as a balloon, then if extra pressure (P) was applied, the **volume** (V) of the balloon always decreased in a direct proportion to the applied pressure. This behavior pattern became known as **Boyle's Law**, which can be stated in a simple algebraic equation: $P \times V = a$ constant, or PV = C. Pressure is usually measured in atmosphere units, and volume in liters, so C is measured in units of liter-atmospheres (l-atm).

Similarly, **Charles' Law** states that if you heat the trapped gas, the volume increases in direct proportion to the temperature change, so that $V/\Gamma = a$ constant, or $V/\Gamma = D$. With volume in liters and temperature in degree units, the constant will be measured in units of liters per degree.

In experimenting with gases, scientists made an interesting discovery. They found that when a gas was trapped in a cylinder with a movable piston and the weight of the piston applied a constant pressure to the gas, as the temperature was lowered from about 70°C to 100°C, the volume of the trapped gas decreased from 60 mL to 30 mL. Other measurements would show that a straight line could be drawn through all the observed volumes. However, this straight line extrapolates to zero volume at a very low temperature! Scientists were surprised to find that this plot worked for every gas they tried, and that the "zero volume" temperature was always the same, namely –273°C. See Figure 15 for a graphical depiction of this phenomenon.

Since it seemed that no temperature can ever go lower than ~273°C (and this has been found to be the rule), this temperature was called "absolute zero." Since that time, it has been agreed that this temperature is a fundamental starting point for temperature measurements, so a scale called the **Kelvin** scale (unit K) is used that starts at ~273°C, which is defined as 0 K. To convert from the more familiar °C (Celsius), you simply add 273. So, on this scale, 70°C is 70 + 273 = 343 K. Notice that since the Kelvin unit



is defined as "degrees Kelvin" no degree (°) sign is placed in front of the K. Since the Kelvin temperature scale is linear and starts from 0, it is this unit that must be used in all equations relating to Charles' Law.

Combining Boyle's Law and Charles' Law algebraically is quite simple. If $P \times V = C$, and $V/\Gamma = D$, then $PV/\Gamma = CD$, a new constant. In using this relationship, we usually consider a set of pressure, volume, and temperature conditions (1) changing to a new set of conditions (2), so we can then write:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}.$$

This relationship can be applied both qualitatively and quantitatively to predict new conditions.

For example, consider the following problems:

1. A sample of gas is in a cylinder with a movable piston (See Figure 15). If the pressure is increased while the temperature is unchanged, will the volume increase, decrease, or remain the same?

ANSWER: The volume will decrease to balance the pressure increase.

2. A sample of gas in a rigid container of volume 1.0 liter has a pressure of 2.0 atmospheres at a temperature of 200 K. What will be the pressure at 400 K?

ANSWER: The volume and amount of gas is constant, so P/T must stay constant. And, if T doubles from T_1 to T_2 , then P must also double, so the new pressure, P_2 , is 4.0 atmospheres.



PARTIAL PRESSURES AND CORRECTION OF GAS VOLUMES COLLECTED OVER WATER

The effect of changing pressure, volume, and temperature for mixtures of ideally behaving gases is more difficult than for making such changes to a single gas. However, Dalton's Law of partial pressures enables us to consider each gas separately. In your independent research, you should find how Dalton's Law is used to explain how mixtures of gases follow the gas laws that we have examined. To isolate a gas, scientists often must use water to trap the gas in a container. Your research should also answer the question of how the true pressure of the gas is determined if water is present.



KINETIC MOLECULAR THEORY (KMT)

In our previous discussion of the ideal gas laws, it was assumed that gas molecules would behave in a very ideal way. A scientific model called the Kinetic Molecular Theory (KMT) of gases also uses this approach. Specifically, KMT makes four assumptions about ideal gases and molecules:

- A pure gas consists of identical molecules that are very small, so the actual volume of the individual molecules is negligible.
- 2. The molecules are moving very rapidly in all directions, but not all at the same speed.
- 3. No forces of attraction or repulsion exist between one molecule or another.
- 4. Gas pressure is a result of collisions of the molecules with the walls of the container. While the molecules change direction as they bounce off the walls, they lose no energy during these collisions.

RELATION OF PARTICLE SPEED TO MASS, COLLISION FREQUENCY, VOLUME, AND PRESSURE

Using these assumptions, it is possible to derive mathematical expressions for the pressure of a gas. We start by defining an average speed of a molecule, u. This can be used to determine the frequency of collisions with the walls, given the size of a container—in this example, a cubic box with sides of length ℓ . Next, we consider the momentum transferred to the walls of the container as a result of the gas molecule collisions.

The **force** on a container wall is a direct result of the momentum change of a molecule as it bounces off the wall.

For a molecule that travels a distance ℓ between collisions, the physics of mechanics tells us this is

$$Force = \frac{mu^2}{\ell} \text{ per second,}$$

where m is the mass of the molecule, and u is its average speed in meters/second directed toward the wall.

In a three-dimensional box, if the area of one side is A, then the volume is $A\ell$. If N molecules are in the box, we can assume 1/3 of them will be traveling in each of the three dimensions. So, the pressure, P, or force/unit area (P = Force/A) on the walls of the container as a result of gas collisions totals:

$$P = \frac{1}{3} \frac{Nmu^2}{V}$$
 or $PV = \frac{1}{3} Nmu^2$.

In qualitative terms, this is the relationship that explains Boyle's Law. It also states that the product of $P \times V$ will depend directly on the mass of the gas molecules and the square of their average speed. We will see later how Charles' Law and the temperature of a gas can also be rationalized using this model.



Amedeo Avogadro, the Italian scientist who developed the hypothesis that the relationship between the masses of the same volume of different gases (at the same temperature and pressure) corresponds to the relationship between their respective molecular weights.



AVOGADRO'S LAW

It is important to note that in our discussion of gases thus far, we have been referring to a fixed amount of trapped gas. What happens if the amount of gas changes? In studying gases, scientists found that the pressure and volume of a gas did change in a predictable way as the amount of the gas was varied. However, mass wasn't the only factor; what mattered was the ratio of the total mass of gas to the mass of the individual molecules of the gas.

For example, four grams of helium gas have the same volume as two grams of hydrogen gas at the same temperature and pressure. Sixteen grams of methane (chemical formula $\mathrm{CH_4}$) are needed to reach the same volume, temperature, and pressure conditions.

This behavior provides the basis for establishing the molecular theory of matter. The molecular theory of matter states that many pure substances, especially gases, exist as small discrete units called molecules, and it is the *number* of these molecules that are present that determines the properties, not the mass. An extension of this concept was made by Amedeo Avogadro, an Italian scientist who suggested that no matter what the nature of the gas, at a given temperature, pressure, and volume, there would be the same number of molecules. This concept is known as Avogadro's hypothesis. Since it turned out to be correct (except for minor perturbations under unusual conditions), it is also know as **Avogadro's Law**.

VOLUME AND MASS OF ONE MOLE OF IDEAL GAS AT STANDARD TEMPERATURE AND PRESSURE (STP)

Quantitatively, chemists define a specific number of molecules (or atoms, particles, or ions) as a unit called the **mole**. For historical reasons, the mole was defined as the number of molecules in a volume of 22.4 liters at 1 atmosphere pressure, and a temperature of 273 K. A more recent agreed-upon definition of one mole is the number of atoms in exactly 12 grams of the carbon isotope ¹²C. Considerable effort has gone into measuring this number with greater and greater precision, but it may never be known exactly. Some scientists have suggested that it should be redefined as an exact number, but history and generations of college chemistry majors who have learned the number are probably against such redefinition.

The number was found to be 6.022×10^{23} molecules. This number is known as **Avogadro's number** in honor of Avogadro, though he did not actually determine it. So, there will be 6.022×10^{23} atoms of helium in a 22.4 liter container of helium gas if the pressure is 1 atmosphere and the temperature 273 K. This helium will have a mass of 4 grams. This number is known as the **molar mass** of helium, which is 4 g/mol. Similarly, there will be 6.022×10^{23} molecules of methane (CH₄) in a 22.4 liter container at 1 atmosphere pressure at 273 K. The methane will have a mass of 16 grams. The molar mass of methane is 16 g/mol.

You can calculate the molar mass of any substance by multiplying the atomic masses of the atoms by the number of atoms of that element in the molecule and then adding them up. So, for example, to determine the molar mass of methane, we have 1 C and 4 H's = 12 (atomic mass C) + 4 \times 1 (atomic mass H) = 16.

The concept of moles as the unit of substance has been widely adopted in chemistry for all substances, not just for gases. So, for example, we can have 1 mole of salt (sodium chloride, NaCl) which will have a mass of $(1 \times 23.0) + (1 \times 35.5)$ or 58.5 grams. The molar mass of NaCl is 58.5 g/mol.

DEPENDENCE OF MOLECULAR KINETIC ENERGY ON TEMPERATURE

The speed of a single, isolated, moving molecule determines its energy. Physics provides us with the equation: energy = $\frac{1}{2}$ mv^2 for a moving particle of mass, m, and velocity, v. Physicists also determined that energy of motion is directly related to temperature, if temperature is measured on the Kelvin scale. So, the energy of a molecule = 3/2 kT, where T is the temperature of the gas (in K), and k is a universal constant known as the **Boltzman constant**.

Since both expressions represent the same molecule's energy, we can put them together and do some algebra to give us an equation for an average molecule's speed, *u*:

$$u = \sqrt{\frac{3kT}{m}}.$$

Since this is derived by taking a square root, it is known as the "root mean square" speed.

If we now consider N molecules, instead of just one, where N is Avogadro's Number, we have 1 mole of gas, and the mass is the molar mass, M. The Boltzman constant can be recomputed for 1 mole of gas and given the symbol R, which is known as the "universal molar gas constant."

The equation for root mean square speed of molecules in a mole of gas is

$$u = \sqrt{\frac{3RT}{M}}$$
.

It is interesting to question how fast gas molecules are moving. If we apply this equation to helium atoms (such as might be in your birthday helium balloon on a warm day), put in a molar mass, M, of 4, and a temperature, T, of 298 K (25°C or 77°F), and get all the units correct, we find a root mean square speed of 1360 meters per second. That's almost a mile per second!

The calculation looks like this:

$$\sqrt{\frac{3 \times 8.31 \text{Kgm}^2 \text{s}^{-2} \text{mol}^{-1} \text{K}^{-1} \times 298 \text{K}}{4 \times 10^{-3} \text{Kgmol}^{-1}}}.$$

The helium atoms inside the balloon are zipping back and forth many thousands of times per second, creating enough pressure to keep the balloon inflated.

THE IDEAL GAS EQUATION

If we now substitute the expression for u into the equation we worked out previously, namely $PV = 1/3 \ Mu^2$ (since Nm = M), we find:

$$PV = 1/3 M \times 3RT/M = RT.$$

This equation is known as the universal gas equation, or the **ideal gas equation**. We should keep in mind that here we have assumed 1 mole of gas. However, we can modify the equation easily for any number of moles of gas by including n moles, so PV = nRT.

Note that this has exactly the same form as the combined Boyle's and Charles' Law equations for a fixed amount of gas, since *R* and *n* will be constant:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}.$$

It was very satisfying to early physicists and chemists that they could model the behavior of gases by making the assumptions of Kinetic Molecular Theory and using equations for the motion of a particle.



THE BEHAVIOR OF GASES UNDER EXTREME CONDITIONS

Under extreme conditions of temperature and pressure, most gases exhibit behavior that doesn't agree with the Ideal Gas Equation. In your independent research, you should explore what happens if gas molecules do not have characteristics exactly similar to the assumptions made in the Kinetic Molecular Theory (KMT). How do some common gases such as He, O₂, CH₄, and CO₂ behave under conditions of extreme high and low temperatures and pressures? What is the explanation for the differences in the way they behave? Why is this important, for example, in our choice of substances to use as refrigerants?



DEPENDENCE OF RELATIVE RATES OF DIFFUSION AND EFFUSION ON MOLECULAR MASS AND TEMPERATURE

One obvious consequence of accepting that a gas consists of large numbers of small particles moving at high speed is that we can predict that gases will rapidly travel great distances if released from a container. The atoms in the helium example discussed previously were moving at a mile a second. We become very familiar with this idea as we experience, almost instantaneously, the smell from a perfume bottle opened across the room or the odor from a skunk killed on the road.

Two concepts are at play in these examples. The one that is demonstrated by the perfume molecules is called **diffusion**. The other, which involves how fast a gas will leave a container through a small leak such as a pinhole, is a process called **effusion**.

The simplest calculation used for the helium example assumes that molecules can travel uninterrupted in a straight line. This works in a vacuum, but doesn't work in practice because outside of a vacuum a molecule makes many collisions with other molecules that divert it from a straight-line path. So, we can easily calculate a *relative* diffusion or effusion rate, but probably would need an elaborate computer model to compute the absolute rate at which the average molecule in a gas will diffuse across a room or effuse from a pin-hole leak.

Applying algebra to the root mean square speed equation, we can see that for two gases, (1) and (2), at the same temperature:

$$u_1 = \sqrt{\frac{3RT}{M_1}}$$
 and $u_2 = \sqrt{\frac{3RT}{M_2}}$.

If we divide the equation on the left by the one on the right to form a ratio, we can write:

$$\frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}}$$

Stated in words, the relative speeds of gas molecules (u) are inversely proportional to the square root of the mass of the gas molecules (M).

As an example, consider this problem: If helium atoms (M = 4) are traveling at 1 mile per second, how fast are H_2 molecules (M = 2) traveling at the same temperature?

ANSWER: We have
$$u_{H_2} = \sqrt{\frac{1}{2}} \times u_{He} = \sqrt{2} \times 1$$
 mile/sec = 1.4 miles per second.

Qualitatively, we can easily see that the heavier the gas molecules are, the slower they will move (at a given temperature), and so we can predict that diffusion and effusion will be slower. Thus, in the above problem we could immediately state that H_2 molecules (M=2) will travel faster than H_2 atoms (M=4) if they are at the same temperature.

For another example, consider the following problem: A scientist at a lab bench opens a bottle of ammonia (NH₃) and a bottle of acetone (C_3H_6O) at the same time. Both have a distinctive smell. Which substance will someone on the other side of the room detect first by its smell?

ANSWER: The molar mass of ammonia = 14 + 3 = 17, while the molar mass of acetone = 36 + 6 + 16 = 58. So, clearly the ammonia molecules have the smaller mass and therefore will be sensed first. Since they are in the same room, the temperature will be the same for both substances.

Liquids

In the previous section focusing on gases, we have considered ideal gases, where the molecules (or atoms in the case of noble gases) of gas are independent and have no intermolecular forces causing them to stick together. In a solid substance, the individual particles can be molecules or atoms or ions. These are held firmly by attractive forces in a network or lattice structure that allows almost no movement from one place to another (translational motion). This is not to say that the particles are still. In fact, there

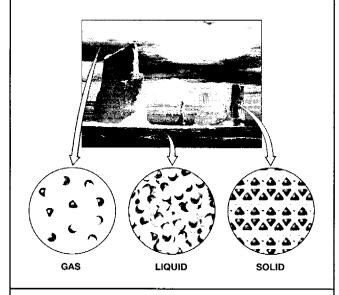
are rapid vibrations and rotations occurring in every solid at all temperatures above absolute zero. However, these vibrations and rotations generally do not change the average position of a particle in relation to its nearest neighbors. The structure of a solid remains much the same, with only slight expansion due to increased vibrations as the temperature is raised, until it changes phase, either melting to a liquid or subliming to a gas. One way to describe a gas in comparison to a solid is that a solid has a high degree of order at short distances (called short-range order) whereas a gas has no ordered structure at all.

A liquid is somewhere in between a gas and a solid. Individual particles or "clusters" of particles weakly "stuck" together by intermolecular forces may move readily (and may vibrate and rotate) throughout the liquid. Experiments show that over a long distance, a liquid has a form of ordering, but over short distances the molecules appear randomly distributed. A liquid may thus be said to exhibit long-range order. The degree of ordering depends very much on the nature of the liquid and the temperature. A liquid such as water, which has strong intermolecular forces due to the hydrogen bonds between the water molecules, has a much higher degree of ordering than does a liquid such as **octane** (a component of gasoline), which has very weak forces between the octane molecules.

The differences between gases, liquids, and solids are illustrated in Figure 16.

In this photo we see both the liquid and solid states of water. We cannot see water vapor. What we see when we look at steam or clouds is tiny droplets of liquid water dispersed in the atmosphere. The molecular views show that

ൃ FIGURE 16 യ



States of matter.

IMAGE FROM CHEMISTRY: THE CENTRAL SCIENCE, 11TH EDITION, BY THEODORE E. BROWN, ET AL., PUBLISHED BY PEARSON PRENTICE HALL, ©2009. the molecules in the gas are much further apart than those in the liquid or solid. The molecules in the liquid do not have the orderly arrangement seen in the solid.

Notice, also, that a usual distinction is that the density (mass per volume) of a solid is the greatest, and a gas the least, with the liquid state being in between.

Solids

Solids have a distinctive structure in the sense that they are highly ordered. While vibrations do occur, we can take an instantaneous snapshot of a solid showing its components (atoms, molecules, or ions) in specific, essentially fixed, positions relative to one another. The component particles are held firmly by the ionic, covalent, or van der Waal's forces described in Section II of this Resource Guide. In

FIGURE 17 600

- a) Diamond
- b) Graphite
- c) Lonsdaleite
- d) C60 (Buckminsterfullerene)
- e) C₅₄₀
- f) C₇₀
- g) Amorphous carbon
- h) Single-walled carbon nanotube

Allotropes of carbon.

IMAGES CREATED BY MICHAEL STRÖCK.

Section II, we also identified four types of solids: **ionic lattice** solids, covalent network solids, **molecular solids**, and metallic solids.

Some chemical compounds and elements can exist in several different solid forms. When an element shows this property, the different forms are called **allotropes** of the element. One excellent example is the element carbon, which has been identified as existing in several different forms, each with its own distinct properties. FIGURE 17 shows eight forms (allotropes) in which pure carbon has been found.

Properties can differ a lot among these allotropes. For example, graphite conducts electricity, but diamond is an **insulator**. Some of these forms (for example, nanotubes) have only recently been identified and may have incredible applications, such as in forming parts of circuits that can control nano-robots.

Carbon is found in so many forms because of the many different bonding arrangements that carbon atoms can have that form stable structures. Each structure does differ slightly in its energy content. The most stable form of carbon, defined in terms of energy at room temperature, is the graphite structure. However, the other forms of carbon do not convert at any appreciable rate into graphite when at room temperature. Natural diamond (that found on Earth) is formed from graphite at high temperatures and pressures underground. Scientists have now learned how to simulate such conditions and thus can manufacture artificial diamonds from graphite.

In addition to bonding ability, another factor that affects how solids are arranged is the size of the particles in the solid. This is especially true for ionic lattice solids and for metallic solids. The forces that hold particles together, and hence the energy, can be maximized when the particles are packed as tightly as possible.

Among the most closely packed structures are pure metals, which have atoms of one single element present. Metals have a regular **crystalline** structure. You can see crystals of zinc on the surface of a galvanized object, such as a bucket or garden watering can. The atoms of a metal are organized in successive layers and can be organized in three simple forms. These forms are determined by how many "nearest neighbors" (i.e., similar particles at the nearest identical distance) an atom has.

In a pure metal, the nearest neighbors are always an identical atom. The three simplest structures seen are **body-centered cubic**, which has eight nearest neighbors (iron is an example): **cubic closest packed**, which has twelve nearest neighbors (gold is an example); and **hexagonal closest packing**, which also has twelve nearest neighbors (zinc is an example). The very tight structure of metals causes them to have the following properties:

- Metals have a **lustrous** appearance (shiny) because of the reflection of light from the regular arrangement of atoms:
- They are good conductors of heat because of their structure. Sound waves are able to travel through the

lattice, and so metals are **sonorous** (they ring when struck):

- They are good conductors of electricity because the closeness of the structure helps atoms give up some of their outer electrons to move freely in the solid as a whole:
- They are very **malleable**—they can be beaten into different shapes; and they are **ductile**—they can be pulled into a wire. The evenly spaced layers of atoms can slip over each other without losing their attractive forces.

Generally, the closer the packing, the "softer" the metal will be, as its layers can more easily slide over one another. You can see in Figure 18 how easily layers can slide over one another in a packed stack of identical spheres. Figure 18 also shows the alternative way that layers can be stacked. The preferred structure of the metal is determined by the size of the atoms.

The reason that impurities are added to metals—or that two or more metals are mixed to form an alloy—is because this disrupts the smooth layers, decreasing the sliding of layers, thereby increasing the strength of the metal for structural purposes. Steel is much stronger than pure iron for this reason.

Thase (Diagrams of One-Component Systems

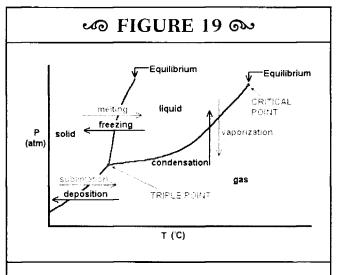
A pure substance can exist as a gas, a liquid, or a solid, depending on the temperature. The relationships between the three **phases** and the transitions between them can be summarized in a phase diagram. It is important to understand how a phase diagram is constructed in order to be able to understand what happens as one phase changes into another. There are several important concepts involved:

 Most phase diagrams are constructed as if there is a single substance present in a sealed container. Not even air is present.

FIGURE 18 6 A Close-packed layer of spheres (a) (b) (c)

Depiction of various metal structures.

IMAGE FROM CHEMISTRY: THE CENTRAL SCIENCE, 11TH EDITION, BY THEODORE E. BROWN, ET AL., PUBLISHED BY PEARSON PRENTICE HALL, ©2009.



A typical phase diagram.

IMAGE COURTESY OF THINKQUEST, ORACLE EDUCATION FOUNDATION, http://library.thinkquest.org.

- 2. Molecules (or particles) are constantly moving from one phase to another in what is known as dynamic transfer.
- Equilibrium conditions can be measured such that the rate at which particles leave a phase is the same as the rate at which they enter the phase.
- 4. At any temperature, any phase of a substance will have a small amount of its gas present. When the substance is not wholly a gas, the gas present is known as a vapor. In a sealed container, this gas will have a gas pressure associated with it.

A typical phase diagram for almost any substance is shown in Figure 19.

With regard to the phase diagram shown in Figure 19:

- The pressure in the container is plotted on the vertical axis, and the temperature of the container of the substance is plotted on the horizontal axis.
- Each phase is represented by a region of pressure-temperature values.
- The curved lines dividing the phase regions represent the values of pressure and temperature when equilibrium exists between *two* phases.
- A single point, called the **triple point**, represents the only value of pressure and temperature at which solid, liquid, and gas can exist *in equilibrium*.
- Horizontal lines represent familiar transitions from one phase to another as a result of changes in *temperature* with pressure remaining constant.
- Vertical lines represent familiar transitions from one phase to another as the temperature remains fixed, but the *pressure* is allowed to change as a result of changing the container size (e.g., by moving a piston).
- A **critical point** is shown at the temperature and pressure at which the distinction between a gas and a

liquid is no longer apparent, and the two phases exhibit the same properties.

WATER AND ITS PHASES AND PHASE TRANSITIONS

The phase diagram for water, one of the world's most common substances, is shown in Figure 20.

Notice that normal melting point, T_m (0°C), and boiling point, T_b (100°C), can be defined as the temperature at which the two-phase equilibrium in question (i.e., curved line) has a pressure equal to atmospheric pressure, which is defined as 1.0 atmosphere. The critical point for water is at a very high temperature (374°C) and pressure (218 atmospheres).

The unusual feature of the phase diagram for water is that as pressure increases, the melting point is reduced below 0°C. The backwards sloping line for the solid-liquid equilibrium rising from the triple point on the diagram in Figure 20 shows this. This effect is easily demonstrated by squeezing two ice cubes firmly together. The ice cubes start to melt where they touch, but if you relax the pressure, then the liquid that was formed re-freezes and welds the two cubes together. This phenomenon also explains why sharp skates glide easily over an ice rink surface.

The effect is a consequence of the strong hydrogen bonds between water molecules that create a rigid open structure for ice. Solid water (ice) has a slightly lower density than liquid water just above the melting point. As a result, ice floats on cold water, which is a very important physical property that has allowed for life to survive on Earth during glacial periods of the Earth's history as well as in the present day. Just imagine if lakes froze from the bottom up! Almost no other substance shows this effect of an increase in pressure lowering the melting point of its solid phase.



Carbon dioxide has a more normal phase diagram than water. Your research into carbon dioxide and its phase diagram should be directed to answer the following questions:

- 1. Why is it so difficult to observe liquid carbon dioxide?
- 2. Why is super-critical carbon dioxide becoming an important commercial substance?



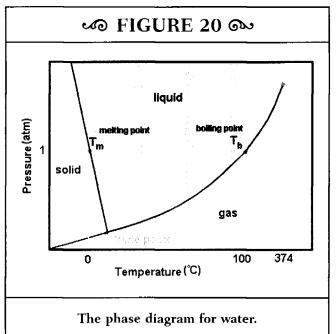


IMAGE COURTESY OF THINKQUEST, ORACLE EDUCATION FOUNDATION. http://library.thinkquest.org>.

Solutions

Solutions are mixtures of substances that are **homogenous**, meaning that no change in properties can be detected in different parts of a sample, no matter how you measure it. In our discussion here, we will only consider solutions consisting of just two substances mixed together. Obviously many solutions can contain more than two substances. For example, lemonade has water, lemon juice, and sugar.

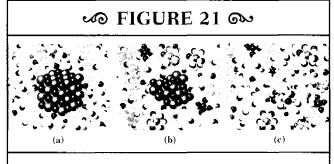
For any solution it is standard to define the substance that is present in the largest amount as the **solvent** and the other substance as the **solute**. We will consider two types of solutions in this section: **aqueous solutions**, in which the solvent is water, and organic solutions, in which the solvent is a non-polar substance such as octane. Most of the solutions considered will feature a solid solute dissolved in a liquid solvent.

Important parameters for all solutions include the following:

- **Solubility** the amount of solute that will dissolve in a certain amount of solvent
- Concentration the relative amounts of solute and solvent present; you will learn that there are several different ways to express the concentration of a solution.
- Changes in physical properties that occur as a consequence of forming a solution.

TYPES OF SOLUTIONS AND FACTORS AFFECTING SOLUBILITY

Chemists generally distinguish between two types of solvents: those that have polar molecules that interact strongly with electrically charged or polar solutes; and those that are non-polar and interact better with non-polar solutes. A very



A depiction of strong hydration.

IMAGE FROM CHEMISTRY: THE CENTRAL SCIENCE, LITH EDITION, BY THEODORE E. BROWN, ET AL., PUBLISHED BY PEARSON PRENTICE HALL, ©2009.

general rule of solubility is "like dissolves like." Water is the most important solvent of the polar type, as the O-H bond is strongly polar. Ionic compounds, such as **salts**, and nonionic substances that are also polar, such as sugar, dissolve well in water. A non-polar solvent such as octane dissolves other non-polar compounds such as vegetable oil.

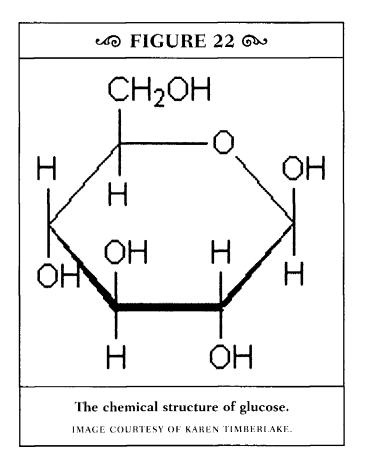
When salts such as sodium chloride dissolve in water, the actions between the water molecules and the ions in the crystal lattice are so strong that the lattice is disrupted, and the positive and negative ions are separated, with each ending up with a sheath of water molecules around it. Chemists call this interaction with water strong **hydration**. This process is illustrated in Figure 21.

A salt crystal surrounded by water (a) is disrupted by interaction with water molecules (b) until the salt ions are completely separated and (c) each ion is surrounded by water molecules.

Most ionic salts are **soluble**. A few ionic salts are **insoluble** since the forces of hydration are insufficient compared to the lattice forces of such salts. Following is a common set of solubility rules that can be applied to many substances:⁷

- (a) All common compounds of Group I and ammonium ions are soluble.
- (**b**) All nitrates, acetates, and chlorates are soluble.
- (c) All **binary** compounds of the halogens (other than F) are soluble with metals, except those of Ag, Hg(I), and Pb. Pb halides are soluble in hot water.
- (**d**) All sulfates are soluble, except those of barium, strontium, calcium, lead, silver, and mercury (1).
- (e) Except for rule (a), carbonates, hydroxides, oxides, silicates, and phosphates are insoluble.
- (f) Most sulfides are insoluble except for calcium, barium, strontium, magnesium, sodium, potassium, and ammonium.

Note that these solubility rules will be important when we begin to predict the results of possible **precipitation** reactions in Section IV of this resource guide.



AQUEOUS SOLUTIONS

In aqueous solutions, each solute, whether classified as soluble or insoluble, reaches a maximum amount that can be dissolved depending upon the temperature. A solution containing exactly the maximum amount of a solute is called a **saturated solution**. If additional compound is added to a saturated solution, it does not dissolve. Since solubility for most substances increases with increasing temperature, if a hot saturated solution is cooled, less solute can stay in solution, so the excess will form a solid. If the solute is a crystalline substance, solid crystals will form. If this is done slowly, large, attractively shaped crystals may be produced. This practice is the basis for crystal growing procedures.

Aqueous solutions are often encountered where more than one salt ion is present. If the ions present in solution happen to form an insoluble substance when combined, then a **precipitate** will form, assuming that the saturation point is exceeded. This is the basis for many tests for the presence of certain ions in water. For example, if water is to be tested for the presence of chloride ions (Cl⁻), adding some silver ions (Ag⁺) will create the insoluble compound silver chloride (AgCl), which will be formed as a precipitate on the bottom of the container.

A similar concept is involved in the formation of **stalactites** and **stalagmites** in caves. Water dripping into caves contains dissolved calcium (Ca^{2+}) ions. Carbon dioxide in the air in the cave dissolves in the water, producing carbonate ions (CO_3^{2-}) . The compound calcium carbonate $(CaCO_3)$ is insoluble and precipitates slowly on the roof or the floor of the cave to form stalactites and stalagmites.

We will learn more about the mixing of solutions that form precipitates in Section IV of this resource guide.

Compounds such as sugar (e.g., glucose) dissolve in water because they, too, contain O-H bonds. These substances are able to engage in hydrogen bonding with the water molecules, and so polar compounds such as sugars and alcohols are soluble in water. The structure of glucose is shown in Figure 22. Notice the five O-H groups, which make glucose similar to water (H-O-H) with five hydrogen bonding sites. Hence, glucose is very soluble in water.

ORGANIC (NON-POLAR) SOLVENTS

Non-polar solvents are generally compounds that contain only carbon and hydrogen. Carbon and hydrogen have similar values for electronegativity, so their bonds have very little polar nature. The intermolecular forces in non-polar solvents are rather weak. As a result, interactions with non-polar solute molecules are weak, and the substances dissolve easily since they can disperse throughout the solvent without forming strong clusters that would create separate liquid layers. A few other solvents fall into this category because their molecules are so symmetric that the dipoles in the bonds geometrically cancel, yielding a **non-polar molecule**.

These non-polar solvents are important for cleaning purposes, as many oils and greases are essentially non-polar and dissolve only in non-polar solvents. The dry-cleaning industry has depended on non-polar solvents for many years to remove grease and oil stains from clothes. A problem with the use of non-polar solvents is ultimately their disposal, as they are quite **toxic** to humans, and the only way to remove them from the environment is by burning them to form carbon dioxide and water.

Chemists have worked hard over many years to invent chemicals that will convert oils and greases to water-soluble compounds so that water may be used as a solvent for cleaning purposes. Soaps and detergents are the result of these efforts. Soaps and detergents have long and/or ring-like molecules that have one end that is relatively polar attached to another end that is relatively non-polar. The non-polar end interacts well with grease or oil molecules, while the polar end allows the molecules, dragging their greasy add-on with them, to dissolve more readily in the polar water. A recent development is the use of **supercritical fluids** as "super" solvents.

METHODS OF EXPRESSING CONCENTRATION

Scientists need to be able to communicate about solutions in a quantitative way so that properties can be precisely compared. There are several methods of expressing the relative proportions of solute and solvent.

Percent Composition

The simplest method to express the proportions of solute and solvent is by mass. The composition of a solution can be expressed as x grams of solute in y grams of solvent. Most often this is a percentage: x grams of solute in 100 grams of solution. For example, a bleach solution may be labeled as containing 5 percent bleach. This means that there are

5 grams of bleach molecule (NaOCl) in every 100 grams of bleach solution.

To be clear in some applications, this may be written as % w/w, where w stands for "weight," a substitute for mass. For example, a toothpaste tube may be labeled 0.5% sodium fluoride %w/w.

Molarity

In scientific applications, chemists are not interested in the grams present but in the moles present. Remember, the number of moles represents the number of molecules, not their mass. Since the solute molecules are the most important, it is the solute moles that are calculated (by dividing the mass of solute, in grams, by the molar mass of the solute, in grams per mole). The volume of liquids is expressed in liters, and so **molarity** is expressed as moles of solute in 1 liter (exactly) of solution. For example, a solution containing 5.85 grams of sodium chloride (molar mass = 58.5 g/mol) in 1 liter (exactly) of a salt solution has a concentration of 0.100 M. The M stands for the unit of molarity (moles/liter).

Molality

Some of the physical properties of solutions, such as boiling point or freezing point, are dependent on the relative numbers of moles of solute and solvent. These properties are called **colligative properties**. A unit known as **mole fraction** expresses this exactly. However, chemists decided they needed a simpler way to communicate how properties change with concentration. For a specific solvent, the mass present is a direct measure of the number of moles, and so the **molality** unit was developed by using a fixed mass of solvent. Concentration in molality units is the number of moles of solute dissolved in 1 kilogram (exactly) of solvent. Molality tracks physical properties in a linear manner. As an example, consider that if a 0.100 molal solution of sugar in water freezes at -0.186°C, a solution that is 0.200 molal will freeze at 2x (-0.186) = -0.372°C.



RAOULT'S LAW AND COLLIGATIVE PROPERTIES

Colligative properties of solutions are properties that depend on the amounts of substances present, i.e., the concentrations of solutions. These properties include boiling point, freezing point, and osmotic pressure. The French chemist François Raoult (1830–1901) formulated specific rules to explain this behavior that became known as Raoult's Law.

In your independent research, you should investigate Raoult's Law and find out how it is used to explain colligative property changes when a nonvolatile solute, such as sugar or an ionic salt,

is dissolved in a solution. Specifically, you should be able to explain how the solute changes the temperatures at which the solution boils or freezes. You should also be able to explain how and why a salt such as calcium chloride (CaCl₂) has a greater effect, per mole of dissolved solute, than a salt such as NaCl. You should also learn how to explain why table salt (NaCl) can be used in cold climates to reduce ice on roads, and you should also find out about some newer and better alternatives to this approach to ice reduction.

Additional research on solutions should lead you to understand and be able to explain how and why distillation works as a method for purifying water. What are some advantages and disadvantages of this method for use in the desalinization of seawater?



Section 111 Summary

- All matter exists in three common states: gas, liquid, or solid.
- These three states of matter each have distinctive properties.
- Gases were studied in the sixteenth and seventeenth centuries, and the properties of gases were found to follow simple rules, or laws. These gas laws are associated with the names of the scientists who worked them out.
- Boyle's Law states that the volume of a gas is inversely proportional to its pressure.
- Charles' Law states that the volume of a gas is directly proportional to its temperature.
- An examination of Charles' Law leads to the notion of an absolute zero temperature at which gas volume would theoretically be zero.
- The temperature scale derived from this idea is known as the Kelvin scale.
- Dalton's Law sets rules for dealing with mixtures of gases.
- The Kinetic Molecular Theory provides a theoretical basis for explaining all of the gas laws. It relies on the idea that a gas consists of very small rapidly moving particles.
- Amedeo Avogadro proposed that under given conditions of temperature and pressure, every gas contains the same specific number of small particles.
- Chemists use the unit of a mole to express an Avogadro's number of particles.

- There are 6.022×10^{23} particles in 1 mole of any substance.
- The average speed of molecules in a gas depends only on the temperature of the gas and the mass of the gas particles.
- Relative effusion and diffusion rates can be calculated if the mass of the gas particles is known.
- Liquids and solids differ from gases in regard to the freedom of motion of the particles of which they are comprised.
- Intermolecular forces between the particles in liquids and solids create an ordering that doesn't exist in gases.
- Solids have a very specific ordering of the particles (ions, molecules, or atoms) of which they are comprised.
- Some substances, such as carbon, can exist in a variety of solid structural forms, each with very distinctive properties.
- Metals have highly ordered structures that give rise to their unique physical properties, such as luster, malleability, and good conduction of heat and electricity.
- Transitions between the various states, or phases, of a sample of a pure substance can be summarized in a phase diagram that shows how the individual phases are in equilibrium as they transition from one to another.
- The processes of melting, freezing, condensing, vaporization, sublimation, and deposition can be explained using a phase diagram.
- Water has unique physical properties that can be explained by its phase diagram and its hydrogen-bond-type intermolecular forces.
- Solutions are formed when a solute dissolves in a solvent.
- Solubility depends on the structure of the solute and the solvent in terms of the polarity of their molecules. A general rule is "like dissolves like."
- Solubility "rules" exist for ionic compounds in water because certain combinations of ions may interact to form precipitates, reducing solubility.
- Solutions have important applications in our society; for example, they have many uses as cleaning products.
- How much solute dissolves in a solvent can be expressed in several different ways: including percent composition, molarity, or molality.
- Physical properties of solutions, called colligative properties, can be examined qualitatively, for example, to explain why adding salt raises the boiling point of water and lowers its freezing point. The physical properties of solutions also can be made quantitative by using Raoult's Law.



Section IV:

Reactions



Introduction

Themistry is all about reactions. Reactions occur when atoms in chemical substances rearrange under certain conditions to form new substances that have different physical and chemical properties. In this section of the resource guide, you will learn how chemists classify different reaction types and how they keep track of the atoms involved in the changes. You will also learn about the conditions that determine whether reactions allow for the complete reorganization of atoms, or only partly so. You will also learn how certain conditions determine how fast reactions will take place and how chemists track energy changes. Lastly, you will learn about energy changes since they enable chemists to predict which reactions will occur and which will not.

Types of Reactions

There are five basic types of reactions that we will study here:

- 1. Synthesis
- 2. Decomposition
- 3. Double replacement
- 4. Single replacement
- 5. Combustion

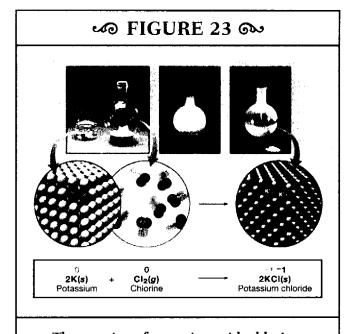
These different types of reactions are distinguished by the pattern of atom rearrangement that takes place. Learning about these patterns will be helpful when you need to predict the products in a reaction. Within these patterns, reactions may be further classified according to the properties of the substances that are reacting or the products that are formed.

Before we look at the patterns of atom rearrangement that can occur in reactions, we need to understand chemical equations. Chemical equations track the atoms in chemicals and the manner in which they are rearranged during a reaction. In chemical equations, reacting substances are indicated on the left-hand side of an arrow. The arrow means "yields" or "to make." The products of the reaction are found on the right-hand side of the arrow.

You might expect an equation to have an equals sign, and in fact, reaction equations used to have equals signs, but chemists decided that it was better to use an arrow to indicate the direction of the reaction. However, these reactions are still referred to as equations because the number and type of atoms on the left side must equal the number and type of atoms on the right side. For example, the equation $A + B \rightarrow C + D$ indicates that A and B are reactants and C and D are products, or what is made in the reaction.

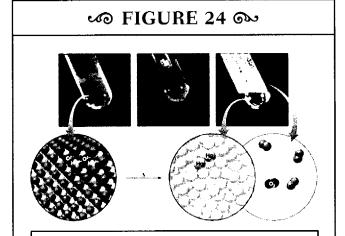
Synthesis, or combination, reactions involve combining two or more reactants to make one product. The equation for a synthesis reaction will have the form $A+B \rightarrow C$. An example of this type of reaction is $2Na+Cl_2 \rightarrow 2NaCl$. A different example is shown in Figure 23, which depicts the reaction of potassium with chlorine.

In **decomposition** reactions, one reactant breaks into several products, taking the general format $A \rightarrow B + C$. One example of this type of reaction is $2H_2O_2 \rightarrow 2H_2O + O_2$.



The reaction of potassium with chlorine.

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2HgO(s) 2Hg(I) + O₂(g)
Mercury (II) oxide Mercury Oxygen

A decomposition reaction in which mercury oxide is broken down into mercury and oxygen.

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FIGURE 24 depicts another decomposition reaction that several chemists used to make oxygen gas when it was first identified in 1774.

Double replacement reactions involve two ionic compounds in solution trading ions. Double replacement reactions have the general formula of AB + CD \rightarrow AD + CB. Pb(NO₃)₂(aq) + 2KI(aq) \rightarrow PbI₂(s) + 2KNO₃(aq) is an example of a double replacement reaction, and this reaction is shown in Figure 25.

Single replacement reactions have one atom moving into a compound and another atom moving out of the compound. Single replacement reactions have two different formats, depending on whether the atom is from a metallic element or a nonmetallic element. If the incoming atom is a metal, M, the format is $M + BC \rightarrow MC + B$. If the starting element is a nonmetal, N, the format is $N + BC \rightarrow BN + C$.

An example of a metal single replacement reaction is $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$. An example of a nonmetal single replacement reaction is $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$. Metal replacement reactions are quite common. A metal replacement reaction with copper wire being added to silver nitrate is pictured in Figure 26.

FIGURE 27 shows the addition of lithium metal to water. In this case, the lithium displaces H⁺ in the water, and elemental hydrogen gas results.

So how does one predict this type of reaction? Familiarity with the **activity series** of the metals is needed. The activity series shown in Figure 28 indicates which reactions will take place. In the reactions shown previously, copper was able to take the place of silver ion. The activity series shows



A double replacement reaction.

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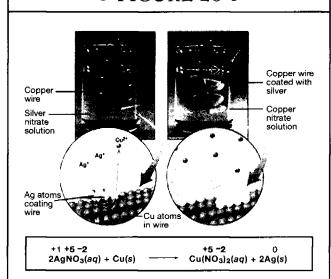
that copper is above silver, which indicates that it is more reactive. If silver metal were to be put in a copper solution, no reaction would occur. The series also indicates which metals can displace H⁺ from water and form hydrogen gas. This means that any atom above hydrogen can displace hydrogen from its source.

Combustion reactions are reactions of a reactant substance with oxygen gas (O_2) . If the reactant is an organic hydrocarbon, then the complete combustion products will be water and carbon dioxide, as is shown here in the combustion of methane (natural gas):

$$CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$$
.

To understand all chemical reactions, we must look at the properties of the substances that are reacting. While there are many possible chemicals that react, here we will learn about three common types of substances and how they react by examining three types of reactions: acid-base reactions, precipitation reactions, and **oxidation-reduction** reactions.

ശ FIGURE 26 യ



A metal replacement reaction with copper wire being added to silver nitrate.

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Acid-Base Reactions

This section introduces **acids** and **bases**, and the concept of **pH**. There are several theories of acids and bases, but here we will only discuss those introduced by Svante **Arrhenius** and Johannes Brønsted and Thomas Lowry (the **Brønsted-Lowry acid-base theory**).

ARRHENIUS ACIDS AND BASES

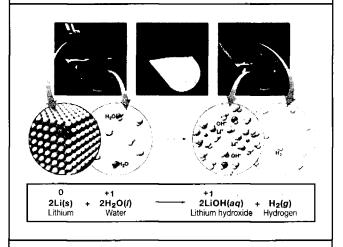
The first theory concerning acids and bases was developed by Svante Arrhenius, who stated that acids break apart in water to yield H⁺ (hydrogen ion) and that bases break apart in water to yield OH⁻ (hydroxide ion). However, this theory was incomplete, as it could not explain why, for example, tests would show that NH₃ (ammonia) solution was basic despite the fact that ammonia does not contain OH⁻.

BRØNSTED-LOWRY ACIDS AND BASES

Brønsted and Lowry took a different approach from Arrhenius and looked at how a substance reacted. Under their theory, an acid is defined as any reactant that donates an H^+ in a reaction, and a base is defined as any H^+ acceptor in the reaction.

This could explain NH_3 in water. In the reaction NH_3 + $H_2O \rightarrow NH_4^+$ + OH^- , the water donates an H^+ to the ammonia, making water an acid and ammonia a base. Ammonia becomes a positive ammonium ion, and the other product is OH^- . The solution has an excess of OH^+ ions and thus is basic. The Brønsted-Lowry model also explains all Arrhenius acids and bases.

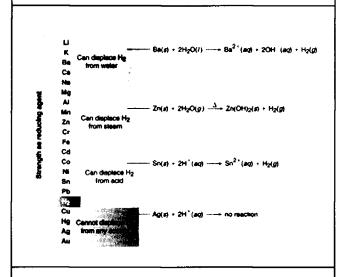
ൃ FIGURE 27 യ



A reaction involving the addition of lithium metal to water.

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∽ FIGURE 28 യം



The activity series for metals.

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In other situations, water can behave as a base by accepting an H^+ and forming H_3O^+ (**hydronium ion**). For example, when acetic acid ($HC_2H_3O_2$) is added to water to make vinegar, the reaction that occurs is: $HC_2H_3O_2 + H_2O \to H_3O^+ + C_2H_3O_2^-$. The acetic acid has transferred an H^+ to water and is acting as an acid in solution. The water has acted as a base by accepting the H^+ and forming a negative

Before titration Point Point Point

 $(I) + M^{+}(aq) + X^{-}(aq)$

A depiction of a titration.

(aq)

IMAGE FROM CHEMISTRY: THE MOLECULAR NATURE OF MATTER AND CHANGE, 4TH EDITION, BY MARTIN SILBERBERG, PUBLISHED BY MCGRAW-HILL, © 2006. IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.

acetate ion. The excess of $\mathrm{H_3O^+}$ ions makes the solution acidic.

The reactions just described indicate that water can be either an acid or base depending on the situation. Because water demonstrates this behavior, we say that water is **amphoteric**.

THE ACIDITY SCALE: PH

 $(aq) + X^{-}(aq) + M^{+}(aq) +$

pH is a measure of the acidity or basicity of an aqueous solution. A logarithmic scale called the pH scale relates the acid concentration, which is defined as the concentration of H_3O^+ in units of moles per liter. The equation is $pH = -\log[H_3O^+]$, where the square brackets indicate concentration in the units of molarity. Chemists usually abbreviate $[H_3O^+]$ to simply $[H^+]$.

The pH scale generally ranges from 0 to 14, where any substance in the range below seven is acidic, which indicates that $[H^+]$ is greater than 1.0×10^{-7} , and so seven is neutral. Substances with a pH above seven are basic, so $[H^+]$ is smaller than 1.0×10^{-7} . This means that a solution with a low pH (acidic) has a lot of H^+ compared to pure water. Conversely, a high pH (basic) solution has very little H^+ .

Note that acid and base solutions always contain both acid and base. In acidic solutions, there is more acid than base, and in basic solutions, there is more base than acid. Neutral solutions (pH exactly = 7.00) have an equal amount of acid and base. This is because in a sample of pure water some of the water molecules actually transfer a few H⁺ ions, making OH⁺, but there are equal concentrations of H⁺ ions and OH⁺, both at 1×10^{-7} M, and thus pure water has a neutral pH = 7.0.

TITRATIONS—MIXING ACIDIC AND BASIC SOLUTIONS TO MEASURE THE CONCENTRATION OF AN UNKNOWN

When acids and bases are mixed together and the pH is measured as different amounts react, this is called a **titration**. In this method, pictured in Figure 29, a given amount of acid with an unknown molarity is placed into a flask, and a solution of a base of known molarity is added by using a **buret**. The long tube being used in Figure 29 is a buret, which has a device called a stopcock that opens and closes, allowing for the acid or base to be added by drops into the Erlenmeyer flask. Usually, the acid is in the flask, and the base is in the buret, but it can be done the opposite way with the base in the flask and the acid in the buret.

The buret allows for precisely measured volumes of a liquid to be introduced. When the moles of acid in the flask are exactly equal to the moles of added base, the reaction has reached the **equivalence point** or **endpoint** of the titration. To help identify this point, chemists use a colored **indicator**. As the reaction reaches the endpoint, the indicator, which is sensitive to a certain pH, changes color. Titrations, as well as the pH state of any solution, can also be performed by using an instrument called a **pH meter**.

Since the moles of acid and base are equal at the equivalence point, we can write the equation $M_aV_a = M_bV_b$, where $_a$ represents total acid, and $_b$ represents total base. M is the molarity of each substance, and V is the volume of each substance. The reason this has to be written for total acid content or H^+ is that sometimes acid substances are **diprotic**, such as H_2SO_4 , or **triprotic**, such as H_3PO_4 . A diprotic acid can provide two H^+ ions to the reaction for every mole of acid added, and a triprotic acid can provide up to three H^+ ions for every mole of acid added. These differences have to be accounted for in the equation.

Let's look at an example. If 50.0 mL of HCl of unknown molarity were titrated with 25.0 mL of 1.50M NaOH, the equation would have $V_{\rm a}=50.0$ mL = 0.0500 L; $M_{\rm b}=1.50~M$; and $V_{\rm b}=25.0$ mL = 0.0250 L. So, $M_{\rm a}$ (0.0500L) = 1.50M (0.0250L). The unknown quantity is $M_{\rm a}$, which means the acid concentration must be 0.750M.

Acid-base reactions are a type of double replacement reaction where an acid + a base yields a salt + water. A salt is the compound formed when the H^+ of an acid is replaced by a metal ion or its equivalent. In the salt NaCl, the H^+ in HCl has been replaced by the metal ion Na $^+$. Many acids can be recognized as substances whose formulas start with atoms of hydrogen, such as HCl, HBr, H_2SO_4 , $HC_2H_3O_2$, and HNO_3 .

In water solutions, the hydrogen atom is transferred to water as a positive ion, H^+ . Strong acids are those that will transfer all of their H^+ to water. Weak acids are those that will transfer only a few molecules of H^+ to water. A base is any substance that accepts H^+ ions. Many bases have an OH⁻ ion, and so $H^+ + OH^- \rightarrow H_2O$. Water is formed as a product in such an acid-base reaction.

The salts that are formed can be acidic, basic, or neutral. The salts that are neutral are those made from a strong acid and strong base. For example, NaCl is made by reacting

HCl and NaOH. KBr is a salt from the reaction of HBr and KOH.



Some salts test acidic or basic when dissolved in water (i.e., they have a pH that is not equal to 7). This is a result of the ions in the salt undergoing a secondary reaction with the water. This secondary reaction is called hydrolysis. Your research should explore which salts make acidic solutions and why, and which salts make basic solutions and why. Is there a specific pattern that can enable a chemist to predict whether a particular salt will yield an acidic or basic solution when dissolved in water?



Precipitation Reactions

Precipitation reactions are also a specialized type of double replacement reaction. In precipitation reactions, two solutions, each containing ions in water, are mixed, and one of the products will be a solid precipitate that forms in the solution, often appearing as a cloud. The precipitate may settle slowly to the bottom of the container. Occasionally, two products may both be precipitates. The ions that do not form precipitates are known as **spectator ions** because no change occurs to them during the reaction—hence they are simply watching what is going on like spectators.

Let's look at an example of a precipitation reaction. If silver nitrate solution is mixed with sodium chloride solution, a balanced equation can be written to represent the reaction:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

However, if the ions in the solution are separately represented, we can write

$$\Lambda g^{+}(aq) + NO_{3}^{-}(aq) + Na^{+}(aq) + Cl^{-}(aq) \rightarrow AgCl(s) + NO_{3}^{-}(aq) + Na^{+}(aq).$$

Then, we notice that the NO₃⁻(aq) and the Na⁺(aq) are the same on both sides of the equation, so these ions are the spectators while the AgCl is the precipitate. The *net* reaction omits spectators and so is written:

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

FIGURE 30 provides another example, showing different equations and the spectator ions.

To predict which ions form precipitates and which will be spectators, chemists use the solubility rules that we discussed in Section III of this resource guide. Here are the solubility rules briefly listed again:

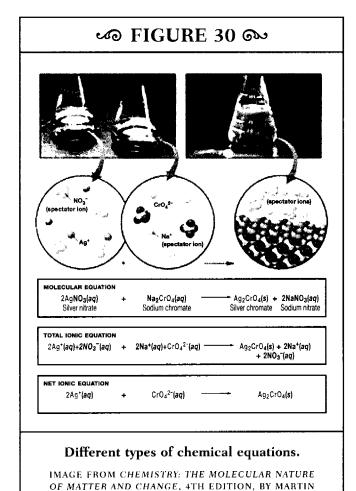
- Group I nitrates, acetates, and chlorates and ammonium ions are soluble, except for those listed previously.
 Carbonates, hydroxides, oxides, silicates, and phosphates are insoluble.
- 2. Ag, Hg(I), and Pb only form precipitates with chloride, bromide, and iodide.
- All sulfates are soluble, except those of barium, strontium, calcium, lead, silver, and mercury (I). Most sulfides are insoluble except for calcium, barium, strontium, magnesium, sodium, potassium, and ammonium.

Let's look at two examples that will demonstrate how these rules can be applied to a couple different reactions. If chloride ions are mixed with sodium ions, rule 1 suggests that no precipitate will form, and the chloride ions will remain as spectators:

$$Na^+ + Cl^- \rightarrow$$
 no reaction

However, if lead ions are mixed with chloride ions, rule 2 says that a precipitate of lead chloride (PbCl₂) will form:

$$Pb^{2+} + 2Cl^{-} \rightarrow Pb(Cl)_2(s)$$



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PRECIPITATES

Some precipitates are dangerous for the environment. In your independent research, you should aim to identify at least two precipitates that are hazardous, and you should learn how they could be removed from the environment. Some precipitates are useful. In your research, you should aim to identify and explain at least one way in which a precipitate is useful to people.



Oxidation-Reduction Reactions

Oxidation-reduction or **redox** reactions involve a transfer of electrons from one molecule to another. How and why do the electrons transfer? Every neutral atom has a number of electrons equal to its proton number. For instance, sodium has eleven electrons of which one is a valence, or outermost, electron. When the sodium atom changes to a sodium ion with a +1 charge and with only ten electrons, it has given up one electron and has undergone oxidation. Oxidation is the loss of electrons; reduction is the addition of electrons. The oxidation part of a reaction frees up electrons from one reactant that are then used in the reduction of the other reactant. For example, the freed electron from the sodium can then be used to reduce a silver ion with a +1 charge and forty-six electrons to silver with forty-seven electrons. An electron transfers from the sodium to the silver ion:

$$Na + Ag^+ \rightarrow Na^+ + Ag$$

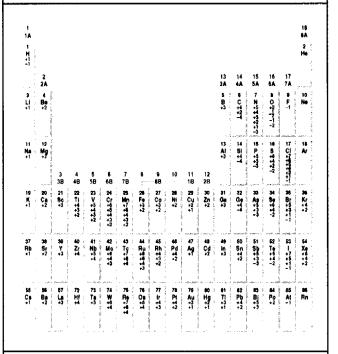
Oxidation-reduction reactions are the types of reactions that run batteries. To be able to determine what is occurring in these reactions, it helps to first identify oxidation numbers. These numbers help determine the number of electrons transferring in a reaction.

OXIDATION NUMBER

Oxidation numbers provide us with an accounting process for atomic electrons. The oxidation number for an element is equal to the number of electrons that must be added or subtracted from the atom in its bonded state to get to the number of electrons of its free, neutral atom. Oxidation numbers can be positive, negative, or equal to zero.

Let's look at some examples. In AgCl, Ag is +1 and Cl is -1, whereas in Al₂O₃, Al is +3 while O is -2. In the metal sodium, Na has an oxidation number of 0.





This table shows oxidation numbers for the elements.

IMAGE FROM CHEMISTRY, 7TH EDITION, BY RAYMOND CHANG, PUBLISHED BY MCGRAW-HILL, © 2002. IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.

Some standard rules help us to determine the oxidation numbers for different elements. Following are two simple rules that are true for nearly all compounds that they form:

- (1) II and group 1 elements have oxidation numbers of +1.
- (2) The oxidation number of O is -2.

The table shown in Figure 31 indicates possible oxidation numbers for the elements.

We can state three other rules concerning oxidation numbers:

- (3) All neutral atoms have an oxidation number of 0.
- (4) Oxidation numbers of atoms in neutral compounds must add up to 0.
- (5) The sum of the oxidation numbers of atoms in polyatomic ions must be equal to the charge on the ion.

In most cases, it helps to set up a mathematical relationship applying rules (3) through (5). For example, in the neutral compound potassium permanganate, KMnO₄, using rules (1) and (2), K is +1 and O is -2. So, to determine the oxidation number of the Mn, we can let Mn be x, and we can set up the equation: $1 + x + (4 \cdot -2) = 0$. Therefore, x = +7.

Once oxidation numbers are determined, then it is possible to determine how many electrons are involved in changing from one atom to another. For instance, if during

a reaction Ag^+ (Ag = +1) is changing to Ag (Ag = 0), as one electron must be added to reduce the positive ion (**cation**) to its atom.

THE ROLE OF THE ELECTRON IN OXIDATION-REDUCTION

In an oxidation reaction, the electron is freed and then is used in the reduction reaction. To balance equations for these types of reactions, it is important to note that the number of electrons freed must equal the number of electrons used. Electrons are conserved in nature. Even in a battery, which is a source of electrons, a complete circuit must deliver the electrons back after they have done any outside work or else the battery cannot operate.

To help construct oxidation-reduction reaction equations, half reactions are written to identify what is happening separately in the oxidation and reduction parts of the reaction and to identify how many electrons are being used. Then, you use a multiplying factor (if needed) to balance out the number of electrons added in one half reaction with the number subtracted in the other half reaction, and the half reactions are added together to make the reaction equation complete.

Let's take a look at an example where copper reacts with silver ion: $Cu + Ag^+ \rightarrow Cu^{2+} + Ag$. The two half reactions are as follows:

1.
$$Cu \rightarrow Cu^{2+} + 2e^{-}$$

2.
$$Ag^+ + 1e^- \rightarrow Ag$$

The second reaction must be multiplied by 2 to have a two-electron change, as that is how many electrons the first reaction is yielding, so we have the following:

3.
$$2Ag^+ + 2e^- \rightarrow 2Ag$$

Summing equations 1 and 3 gives us the following balanced equation:

$$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$$

Electrochemistry

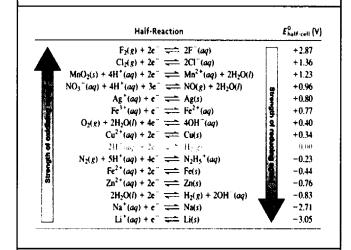
INTRODUCTION

Electrochemistry involves the use of the oxidation–reduction reactions we have just introduced. In electrochemistry, we wish to understand what governs which reactions will occur and how this can be used, as for example in the process of **electroplating**. Perhaps the most common example of electroplating is **chromeplating**, which is used to create shiny parts on appliances and automobiles.

CELL POTENTIALS

The tendency of an electron to leave an atom or to join an atom is measured as an electrical energy in units of **volts**. Since this energy can be measured even when an actual transfer has barely taken place, this has always been called a **cell potential**. It represents the potential energy that could do some work. This concept is the basis for a battery. Until the battery is used, it retains its potential or **voltage**.

∽ FIGURE 32 യെ



Selected Standard Electrode Potentials.

IMAGE FROM CHEMISTRY: THE MOLECULAR NATURE OF MATTER AND CHANGE, 4TH EDITION, BY MARTIN SILBERBERG, PUBLISHED BY MCGRAW-HILL, © 2006. IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.

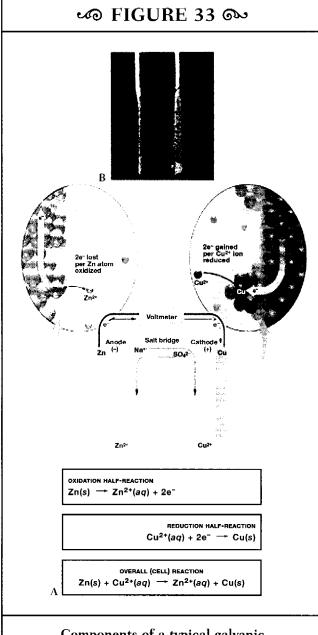
When all of the reaction has been completed and the system reaches chemical equilibrium, the battery has run down (is dead).

Each half reaction described in the previous section can be assigned a potential. Such a potential can be found in a table of reduction potentials, where the direction of the reaction is assumed to be electron gain or reduction. These are actually relative voltages that have been measured under the standard conditions of a 1 molar solution of the appropriate ion and at 1 atmosphere of pressure. All standard half-cell potentials are measured relative to the reaction $2H^+ + 2e^- \rightarrow H_2$ (gas).

A brief table showing standard reduction potentials is shown in Figure 32. Since all of the reactions are reductions, the reaction that is oxidation is just the reverse of these, and the potential of the cell is the same, but the sign is changed. For example, if the oxidation reaction is $Cu \rightarrow Cu^{2+} + 2e^-$, then the E° (standard reduction potential) of the half cell is -0.34 V.

Half cell potentials are combined to provide information on a complete reaction. The potential is not affected by the balancing of the reaction equation. The possible voltage is the difference between the potential of oxidation and the potential of reduction. If the total voltage is a positive value, then the reaction will proceed in the forward direction (from left to right as written). If the voltage is negative, then the reaction will proceed in the reverse direction, or from product to reactant.

Let's take a look at some examples that illustrate this. For the reaction $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ described earlier, the voltage of the reaction is determined by first identifying the potential of each half reaction:



Components of a typical galvanic electrochemical cell.

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$$Cu \rightarrow Cu^{2+} + 2e^{-}$$
 $E^{\circ} = -0.34 \text{ V}$
 $Ag^{+} + 1e^{-} \rightarrow Ag$ $E^{\circ} = +0.80 \text{ V}$

The sign of the Cu half-reaction voltage was changed since it was the oxidizing half reaction. Now, the two half reactions are added together (-0.34 + 0.80) for a complete cell voltage of + 0.46 V. Since this is a positive number, this is a naturally occurring reaction or, as chemists say, **spontaneously** occurring.

An example of a **nonspontaneous** reaction is $Li^+ + Fe^{2+}$ $\rightarrow Li + Fe^{3+}$. The two half reactions are as follows:

$$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$$
 $E^\circ = -3.05 \text{ V}$
 $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^ E^\circ = +0.77 \text{ V}$

The combination of the two half reactions (-3.05~V+0.77~V) gives us a cell voltage of -2.28~V. This reaction will occur naturally only in the reverse direction, or Li + Fe³⁺ \rightarrow Li⁺ + Fe²⁺ when the cell voltage is +2.28 V.

FIGURE 33 illustrates the components of a typical **galvanic electrochemical cell** and shows how its voltage could be measured. In a galvanic cell, the electrodes at which the two half reactions are occurring are identified by the terms **anode** and **cathode**. The anode is the electrode where oxidation occurs, and the cathode is the electrode where reduction occurs.

This oxidation half-cell has a potential of +0.76 V which combined with the reduction half cell of +0.34 V gives the overall cell potential of +1.10 V. This is a positive value, so this reaction will happen spontaneously under standard conditions. This cell was invented by J. F. Daniell in 1836 and was actually the earliest reliable battery.

Spontaneous reactions are clearly quite useful in providing us with batteries to power our electrical devices. Are non-spontaneous reactions ever useful? Yes, indeed! An entire industry of electroplating has developed around using cells driven in reverse by applying an external voltage. Such a cell is known as an electrolytic cell, and the process happening in it is called **electrolysis**. In these reactions, the external voltage of the source must be larger than the potential of the cell that must be driven "in reverse." As an example, let's consider a reaction used in zinc plating: the reaction $Zn \rightarrow Zn^{2+} + 2e^-$ where $E^\circ = -0.76$ volts must be driven in reverse to plate out solid zinc from a Zn^{2+} solution.



Electroplating is used in many industries. In your independent research, you should identify two different industries that use electroplating and explain how they do it. You should understand how much energy is needed and be able to show a calculation indicating how the energy added can be used to determine the number of grams of material that is electroplated.





The discussion in this section assumed that all reactants and products were in their "standard states." Walther Nernst (1864–1941) developed a way to make predictions about oxidation-reduction reactions. An equation known as the Nernst equation resulted from his work. Your independent research should lead you to learn about the Nernst equation and explore examples of how it is used to predict electrochemical phenomena under any reaction conditions.



Stoichiometry

In the kitchen, when a cook doubles a recipe to make twice as many cookies, he/she is practicing **stoichiometry**. We have already been using stoichiometric principles in the earlier sections of this resource guide, as we tracked the numbers of atoms taking part in chemical and physical processes. This is similar to what an accountant at a bank does in keeping track of dollars going into or out of a bank account. In this section, we will review the principles of stoichiometry and illustrate how important they are in helping us to determine the amounts of substances formed in chemical reactions. Stoichiometry uses a balanced equation to keep track of chemical substances changing from one type of compound to another type of compound.

A balanced equation is just a recipe that keeps track of what is put together as ingredients (reactants) and what is made in the end (products). The units used are moles. In an earlier section, you learned that chemists invented mole units to keep track of numbers of atoms or molecules. When hydrogen burns in air to form water, the balanced equation is:

$$2H_2 + O_2 - 2H_2O$$

This "translates" to: 2 moles of hydrogen combine with 1 mole of oxygen to make 2 moles of water. If you only had 1 mole of hydrogen, then only ½ mole of oxygen is needed to react, and only 1 mole of water would be formed. This is an example of how stoichiometry can be used. Now, we will review a few specific applications of stoichiometric principles.

IONIC AND MOLECULAR SPECIES PRESENT IN CHEMICAL SYSTEMS: NET IONIC EQUATIONS

If a chemical substance is an ion in solution, it is shown as an ion and not a compound. For instance, in solution, NaCl is $Na^+(aq)$ and $Cl^-(aq)$. However, the numbers (moles) of positive and negative charges in any solution must be equal, so $MgCl_2$ in solution is Mg^{2+} and 2Cl. **Net ionic equations** are written to eliminate any spectator ions. (You will recall our earlier discussion of spectator ions in precipitation reactions.) For molecular and ionic species, the number (of moles) of atoms of each chemical type in reactants (left side of an equation) must be equal to the numbers of atoms in the products (right-hand side). Atoms are always conserved in chemical reactions.

Let's look at an example showing equations for the reaction that occurs when solutions of silver nitrate and sodium sulfide are mixed to yield a solution of sodium nitrate and a precipitate of silver sulfide:

$$2AgNO_3(aq) + Na_2S(aq) \rightarrow 2NaNO_3(aq) + Ag_2S(s)$$

The ionic equation is the following:

$$2Ag^{+} + 2NO_{3}^{-} + 2Na^{+} + S^{2-} \rightarrow 2Na^{+} + 2NO_{3}^{-} + Ag_{5}S(s)$$

Canceling spectator ions leaves the following balanced ionic equation:

$$2Ag^+ + S^{2-} \rightarrow Ag_2S(s)$$

BALANCING REDOX EQUATIONS

When redox (i.e., oxidation-reduction) equations are balanced by using the half-reaction method, the electron change of the two half reactions must be balanced as well. This was explained previously in our discussion of single replacement reactions.

MASS AND VOLUME RELATIONS

Compounds can be represented by either molecular or **empirical formulas**. Empirical formulas provide the lowest mole ratio of atoms to each other in a formula. For example, acetylene, C_2H_2 , could be CH. However, chemists know that CH doesn't actually exist as a stable molecule, whereas C_2H_2 does, even though the mathematical relationship is the same. Determining this from empirical data is a common experiment in chemistry.

Chemical equations and stoichiometry are used to determine **yields** in more measurable units by using molar masses. For instance, using the example given previously, $2H_2 + O_2 \rightarrow 2H_2O$, we can state that "4 grams of hydrogen react with 32 grams of oxygen to make 36 grams of water." Note that this restatement uses the established facts that the molar mass of H = 1, and the molar mass of O = 16. Simple math can then be used to establish that $2H_2$ has the mass of 4 moles of H, that H0 has the mass of 2 moles of H0, and that molar masses are additive in calculating that water has a molar mass of 18.

A very common problem in chemistry is determining a "limiting reactant" where exact stoichiometric amounts are not present. This is also easy to relate to cooking. If a recipe calls for four cups of flour, and you only have two cups, it will limit the amount of product, for example cake, that you can make. A chemist uses a balanced equation to determine how much of each ingredient (i.e., reactant) will be needed for a reaction and how much product can be made.

Let's consider the same sample reaction we just discussed: $2H_2 + O_2 \rightarrow 2H_2O$. If there are initially 12 g of H_2 and 32 g of O_2 , which reactant is limiting? And, how much water product can be made?

ANSWER: Since balanced equations are in moles, 12 g of 11_2 must be changed to moles by dividing by its molar mass of 2. This gives us 6 moles of H_2 . Doing the same for the O_2 , we have 32 divided by the molar mass of oxygen, which is 32. This gives us 1 mole of O_2 . With 6 moles of H_2 , the balanced equation indicates that 3 moles of O_2 would be needed for a complete reaction since the ratio of H_2 to O_2 is 2:1. Since we only have 1 mole of O_2 , there is not enough O_2 to react with all of the H_2 . This means that 4 moles of H_2 will be left over since only 2 moles of H_2 can react with the 1 mole of O_2 . Therefore, the H_2 is said to be in excess, and the O_2 is said to be the "**limiting reagent**." The total amount of product that can be made, called the reaction yield, will be 2 moles of water, or $2 \times 18 = 36$ grams.

Equilibrium CHEMICAL EQUILIBRIUM

Many reactions do not go to completion but instead start to reverse as product is made. So while reactant turns into product, some of the product turns back into reactant. Obviously this is important, as the amount of chemical you find you have when the reaction has apparently stopped might not be the same amount that we calculated in the previous section.

For our example reaction $A+B\to C+D$, this means that soon after the reaction starts in the forward direction as written above, some products may start to react together in the reverse process, $C+D\to A+B$. Chemists often combine these two reactions together into one equation by using a double-headed arrow to indicate the two directions in which the reactions can proceed: $A+B\to C+D$.

It is crucial to recognize that chemical reactions do not happen instantaneously. In the next section, you will learn about what controls how fast reactions take place. The speed, in terms of how much material reacts in each time unit (e.g., one minute) is known as a reaction rate. Most reactions will eventually reach a point when the rate of the forward reaction is the same as the rate of the reverse reaction. When this point occurs, the reaction is said to be at equilibrium. Even though both reactions are still occurring, the system appears not to change. This does not mean that the amount of any reactant left equals the amount of any product formed, but rather that the rates are equal.

EQUILIBRIUM CONSTANTS

Chemists found that, at equilibrium, if they multiplied the equilibrium concentrations of products together and divided this number by the mathematical product of the concentrations of the reactants, as long as certain conditions were unchanged, this ratio had a constant value. This number is known as the **equilibrium constant**, and the relation from which it can be calculated is known as the equilibrium constant expression.

For our typical reaction A + B + C + D, the equilibrium constant expression is as follows:

$$K = \frac{[C][D]}{[A][B]}$$

Here *K* is the equilibrium constant, and the symbols [A], [B], [C], and [D] represent the concentrations of the individual reactants and products at their equilibrium values. Each reactant or product molecule involved in reacting or being produced must be included in the expression.

So, for example, for the reaction $A + B \rightarrow 2C$ the *K* expression is:

$$K = \frac{|\mathbf{C}|[\mathbf{C}]}{|\mathbf{A}|[\mathbf{B}]}$$
, which simplifies to $K = \frac{|\mathbf{C}|^2}{|\mathbf{A}|[\mathbf{B}]}$

For the completely general reaction $aA + bB \rightarrow cC + dD$, the equilibrium expression is:

$$K = \frac{[\mathbf{C}]^{\mathbf{c}}[\mathbf{D}]^{\mathbf{d}}}{[\mathbf{A}]^{\mathbf{a}}[\mathbf{B}]^{\mathbf{b}}}$$

The constant, K, has a different numerical value for every different reaction. Chemists give special names and expressions to these equilibrium constants, depending on the type of reaction involved. If the reaction involves acids, then the equilibrium constant is called Ka. If the reaction involves a base, then it is called Kb. If the reaction involves forming a precipitate, then the constant is called Ksp. If it is a reaction involving pressures of gases, then the equilibrium constant is called Kp. If it involves solutions and concentrations, then the equilibrium constant is called Kc, although this is often just written K.

The value of the constant is an indicator of how far the reaction will proceed, moving forwards or backwards. If the value is greater than I, then the reaction favors product; the value of the mathematical product of the concentrations of products is greater than the value of the mathematical product of the concentrations of the reactants. In simple language, this means that when equilibrium has been reached, there is more product than reactant. If the value of *K* is less than one, then there is more reactant than product, and so the reverse reaction has been favored. This is true regardless of the type of equilibrium reaction.

QUANTITATIVE TREATMENT OF EQUILIBRIUM

To solve equilibrium problems, a general understanding of an equilibrium expression is needed. As a reminder, the [A] indicates the concentration of substance Λ in moles per liter (Molarity, M). Any substance that is a *pure* solid or *pure* liquid is dropped from the equilibrium expression. This is because its concentration *in itself* does not change or is said to have an activity of 1, and therefore including it doesn't change the value of the equilibrium constant. Thus, it is extremely important that phases be known and shown in balanced equations.

EQUILIBRIUM CONSTANTS FOR GASEOUS REACTIONS: Kp, Kc

As mentioned previously, Kp is a unique equilibrium expression for gas pressures, p, and is written slightly differently. For the same general equation $aA + bB \leftrightarrow cC + dD$,

$$Kp = \frac{p_{C}^{c} p_{D}^{d}}{p_{A}^{a} p_{B}^{b}}$$
, where P_{C} is the pressure of gas c .

There is an equation that allows us to convert from a Kc to a Kp value. This equation is $Kp = Kc(RT)^{\Delta n}$, where Δn represents the change in moles of gas in the balanced equation of (total products) – (total reactants). When the change in moles of gas is 0, then Kp = Kc.

EQUILIBRIUM CONSTANTS FOR ACIDS AND BASES

Because acid and base reactions involve H⁺ and OH⁻, calculations of the pH or acid content can also be performed, and *Ka*, the equilibrium constant for acids, or *Kb*, the equilibrium constant for bases, can also be evaluated. The constant itself lets us know the characteristic of the acid or base just by the size of the number—the smaller the number, the weaker the acid or base. Remember that when weak acids are in water, not all of the H⁺ goes into solution. The less H⁺, the smaller the *Ka* will be.

For example, when acetic acid is in water (in vinegar), the balanced reaction is $HC_2H_3O_2 \rightarrow H^+ + C_2H_3O_2^-$ for which the Ka expression is:

$$K_{\rm a} = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]},$$

and the Ka value is found to be 1.8×10^{-5} . This means that the amount of H⁺ is very small compared to the concentration of the unreacted acid (in the denominator), and so acetic acid is classified as a weak acid. The less the acid reacts, the smaller the Ka value. This means that the weaker the acid, the smaller the Ka value will be. The pH of a solution of a weak acid can be determined by using the Ka expression and the equation pH = $-\log[H^+]$. For a base reaction such as NH₃ + H₂O \leftrightarrow NH₄⁺ + OH⁻, the equilibrium expression is:

$$K_{\rm b} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Here, too, the amount of the $\mathrm{NH_3}$ (ammonia) that reacts is an indication of the strength of the base. If it does not react very much, the Kb value will be small. For ammonia, the Kb value is only 1.8×10^{-5} . It is a weak base. The weaker the base, the smaller the Kb value will be. This is another way that we can compare the strengths of acids and bases.

SOLUBILITY PRODUCT CONSTANTS AND THEIR APPLICATION TO PRECIPITATION AND THE DISSOLUTION OF SLIGHTLY SOLUBLE COMPOUNDS

Ksp equilibrium constant values indicate how much a solid will ionize in solution. For an insoluble substance this also indicates its solubility. For example, for the reaction $AgCl(s) \rightarrow Ag^{+}(aq) + Cl^{-}(aq)$, the Ksp expression is Ksp = $[Ag^{+}][Cl^{-}]$. Since the value is only 1×10^{-10} , silver chloride

is an insoluble substance, though the *Ksp* indicates a very small amount does indeed dissolve. The smaller the *Ksp* value, the more insoluble the compound is.

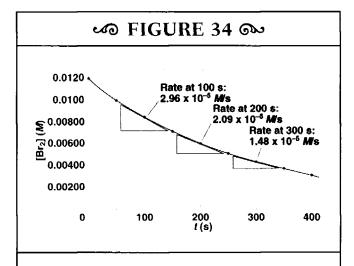
As mentioned in the earlier discussion of precipitation reactions, if ions of silver and chloride become mixed in the same solution, the small value for *Ksp* indicates that a precipitate will form. *Ksp* values can be used to predict exactly how much precipitate will form in a precipitation reaction.

Kinetics

Kinetics is the study of how fast reactions happen and what affects the rate at which the reaction proceeds. An algebraic equation called a **rate law** can be determined from studying experimental evidence about a reaction. A rate law shows how changing the concentration of a reactant affects how fast the reaction happens. There is no set rate law for all reactions; the rate law is different for different reactions.

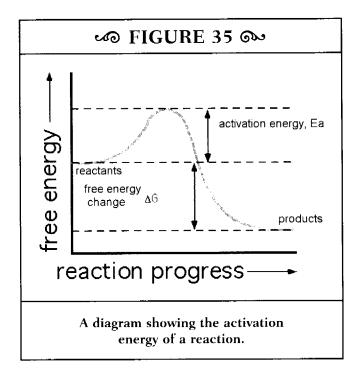
THE CONCEPT OF RATE OF REACTION

As a reaction takes place, there is a loss of reactant and a gain of product. The rate of the reaction can be determined by tracking the change in concentrations over time. In a car, the rate or speed is the distance traveled over a period of time. If this changes, one can use an average speed. The same can be done for a reaction, but instead of measuring miles traveled, with reactions we measure the change in concentration over a unit of time. This will give us the average rate of change. To determine an **instantaneous rate**, or the rate at one moment in time, a chemist can construct a graph as is shown in Figure 34. By making a line tangent to the curve at a point and finding the slope of the line, you get the instantaneous rate.



A graph showing instantaneous rates of reaction.

IMAGE FROM CHEMISTRY, 7TH EDITION, BY RAYMOND CHANG, PUBLISHED BY MCGRAW-HILL, © 2002. IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.



EFFECT OF TEMPERATURE, CONCENTRATION, AND SURFACE AREA CHANGES ON REACTION RATES

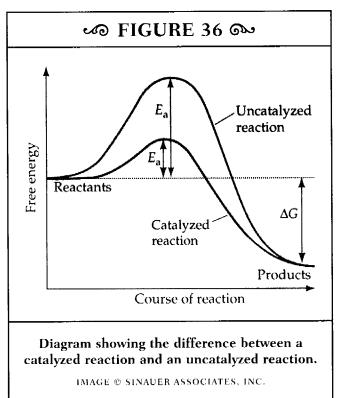
The best model for chemical reactions, called the **collision model**, is similar to the Kinetic Theory described earlier in Section II of this guide. Molecules must collide in order to react. Thus, if we examine how changing conditions affect the rate of collisions, we will be able to predict how the reaction rate will change. For example, every day we put foods in the freezer to keep them longer (i.e., slow down the rate of decomposition or decay), or we heat food to increase the rate of chemical change by cooking it. Increasing temperature generally increases reaction rates.

As another example, consider that an antacid tablet such as Alka-Seltzer® clearly dissolves much more quickly in hot water than it does in cold water. The tablet also reacts much more quickly if it is broken into small pieces. An increase in surface area increases the number of possible collisions per second with the water molecules.

Concentration affects rates of collision and reaction because if fewer molecules are present in a given volume, then it is less likely that they will collide and react.

REACTION POTENTIAL ENERGY DIAGRAMS, ENERGY OF ACTIVATION, AND THE ROLE OF CATALYSTS

For a successful reaction to occur, reacting molecules must not only collide, but the collision must also happen with a certain amount of energy. Each collection of chemically reacting molecules also has a specific energy before and after reaction. These energies are commonly shown on a **potential energy diagram**. As the reaction progresses, the energy changes are shown from left to right on the diagram. The "essential" energy of collision is called the **activation energy**, $E_{\rm a}$, and is shown in Figure 35.



The rate of reaction depends on the size of the activation energy. The higher the activation energy is, the slower the reaction will progress.

Some substances can be added to reactions to lower the required activation energy and thus speed up the reactions. These substances are called **catalysts**. A catalyst provides a lower energy pathway for the reaction. Imagine you need to get to a certain destination, and there is a mountain in front of you. Instead of climbing to the top to get to the other side, you might find a route through a lower pass between the mountains to the other side. This approach would take much less energy and would be faster. The diagram in FIGURE 36 shows that two reactions may begin in the same place and end at the same place as before, but the catalyzed reaction requires less energy.

There are many types of catalysts. Most chemical reactions in the human body use catalysts called enzymes to increase reaction rates at body temperature.

Thermodynamics

Measuring energy changes in chemical reactions is called **thermochemistry**. **Thermodynamics** is based on the ideas of energy and temperature related to the particle motion in the Kinetic Molecular Theory model and ultimately explains whether or not a reaction will happen. Because movement is so important, scientists have to define very carefully what they are studying. A part of the universe such as a number of moles of a chemical in a beaker is called the *system*. All of the rest of the universe is called the *surroundings*. To a thermodynamics expert: System + Surroundings = Universe. The movement of matter and/or energy between a

system and its surroundings must be understood in order for the laws of thermodynamics to make sense.

Thermochemistry allows us to determine the amount of heat released or taken in by a chemical reaction. **Endothermic** processes are reactions that have a net absorption of heat, whereas **exothermic** reactions have a net release of heat. Energy diagrams for reaction progress show this change. Reaction (a) in the diagram shown in Figure 37 is exothermic. Reaction (b) is endothermic. Note that a catalyzed reaction, which we discussed earlier, has the same overall loss or gain in thermal energy (heat released or taken in).

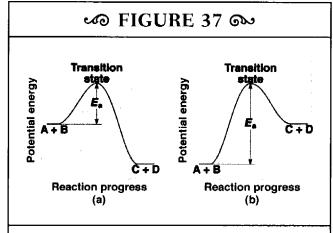
STATE FUNCTIONS

State functions are properties of a system that are only determined by where the system is, not how it arrived there. Examples of state functions are volume, V, energy content, E, and pressure, P. When a state function quantity changes, chemists use the Δ symbol to indicate the amount of change. It is also true that for ΔV , ΔE , and ΔP , the pathway is not important; only the initial state and the final state are important. It is usually easy to measure most Δ values, but it is not easy to establish the magnitude of state function parameters for a system.

THE FIRST LAW OF THERMODYNAMICS

Systems have an energy content that can be tracked by measuring the heat given off or taken in by a chemical reaction. This energy content, called **enthalpy**, can be assigned to individual chemical substances. Enthalpy is a state function and is given the symbol H. Changes in enthalpy are thus denoted ΔH . The enthalpy is directly proportional to the amount (in moles) of chemical present. If heat given off by a reaction is measured, the total enthalpy change per mole can be assigned energy units, usually kilojoules.

For example, if 1 mole of carbon (e.g., as coal) is changed to 1 mole of carbon dioxide, then 393.5 kilojoules of energy are given off as heat. This is known as a "heat of reaction."



Diagrams of an exothermic reaction (a) and an endothermic reaction (b).

IMAGE FROM CHEMISTRY, 7TH EDITION, BY RAYMOND CHANG, PUBLISHED BY MCGRAW-HILL, © 2002. IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.

Each compound can be assigned an enthalpy change of formation or **heat of formation** ($\Delta H_{\text{formation}}$) if 1 mole of the compound is made from its elements using standardized conditions.

The process by which these enthalpy changes are measured is called **calorimetry**. For example, to measure the heat of a reaction between an acid and a base, the change in temperature of the reacting mixture (ΔT) is measured. This change is compared to the change in temperature of a sample of the same mass of pure water as heat is added from a burner. Chemists have established that it takes 4.18 joules of heat to raise the temperature of 1.00 gram of water by 1.00 degree Celsius, so they do not have to re-measure this each time.



A law called Hess' Law allows us to calculate enthalpy changes that have not been measured based on our knowledge of other enthalpy changes that have been measured and using a cycle of reactions. In your independent research, you should find out about Hess's Law and learn how it is used.



THE SECOND LAW OF THERMODYNAMICS

The second law of thermodynamics involves the idea of **entropy** or chaos. Chemists found that the measurement of heat changes (enthalpy) alone was insufficient to account for everything that happened in chemical and physical changes of matter. So, chemists associated an energy content with the actual ordering process within and among the particles of matter. A new state symbol, *S*, was defined for entropy, which indicates the energy associated with disorder.

The smaller the entropy value is, the greater the order. All natural processes result in an increase in entropy for the universe, but not necessarily for the system. Remember: System + Surroundings = Universe. So, a decrease in entropy can occur in a system if a matching increase occurs in the surroundings. Solids are the most ordered state of matter; liquids are the next most ordered, then solutions, and lastly gases are the most disordered.

The final decision as to whether or not a chemical reaction will occur must be based on a consideration of changes in *both* enthalpy (ΔH) and entropy (ΔS). Changes in these state functions can be positive or negative, so signs are very important. J. W. Gibbs, an early American chemist, determined a simple relationship of enthalpy and entropy, and the resultant state function is named in his honor. This state function is **Gibbs Free Energy**, which has the symbol G.

∞ FIGURE 38 ∞

H	ΔS	ΔG	Example
+ +	+	Reaction proceeds spontaneously at high temperatures. At low temperatures, reaction is spontaneous in the reverse direction.	$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$
+	-	ΔG is always positive. Reaction is spontaneous in the reverse direction at all temperatures.	$3O_2(g) \longrightarrow 2O_3(g)$
-	+	ΔG is always negative. Reaction proceeds spontaneously at all temperatures.	$2H_2O_2(l) - 2H_2O(l) + O_2(g)$
-	-	Reaction proceeds spontaneously at low temperatures. At high temperatures, the reverse reaction becomes spontaneous.	$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

Factors affecting the sign of ΔG in the relationship $\Delta G = \Delta H - T \Delta S$.

IMAGE FROM CHEMISTRY, 7TH EDITION, BY RAYMOND CHANG, PUBLISHED BY MCGRAW-HILL, © 2002-IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.

We can now determine whether or not a chemical reaction will occur by using the equation connecting changes in G, H, and S, or: $\Delta G = \Delta H - T\Delta S$, where T is the temperature (in Kelvin units).

The sign of ΔG determines if a reaction (as written left to right) will be spontaneous or not. Because the signs of the other terms can be known, and temperature is always positive (+), we can establish the conditions needed for a reaction to be spontaneous, as is shown in the table in Figure 38. If ΔH is negative and ΔS is positive, then the reaction occurs spontaneously. But, if ΔH is positive and ΔS is negative, then the reaction is nonspontaneous. When both ΔH and ΔS have the same sign, whether or not the reaction is spontaneous depends upon the temperature.

Relationship of Change in Free Energy to Equilibrium Constants and Electrode Potentials

Free energy has a direct relationship to equilibrium constants and electrode potentials. If a reaction has a negative

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ΔG°	κ	E _{cell}	Reaction under Standard-State Conditions
Negative	>1	Positive	Favors formation of products.
0	= 1	0	Reactants and products are equally favored
Positive	<1	Negative	Favors formation of reactants.

#### Relationships among $\Delta G^{\circ}$ , K, and $E^{\circ}_{cell}$ .

IMAGE FROM CHEMISTRY, 7TH EDITION, BY RAYMOND CHANG, PUBLISHED BY MCGRAW-HILL. © 2002. IMAGE REPRODUCED WITH PERMISSION FROM MCGRAW-HILL.

 $\Delta G$ , then this indicates that the reaction is spontaneous in the forward direction. A positive galvanic cell voltage, E, also indicates that a reaction is spontaneous in the forward direction. As a reaction approaches equilibrium, its  $\Delta G$  and E will approach zero. At equilibrium,  $\Delta G$  and E will be equal to zero.

For prediction purposes and to compare different reactions, it is useful to consider values of  $\Delta G$  and E when reactants and products are all in their defined "standard states." Then, it is possible to say whether a reaction from that condition will have a large (>1), or small (<1) value for its equilibrium constant, and we can determine the sign of its cell potential. These results are summarized in the table shown in Figure 39.

The equation that allows you to change from free energy to equilibrium constants at standard states is  $\Delta G^{\circ} = -RT \ln K$ . And, the equation to change to standard cell potential is  $\Delta G^{\circ} = -nFE^{\circ}_{\text{cell}}$ . The negative signs allow for a change from the negative value of free energy to the corresponding positive values of K and E.

### Section N Summary

- The starting substances in reactions are called reactants, and the ending substances are called products.
- Synthesis reactions combine substances to make one product while decomposition reactions start with one reactant and break it apart.
- Double replacement reactions have two reactants that trade **anions** to make new products. Single replacements reactions have an atom and a compound as reactants and then trade to have a different atom and compound as products. Metal atoms can replace metal ions based on the activity series of the metals.
- Combustion reactions involve substances reacting with oxygen. If the reaction is with an organic molecule, the products will be carbon dioxide and water.

- The pH scale is a measure used to define the acidity of a water solution. Acidic solutions have a pH that is lower than seven, and basic solutions have a pH that is greater than seven. Arrhenius and Brønsted-Lowry theories can explain why some substances are acids and others are bases.
- The equation for determining the pH is pH =  $-\log[H_3O^+]$ , where  $H_3O^+$  is called hydronium or, simply, hydrogen ion concentration in pH =  $-\log[H^+]$ .
- An amphoteric substance can behave as either an acid or a base. Water is an example of an amphoteric substance.
- A titration is a method used when acids and bases are mixed together and the pH is monitored as different amounts react. A titration uses a piece of equipment called a buret to measure the volume of solution added
- In a titration, when the amount of acid equals the amount of base, the equivalence point or endpoint has been reached. This can be indicated by graphing, by using a pH indicator that changes color at the appropriate pH, or by using a pH meter.
- Salts are a product of acid-base reactions. Salts can be acidic, basic, or neutral depending upon the acid or base that was used to make them.
- Precipitation reactions involve two ionic solutions being put together and an insoluble product forming. The reactant ions that do not change form (remain as ions) are called spectator ions. There are solubility rules that help to predict whether or not a precipitate will form.
- Oxidation-reduction reactions, or redox reactions, involve the transfer of electrons. To determine which substance is freeing up electrons (being oxidized) for another reacting substance (being reduced) to use, the oxidation numbers must be determined.
- Electrochemistry uses redox reactions to perform many common types of useful reactions, including those used in batteries and electroplating.
- Cell potentials are the potential energy a reaction has to do work. Cell potentials are measured in volts. If the total voltage of a reaction is positive, then the reaction will be spontaneous.
- Oxidation occurs at the anode, and reduction occurs at the cathode in an electrolytic cell.
- When a cell will not run due to a negative potential, energy can be applied to make it run. This is called electrolysis.
- Using a balanced equation, stoichiometry is a method of converting from moles of one substance in the equation to moles of another substance in the equation. Conversions can then be made from moles to obtain masses in grams.
- Stoichiometry is used both to predict the yield of reactions and to determine limiting reagents. A limiting

- reagent is a substance that, due to its starting amount, will limit the amount of product that can be made.
- Reactions do not go to completion and instead run in both forward and backward directions. When the rate of the forward reaction is equal to the rate of the backward reaction, the reaction is said to have reached equilibrium.
- Reactions at equilibrium have a special number that can be calculated called an equilibrium constant, *K*. The formula for the equilibrium constant for the general reaction aA + bB cC + dD is:

$$K = \frac{[\mathbf{C}]^{c}[\mathbf{D}]^{d}}{[\mathbf{A}]^{a}[\mathbf{B}]^{b}}$$

- The type of reaction determines if the value is called a Ka (acid), Kb (base), Kp (gas), Kc (concentration), or Ksp (precipitate). The size of the number indicates which half reaction (forward or backward) is more favored to occur. If the constant is larger than 1, then the forward reaction is favored. If the constant is less than 1, then the backward reaction is favored. If the constant is equal to 1, then they are equally favorable.
- Reactions occur at different rates. Reaction rates are studied in the field of kinetics. All reactions have different factors that affect their rate. By studying reactions, scientists can determine rate laws that show how concentrations affect the rates of reactions.
- Kinetics is understood using collision theory, which states that molecules need to collide with each other with the right orientation and sufficient energy in order for a reaction to occur. Some substances, called catalysts, can speed up reactions by lowering the activation energy needed for the reaction.
- Thermochemistry measures overall energy changes in reactions. Reactions that absorb energy are called endothermic, and reactions that release energy are called exothermic.
- Calorimetry is a method of measuring the heat of a reaction.
- State functions are properties of systems where the pathway is not important, but the beginning and ending state are. Examples include temperature, volume, pressure, enthalpy, entropy, and free energy of systems.
- The energy content of a system is called enthalpy, and the disorder of a system is called entropy.
- Calculations of Gibbs Free Energy can be used to determine if a reaction will be spontaneous. If the Gibbs Free Energy value is negative, then the reaction will occur. If it is equal to zero, then the reaction is at equilibrium. And, if it is positive, then the reaction will not occur. Different relationships are used depending on the data that is available. If cell potential is known, then the equation is  $\Delta G^{\circ} = -nFE^{\circ}_{cell}$ . If the equilibrium constant is known, then the equation is  $\Delta G^{\circ} = -RT\ln K$ . And, if entropy and enthalpy are known, then the equation is  $\Delta G = \Delta H T\Delta S$ .



## Conclusion

Themistry is the science of matter, involving the study of how matter is structured and how those structures behave through chemical and physical changes. Chemistry has been both an art and a science. For thousands of years, creative people experimented with substances that occurred naturally on Earth and found that these substances could be changed to improve their appearance and to benefit human lives in both aesthetic and practical ways. Individuals, especially Greek philosophers, created theories to explain what was observed and came up with ideas about the inner structure of matter.

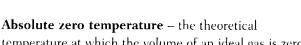
It wasn't until the seventeenth century that these early theories about the nature and structure of matter were further extended, as more precise methods of measurement became available. As a result, many "rules" by which nature operated became clear. In the eighteenth century, Antoine Lavoisier and others, working mostly in France and England, brought order to the rapidly expanding field of experiments and theories by proposing, among many other things, a nomenclature system for chemical compounds and a better explanation for combustion reactions.

Modern chemistry has since developed many methods of exploring the properties of natural materials and many ways to manufacture new ones that are of considerable benefit to human existence. Modern chemists have refined the "rules" (physical laws) of science, using theoretical models to enable reliable predictions about the properties and behaviors of materials. These successful models rely on a consideration of how minute particles of matter interact with each other through the application of only a few types of force and using fundamental properties of both waves and particles. Mathematics, with the aid of fast computers, can turn these particulate models into realistic representations of matter and its behaviors on the large scale at which people observe them.

In this resource guide, you have studied how such models can explain or predict the way that matter is shaped and held together by bonding and/or intermolecular forces. You have learned how to apply the models to rationalize the varied ways in which matter can undergo both spontaneous and non-spontaneous change, and you have learned what determines the rates at which such changes happen. You have also learned how different forms of energy, including enthalpy and entropy, are involved in determining the structures of matter and should now better understand the laws governing changes among those structures. •







temperature at which the volume of an ideal gas is zero and all molecular motion stops

**Absorption or emission spectrum** – when radiant energy is taken in or released by absorbing or giving off a photon, the signature pattern of absorption or emission can be used in identifying an element

**Acids** – a class of compounds that, when in water solutions, produce an excess of hydrogen ion (H⁺); such solutions will test below 7 on the pH scale; originally the name acid meant "sour taste"

**Activation energy** – in a collision of reactants, the amount of energy required in order for a reaction to form transition states and proceed to products

**Activity series** – an arrangement of metals sequenced in decreasing order of their ability to act as a reducing agent in aqueous solutions

**Alchemy** – a practice that was particularly popular during the middle ages that was focused on trying to change other metals into gold or silver

**Allotrope** – one of several possible forms of an element; for instance, carbon can be graphite, diamond, or buckminsterfullerene (commonly called "buckyballs")

**Alloy** – a solid or liquid mixture of two or more metals; some common alloys include steel, brass, and bronze

**Alpha decay** – a type of radioactive decay where the nucleus emits two neutrons and two protons which is the equivalent of a helium nucleus,  ${}_{2}^{4}\mathrm{He}^{2+}$ 

**Alpha particle** – the helium nucleus, ⁴₂He²⁺, emitted in an alpha decay

**Amphoteric** – describes a substance that can behave as either an acid or a base

**Anion** – any negatively charged ion

**Anode** – the positive electrode of an electrolytic cell and the site of oxidation

Aqueous solution - a solution made by dissolving a substance in the solvent water

Arrhenius, Svante – 1903 Swedish Nobel Laureate chemist known for his theory on acids and bases

**Atom** – the smallest whole unit of an element that contains a nucleus and electrons around the nucleus **Atomic mass** – the average mass of an atom of an element; it is the weighted average of the isotopes based on the proportion of each isotope in a given element sample

Atomic mass unit - abbreviated amu and also called a Dalton, it is the mass equal to 1/12 the mass of a carbon-12 atom

**Atomic number** – the number of protons in the nucleus of an atom; this identifies the type of atom; for example, all atoms with seven protons are nitrogen

**Atomic radius** – the radius of an atom from the center of its nucleus to the edge of its outermost electron orbital

**Avogadro's Law** – this gas law states that when two gases at the same temperature and pressure have the same volume, the number of molecules of each gas is also equal

**Avogadro's number** – a number representing the number of atoms in 12.0 g of C; the number is  $6.022 \times 10^{23}$ 

**Bases** – a class of compounds that, when in aqueous solutions, produce an excess of OH⁻ ions; such solutions will test above 7 on the pH scale

Beta decay - radioactive atom decay that emits a charged particle; if this particle is negative, then it is an electron; if it is positive, then it is called a positron

**Beta particle** – a negatively charged or positively charged particle with the mass of an electron emitted in beta decay

**Binary** – describes a substance made of two elements

**Body-centered cubic** – a unit cell of a crystal where there is a particle at each corner of a cube and one in the center

Bohr model of the atom - a model of the atom where electrons orbit the nucleus like planets around the sun, but at predetermined distances; electrons can transition from one orbit to a different empty one

**Boiling point** – the temperature at which the vapor pressure of a liquid substance equals the atmospheric pressure; normal boiling points are defined at I atmosphere of pressure

**Boltzmann constant** – the value of the constant in an equation connecting the entropy (S) of a system to the number of possible ways in which that substance can be recognized; it is the k in the equation  $S = k \ln W$ 

**Bonds**– (*referring to chemical bonds*) – attractive forces between a pair of atoms that provide a link between them so that energy is required to break them apart

**Boyle's Law** – a gas law stating that the volume of a gas is inversely proportional to its pressure when the number of gas molecules and the temperature remain constant;  $P_1V_1 = P_2V_2$ 

**Brass** – an alloy of primarily copper and zinc with varying proportions of each metal

**Brønsted-Lowry acid-base theory** – a model of acid-base theory based on the transfer of a proton; according to this theory, an acid is a proton donor and a base is a proton acceptor

**Bronze** – an alloy of primarily copper and tin with varying proportions of each metal

**Buret** – a liquid measuring device used in titrations; the liquid is dispensed through the bottom of the device by the use of a stopcock that open and closes to control the flow of liquid

**Calorimetry** – the measurement of heat evolution or absorption in chemical reactions

**Catalyst** – a substance that increases the rate of a chemical reaction by lowering the activation energy needed for the reaction by altering the reaction pathway

**Cathode** – the negative electrode of an electrolytic cell and the site of reduction

Cation – any positively charged ion

**Cell potential** – electromotive force (emf) or cell voltage; the electrical potential difference between the two electrodes in an electrolytic cell

**Charles' Law** – a gas law stating that the volume of a gas and its temperature are directly proportional when the number of molecules and the pressure are constant

**Chemical bond** – when an attractive force between two atoms exists such that the combination of atoms behaves as a unit that requires force (energy) to break apart

**Chemical formula** – a written expression using symbols and subscripts to indicate the number and type of each atom in a chemical unit

**Chemical kinetics** - the study of the rate of change in chemical reactions

**Chromeplating** – the technique of electroplating a thin layer of chrome onto an object

**Colligative properties** – the physical properties of a solution that depend only on the concentration of solute species and not on what they are

**Collision model** -- a model used to explain reaction rates using the number of molecules colliding per second with a certain required minimum energy

**Combustion** – a reaction involving oxygen as a reactant

**Concentration** – the measurement of how much solute is in a given amount of solution; depending on the type of problem, different types of concentration units can be used, such as molarity and molality

**Covalent bond** – the sharing of a pair of electrons by two atoms in a chemical bond

**Covalent network** – an extended lattice formed of covalent bonds, such as in SiO₂ or diamond

**Critical point** – the endpoint of the liquid-gas curve on a phase diagram; it is the point at which a vapor and liquid become indistinguishable

**Crystalline** – a solid with a well-defined lattice shape due to an orderly arrangement of its particles within the solid

**Cubic Closest Packed** – an arrangement in a crystal structure formed by alternating face-centered cubic layers: this arrangement provides the maximum density of particles, hence the name "closest" packing

**Decomposition** – a type of chemical reaction where a substance breaks down into simpler substances; such reactions have one reactant and two or more products

**Diffusion** – the movement of one fluid through another from an area of high concentration to an area of lower concentration

**Dipole** – formed when the electrons in a molecule are arranged so that one part of the molecule has a positive charge and another part has a negative charge

**Dipole moment** – the product of one of the equal but opposite charges on two atoms in a molecule and the distance separating them

**Diprotic** – describes an acid capable of donating 2  $H^+$  ions per molecule; for example, sulfuric acid,  $H_2SO_4$ 

**Dispersion force** – the intermolecular attraction between any two particles resulting from the polarization of the electrons in the electron clouds: also called *London dispersion forces* 

**Double replacement** – the reaction of two ionic compounds to form two new ionic compounds

**Ductile** – the property of a metal that describes how easily the metal can be pulled into a wire without breaking

**Effusion** – the process by which a gas escapes from its container through a tiny hole

**Electrochemistry** – the division of chemistry that studies the relationship between reactions and electrical parameters such as voltage and current

**Electrolysis** – the nonspontaneous splitting of a substance by adding electrical energy; an example is the splitting of water into hydrogen and oxygen by applying a voltage and current

**Electron** – a fundamental particle of matter that has a negative charge; discovered by J. J. Thompson in 1896

**Electron affinity** – the energy change, in kJ, when one mole of atoms of an element adds a mole of electrons (one per atom)

**Electron cloud** – the representation of the areas occupied by electrons in the atom; the types of electron clouds include ones with different shapes, called "orbitals"

**Electronegativity** – the ability of an atom to attract shared electrons in relative units

**Electroplating** – the deposition of a thin coat of a metal onto the surface of an object by passing an electrical current through a salt solution containing the metal ion to be deposited

**Electrostatic force** – a force created by the attraction of oppositely charged ions or charged particles

**Element** – the simplest type of substance with unique physical and chemical properties

**Empirical formula** – a chemical formula reduced to lowest terms; for example, C₂H₂ reduced to CH

**Endothermic** – describes a chemical reaction that absorbs heat from the surroundings

**Endpoint** – see *equivalence point* 

**Enthalpy** – a thermodynamic quantity (*II*) that measures the heat content under constant pressure

**Entropy** – a thermodynamic quantity (S) that is a measure of the state of order in a substance

**Equilibrium** – the condition of a chemical reaction when the forward rate and the backward rate are equal, so there is no apparent change

**Equilibrium constant** – the number obtained if concentrations and/or pressures of the products are divided by those of the reactants when the reaction is at equilibrium

**Equivalence point** – the point in a titration when the moles of the two reactants are equal; in an acid-base titration, the moles of H⁺ transferred equals the moles of OH⁻; also called the *endpoint* 

**Exothermic** – describes a reaction that releases heat to the surroundings

**Force** – mass x acceleration or time derivative of momentum; force creates change in a physical system

**Galvanic electrochemical cell** – a spontaneous reaction used to generate electrical potential (voltage) and deliver energy

**Gas** – a state of matter characterized by molecules being spread out from each other and therefore highly compressible; takes the shape of any container it is in, creating uniform pressure in all directions

**Gibbs Free Energy** – a thermodynamic quantity that indicates whether a reaction will be spontaneous and how much work it can do; a spontaneous reaction has a release of free energy ( $\Delta G$ ) found from the equation  $\Delta G = \Delta H - T\Delta S$ , where  $\Delta H$  is the enthalpy change, and  $\Delta S$  is the entropy change

**Glass** – an amorphous mixture of primarily sand, soda ash, and lime, melted together

**Group** – a column on the periodic table; for example, alkali metals (group 1) or noble gases (group 18)

**Heat** – an amount of energy transferred from a hot system to a cooler system resulting in changes in molecular motion; measured in kilojoules

**Heat of formation** – the enthalpy change that occurs when 1 mole of a compound is produced from its elements

**Hexagonal Closest Packing** – a crystal structure based on a hexagonal unit cell layering

**Homogenous** – describes a solution that features a uniform dispersion of its components

**Hybridization** – the process when electron orbitals like s and p mix together to form new orbitals such as sp; the total number of orbitals is conserved, so 1s + 1p electron yields 2sp hybrids

**Hydration** – the process of being surrounded by water solvent molecules; can also form solid crystals called hydrates

**Hydrogen bond** – a type of dipole-dipole force between molecules involving hydrogen attractions to a strongly electronegative atom, either a nitrogen, oxygen, or fluorine atom in another molecule; for example, the bond that occurs when the hydrogen in a water molecule is attracted to the nitrogen atom in an ammonia molecule

**Hydronium ion** – a hydrogen ion (H⁺) bonded to a water molecule in hydration to H₃O⁺

**Ideal gas equation** – an equation that relates temperature, pressure, volume, and moles of gas together (PV=nRT)

**Indicator** – a species that has a different color at different pH values, typically one color in acid and a different color in base; this color change can indicate the endpoint of a reaction

**Induced dipoles** – formed when a previously symmetrical arrangement of charge in an atom or molecule is distorted by the presence of a nearby unsymmetrical charged species

**Insoluble** – describes a substance that does not dissolve in a solvent such as water

**Instantaneous rate** – the rate of a reaction at any one point in time

**Insulator** – any substance with a low thermal conductivity and/or low electrical conductivity

**Intermolecular force** – the type of force that attracts or repels one molecule to another molecule

**Intramolecular force** – the force in a bond that holds the atoms within a molecule together

**Ion** – a charged particle formed when an atom either gives up one or more electrons or gains one or more electrons





**Ionic bonds** - bonding that occurs between two oppositely charged ions

**Ionic lattice** – an arrangement formed of alternating cations and anions in an ionic crystal

**Ionization energy** – the amount of energy needed to remove a mole of electrons from a mole of atoms (one electron per atom) in the gaseous state; measured in kilojoules per mole

**Isotope** - an alternate form of the same element that has a different mass due to its having a different number of neutrons

**Kelvin** – a scale that is the SI (international system of units) unit of temperature; the lowest temperature is zero K; water boils at 373K

**Kinetic energy** – the energy of any particle of matter due to its motion

**Kinetic Molecular Theory** – a model that explains how the random motion of particles gives gases volume and pressure: it also explains rates of diffusion and effusion of gases

**Kinetics** – a branch of chemistry that studies the rates of reactions

**Law of conservation of mass** – the concept that during normal (non-nuclear) reactions, the mass of the reactants will equal the mass of the products: in nuclear reactions mass-energy is conserved

**Lewis structures** – a structural formula representation used to depict the arrangement of atoms, electrons, and bonds in a molecule using dots to represent electrons and a line to represent the sharing of an electron pair

**Limiting reactant (limiting reagent)** – the reactant in a reaction that determines the maximum amount of product that can be made

**London dispersion force** - see dispersion force

**Lustrous** – refers to the appearance of a substance's surface in reflected light; metals have luster

**Malleable** – describes the property of a substance indicating it can easily be flattened into thin sheets

**Mass number** – the total number of protons and neutrons in the nucleus of an atom

Mass spectrometer — an instrument used to measure the relative masses of ions; this instrument makes the ions from neutral molecules and subsequently measures their mass; the mass and structure of the original neutral molecules can be deduced from the measurements of the ions formed

**Melting point** – the temperature at which a substance transitions from a solid to a liquid; usually measured at 1 atmosphere pressure

**Mesopotamia** – the area located between the Tigris and Euphrates Rivers, which is now Iraq and parts of Syria, Turkey, and Iran

**Metallic bond** – a type of bond that forms between metal atoms in a solid metal, where the atoms act as if positive ions form and electrons are mobile in a "sea" and so are able to conduct an electric charge

**Mixture** – two or more substances that are intermingled with each other but not chemically bonded together

**Molality** – a measure of concentration in moles of solute per kilogram of solvent

**Molarity** – a measure of concentration in moles of solute per liter of solution

**Molar mass**— the relative mass of a molecule (or an atom) compared to the mass of a single atom of ¹²C set as a standard at 12.0000

**Mole** – the amount of a substance containing  $6.022 \times 10^{23}$  particles of the substance

**Mole fraction** – a measure of concentration based on the number of moles of one substance compared to the total number of moles of all substances present

**Molecular solid** – a solid held together by weak intermolecular forces between the molecules; typically molecular solids have relatively low melting points and boiling points

**Molecule** – a compound consisting of two or more atoms held together by a chemical bond

**Neolithic age** – the last part of the Stone Age beginning around 10,000 B.C.

**Net ionic equation** – a type of chemical equation where spectator ions have been eliminated to highlight the actual chemical reaction taking place

**Neutron** – a fundamental particle of matter found in the nucleus that has a mass of 1.009 amu but no electric charge; first identified by Sir James Chadwick in 1932

**Nonpolar molecules** – a molecule without a net permanent dipole

**Nonspontaneous** — describes a reaction that does not naturally occur in the direction indicated (from left to right)

**Nucleus** – the very dense positively charged center part of the atom that contains the protons and neutrons and, consequently, most of the mass of the atom

**Nuclide** – a nuclear species of an atom characterized by mass (protons + neutrons), charge (protons), and energy content

 $\begin{array}{l} \textbf{Octane} = \text{a hydrocarbon with eight carbons in its} \\ \text{chain, all of which are single bonded together and fully} \\ \text{surrounded by hydrogen atoms; the chemical formula for} \\ \text{octane is } C_8\Pi_{18} \\ \end{array}$ 

**Orbital** – in an atom, the area in space around the nucleus where an electron is found; each orbital holds only one or two electrons

**Organic** – describes compounds in which carbon is bonded, often to itself, in chains or rings; organic compounds usually contain many hydrogen atoms

**Oxidation number** – the number of electrons that must be added to or subtracted from an atom of an element in a compound to convert it to its elemental state of zero charge

**Oxidation-reduction** – a type of reaction where one reactant loses one or more electrons while another reactant gains those electrons; often abbreviated "redox"

**Oxidation state** – essentially equivalent to oxidation number (see above); the number that indicates the *oxidation state* is equal to the effective charge on an atom

**Period** – a row on the periodic table; period 1 has only H and He

**Periodic table** – a table showing all the elements, symbols, and masses arranged into seven rows and eighteen columns according to increasing proton number and similar electron configurations

**pH** – a number scale that indicates whether a solution is acidic or basic; in aqueous solutions, a pH below seven indicates an acid while a pH above seven indicates a base

**pH meter** – an instrument used to indicate the pH of a solution

**Phase** – a state of matter such as a solid, liquid, or gas

**Phase diagram** – a diagram that shows the temperature and pressure conditions of a pure substance as it changes from one state of matter to another

**Photoelectric effect** – the emission of an electron from a metal's surface as a result of the absorption of a photon (electromagnetic radiation); this phenomenon was first explained by Albert Einstein; the amount of energy varies depending on the metal

**Photon** – a discrete unit or quantum of electromagnetic radiation; forms include microwaves, light waves, X rays, and gamma rays; photons are emitted when electrons move from one energy state to another in an atom, for example when an excited electron drops from the sixth energy level to the second energy level; the energy released to make the change is the energy of the emitted photon

**Polar molecule** – a molecule with an unequal overall distribution of permanent positive and negative charge; such a molecule has a net dipole moment that is not equal to zero

**Positron decay** – a radioactive decay of a nucleus emitting a particle whose mass and spin are the same as the electron but has a positive charge

**Potential energy diagram** – a diagram that shows the energy of the reactants, the activation energy, and the energy of the products in relationship to one other; this type of diagram can show if a reaction is endothermic or exothermic and how fast the reaction will occur

**Precipitate** – an insoluble solid substance formed by chemical reaction within a liquid solution

**Precipitation reaction** – a reaction of two soluble ionic solutions that, when combined, form an insoluble product called a precipitate

**Product** – a substance that is formed in a chemical reaction

**Proton** – a subatomic particle found in the nucleus that has a positive charge and a mass of 1.008 amu

**Quantum** – the smallest possible unit of energy; this unit can apply to any form of energy

**Quantum Mechanical Model** – a model of the atom that explains the presence of electrons as "stationary" waves in orbitals

Radioactive atom – an atom whose nucleus will break up to form a more stable arrangement of the nuclear particles; was first discovered by Antoine Henri Becquerel in 1896; radioactivity is not affected by the chemical state of the atom, and thus processes like burning do not destroy an atom's radioactivity

Radioactive decay – the process through which an atomic species transmutates into a new species; for instance, U-238 decays initially to thorium-234 and then through a series of steps into Pb-206

**Rate law** – an equation that shows the factors that affect the rate of the reaction and how they affect it

Reactants - the starting substances in a chemical reaction

**Redox** – see oxidation-reduction

**Rotational motion** – the movement of a molecule circling around its central atom, or an individual group within a molecule rotating about a bond

**Salt** – an ionic compound resulting from the combination of an acid and base; sodium chloride and lithium fluoride are both salts; sodium chloride is often referred to as "common salt" or just "salt"

**Saturated solution** – a solution containing the maximum amount of solute that can be dissolved at a given temperature

**Single replacement** – a type of reaction where an element combines with a compound and the element displaces a species in the compound that then becomes an element

**Solubility** – the maximum amount of a solute that can dissolve in a solvent at a particular temperature; often expressed in grams per liter

**Soluble** – describes the ability of a solute to dissolve in a solvent at a particular temperature

**Solute** – the substance being dissolved in a solvent to form a solution

**Solution** – a homogeneous (uniformly dispersed) mixture

**Solvent** – the substance doing the dissolving in a solution; the solvent generally forms the main component of a solution

Sonorous - producing sound

**Spectator ions** – ions that are not involved in a reaction but are in the solution

**Spontaneous** – a reaction that occurs naturally

**Stalactites** – formations in a cave that descend from the ceiling and are made of calcium carbonate and other minerals

**Stalagmites** – formations in a cave that grow upward from the floor due to dripping minerals, including calcium carbonate

State function – a property of a system that is determined by its current state and not how it got there; volume,  $V_i$  is an example of a state function; since chemistry mostly involves studying change, it often focuses on change in a state function; this also does not depend on the pathway; for instance, only  $V_{\text{initial}}$  and  $V_{\text{final}}$  are important for determining  $\Delta V$ 

**Steel** - a metal alloy made with primarily iron and carbon whose composition varies; other components can include nickel, chromium, cobalt, molybedum, or zirconium

**Stoichiometry** – the establishment and use of mass-mole relationships in balanced chemical equations

**STP** – standard temperature and pressure for a gas which is at 0°C (273K) and 1 atm

**Supercritical fluid** – a dense gas that is above its critical temperature and pressure and behaves as a viscous fluid

**Synthesis** – a type of reaction where two or more reactants make a new product

**Temperature** – a measure of how hot or cold a substance is relative to another substance; actually a measure of the thermal (motion) energy content of a system

**Temporary dipoles** — a momentary induced dipole created when a polar molecule distorts the electron cloud of a nonpolar molecule

**Thermochemistry** – a branch of chemistry focusing on the heat transferred during chemical reactions

**Thermodynamics** – the study of the mathematical relationships among work, temperature, and different forms of energy in chemical systems, especially chemical reactions

**Titration** – a method used in chemistry to determine an unknown concentration of a substance, by comparing it with another substance; typically used with acid and base solutions

**Toxic** – describes substances that are able to cause damage to living tissue, especially upon ingestion

**Translational motion** – the net movement of an object through space (in a solid, liquid, or gas), excluding vibration or rotation

**Transmutation** – the natural or artificial process of changing an atom from one type of atom to another as the result of a nuclear reaction

**Triple point** – the one temperature and pressure at which all three phases of a substance (solid, liquid, and gas) exist at the same time in equilibrium with each other

**Triprotic** describes an acid that has the ability to donate 3 protons ( $H^+$  ions); for example,  $H_3PO_4$ 

**Valence electrons** – the outermost electrons in an atom that are used in bonding

Valence Shell Electron Pair Repulsion Model (VSEPR) – a model that is used to explain or predict the shape of many molecules and ions by minimizing electron pair repulsions around the central atom

van der Waals force – weak forces of attraction between atoms or molecules that result from permanent or temporary electrostatic forces

**Vibrational motion** – the motion of two atoms in a molecule that results from the stretching or bending of a bond in the molecule

**Voltage** – (or cell voltage) the electrical potential energy difference between two electrodes when no current is flowing

**Volts** – the SI unit of electrical potential

**Volume** – the space occupied completely by a sample of matter

**Work** – the energy used when a force moves an object a certain distance; in chemistry, a gas expanding and moving a piston is an example of work; work is measured in joules or kilojoules

**Yield** – the amount of product made in a reaction

Period

Actimum

Ac

12271

Actinides ~

Thorium

Th

232.04

rotactinium

Pa

231.04

Franion

U

. 238.03

Septuniun

Np

[237]

Platonium

Pu

[244]

Americium

Am

[243]

Curium

Cm

[247]

Berkelium

Bk

[247]

Litifornium

Cf

[251]

insteimun

Es

[252]

Fermium

Fm

[257]

lendeleviur

Md

[258]

Nobelium

No

[259]

awrencium

Lr

[262]





- 1. Ortus medicinae, published 1648, English translation, Oriatrike or Physick Refined, published 1662. Quoted in Partington, James R. A Short History of Chemistry (Mincola: Dover, 1989).
- 2. "The Leuven Local Heroes in Thermal Sciences and Engineering," Katholieke <u>Universiteit Leuven</u> (1999) 13 Oct. 2008 <a href="http://people.mcch.kuleuven.be/~erik/local_heroes.html">http://people.mcch.kuleuven.be/~erik/local_heroes.html</a>; and "Inspired by the Art of Fire," <u>Chemistry Chronicles, American Chemistry Society</u>, (Aug. 2004) 13 Oct. 2008 <a href="http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%22Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%22Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%22Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%22Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%22Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%20Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%20Jan%20Baptista%20Van%20Helmont%20%22>">http://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%20Jan%20Baptista%20Van%20Helmont%20%22>">https://pubs.acs.org/subscribe/journals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%20Jan%20Baptista%20Van%20Helmont%20%22>">https://pubscribe/gournals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%20Jan%20Baptista%20Van%20Helmont%20%22>">https://pubscribe/gournals/teaw/13/i08/pdf/804chronicles.pdf?sessid=1314/search=%20Jan%20Baptista%20Van%20Helmont%20%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Van%20Aaptista%20Aaptista%20Aaptista%20Van%20Aaptista%20Aaptista%20Aa
- 3. Charles C. Gillispie, "Science in the French Revolution," <u>Proceedings of the National Academy of Sciences</u>, May 1959; 45(5): 677–684; available at <<u>http://www.pubmedcentral.nih.gov/articlerender.fcgi?artid=222616</u>>, accessed 13 Oct. 2008.
- 4. "Radioactivity: Historical Figures," <u>Access Excellence Classic Collection</u>, 15 Aug. 2008 < <a href="http://www.accessexcellence.org/AE/AEC/CC/historical_background.php">http://www.accessexcellence.org/AE/AEC/CC/historical_background.php</a>>.
- 5. "Metal Structures," <u>Dr. Chung Chieh's Applied Chemistry Course Website, University of Waterloo</u>, 13 Oct. 2008 <a href="http://www.science.uwaterloo.ca/~cchieh/cact/applychem/metals.html">http://www.science.uwaterloo.ca/~cchieh/cact/applychem/metals.html</a>>.
- 6. "3-D Crystal Lattice Images." <u>Building a Physics Education Network by Using the World Wide Web</u>, 13 Oct. 2008 <a href="http://members.kr.inter.net/joo/physics/curri-sub/crystal/lattice.html">http://members.kr.inter.net/joo/physics/curri-sub/crystal/lattice.html</a>>.
- 7. The rules cited here are from <<u>http://answers.yahoo.com/question/imdex?qid=20080223081428AAhLKRG</u>> and <<u>http://Files.chem.vt.edu/RVGS/ACT/notes/solubility_rules.html</u>>.





- Brown, Theodore L., Eugene LeMay, Bruce E. Burston, Catherine Murphy, and Patrick Woodward. <u>Chemistry: The Central Science</u>. 11th Edition. Pearson Prentice Hall, 2008.
- Chang, Raymond. Chemistry. 7th Edition. Boston: McGraw-Hill, 2002.
- "3-D Crystal Lattice Images." <u>Building a Physics Education Network by Using the World Wide Web</u>. 13 Oct. 2008 <a href="http://members.kr.inter.net/joo/physics/curri-sub/crystal/lattice.html">http://members.kr.inter.net/joo/physics/curri-sub/crystal/lattice.html</a>.
- Gillispic, Charles C. "Science in the French Revolution." <u>Proceedings of the National Academy of Sciences</u>. May 1959; 45(5): 677–684; available at <a href="http://www.pubmedcentral.nih.gov/articlerender.fcgi?artid=222616">http://www.pubmedcentral.nih.gov/articlerender.fcgi?artid=222616</a>); accessed 13 Oct. 2008.
- "Inspired by the Art of Fire." <u>Chemistry Chronicles, American Chemistry Society</u> (Aug. 2004) 13 Oct. 2008 < <a href="http://pubs.acs.org/subscribe/journals/tcaw/13/i08/pdf/804chronicles.pdf?sessid=1314#search=%22Jan%20Baptista%20Van%20Helmont%20%22">http://pubs.acs.org/subscribe/journals/tcaw/13/i08/pdf/804chronicles.pdf?sessid=1314#search=%22Jan%20Baptista%20Van%20Helmont%20%22>.
- "The Leuven Local Heroes in Thermal Sciences and Engineering." <u>Katholieke Universiteit Leuven</u> (1999) 13 Oct. 2008 <a href="http://people.mech.kuleuven.be/~erik/local_heroes.html">http://people.mech.kuleuven.be/~erik/local_heroes.html</a>.
- "Metal Structures." <u>Dr. Chung Chieh's Applied Chemistry Course Website, University of Waterloo</u>. 13 Oct. 2008 < <a href="http://www.science.uwaterloo.ca/~echieh/cact/applychem/metals.html">http://www.science.uwaterloo.ca/~echieh/cact/applychem/metals.html</a>>.
- Partington, James R. A Short History of Chemistry. Mineola: Dover, 1989.
- "Radioactivity: Historical Figures." <u>Access Excellence Classic Collection</u>. 15 Aug. 2008 < <a href="http://www.accessexcellence.org/AE/AEC/CC/historical_background.php">http://www.accessexcellence.org/AE/AEC/CC/historical_background.php</a>>.
- Silberberg, Martin. Chemistry: The Molecular Nature of Matter and Change. 4th Edition. Boston: McGraw-Hill, 2006.

