TOPIC 13

Matter and Atomic Structure

Everything in the universe—large and small, living and nonliving—is composed of matter. Matter is defined as anything that has mass and volume. In this topic, you will learn about the basic components of matter and how those components combine to create distinct forms of matter with specific properties. You will also learn about the different states of matter and what happens when matter changes from one state to another.

SUBTOPIC A THE COMPONENTS OF MATTER

Covers National Science Content Standards UCP.1, UCP.2, UCP.3, UCP.5; B.1, B.2, B.3, B.6

Unifying Concepts and Processes

UCP.1	Systems, order, and organization
UCP.2	Evidence, models, and explanation
UCP.3	Change, constancy, and measurement
UCP.5	Form and function

Physical Science

B.1	Structure of atoms
B.2	Structure and properties of matter
B.3	Chemical reactions
B.6	Interactions of energy and matter

VOCABULARY

element	ground state
atom	excited state
proton	valence electron
neutron	noble gas
electron	isotope
nucleus	mass number
atomic number	atomic mass
energy level	radioactive decay

Any sample of matter that is made of only one kind of component is referred to as a pure substance. A pure substance has the same composition and properties throughout a sample and from sample to sample.

Elements

The simplest pure substances are elements. An **element** is a pure substance that cannot be broken down into simpler components by chemical means. There are 92 naturally occurring elements in the universe. Examples of naturally occurring elements include hydrogen, silicon, tungsten, and gold. Another 23 elements have been produced artificially in laboratory experiments.

Each element is represented by a symbol that consists of one, two, or three letters. For example, the symbols for hydrogen and silicon are H and Si, respectively. The symbols for some elements are derived from the names for those elements in other languages. For example, the symbol for tungsten (W), comes from the German word for tungsten, *wolfram*, and the symbol for gold (Au) is based on the Latin name for gold, *aurum*.

By far the most abundant element in the universe is hydrogen, as you can see in Figure 13-1A, shown on the next page. The second-most abundant element is helium. Together, hydrogen and helium make up more than 99 percent of the matter in the universe. In contrast, the two most abundant elements in Earth's crust are oxygen and silicon, as Figure 13-1B shows. These two elements account for nearly 75 percent of the mass of the crust. Most of the rest of the crust is composed of metals, such as aluminum, iron, and calcium.

Atoms

All elements are made up of atoms. An **atom** is the smallest particle of an element that has all of the characteristics of that element. Each atom, in turn, is constructed from even smaller particles, which are of three types: protons, neutrons, and electrons. A **proton** has a positive electrical charge and a mass of 1 atomic mass unit (1 amu, or 1.7×10^{-24} g). A **neutron** has no electrical charge and a mass very close to 1 amu. An **electron** has a negative electrical charge that

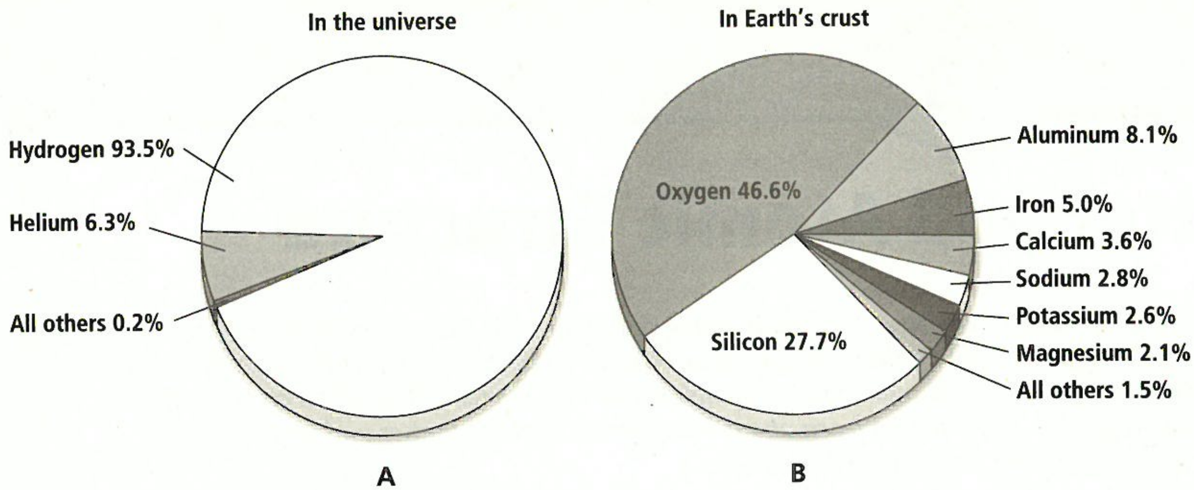


Figure 13-1 Hydrogen and helium are the most abundant elements in the universe (A). Oxygen and silicon are the most abundant elements in Earth's crust (B).

is exactly the same magnitude as the positive charge of a proton. Electrons are much smaller, though; the mass of an electron is only 0.0005 amu.

Figure 13-2 shows that protons and neutrons are found in the central part of an atom, called the **nucleus** (*plural,* nuclei). The nucleus actually comprises a very small fraction—about 0.01 percent—of the total volume of an atom. Most of an atom's volume is occupied by electrons, which are distributed around the nucleus in specific regions known as orbitals. Therefore, the size of an atom depends on the number and distribution of its electrons. Because the nucleus contains positively charged protons and uncharged neutrons, the nucleus itself has a positive charge. However, each atom has an equal number of protons and negatively charged electrons, so an atom as a whole has no net charge.

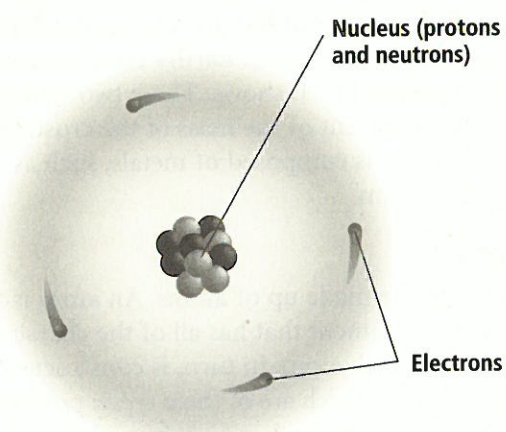


Figure 13-2 An atom is composed of protons and neutrons, which are found in the nucleus, and electrons, which are distributed around the nucleus. Note that the size of the nucleus is greatly exaggerated in this drawing. The nucleus actually comprises a tiny fraction of an atom's total volume.

Different elements contain different numbers of protons in the nuclei of their atoms. For example, a hydrogen atom has only one proton in its nucleus, a nitrogen atom has seven, and a uranium atom has 92. Each element is given an **atomic number** that equals the number of protons in the nucleus of an atom of that element. Thus, the atomic numbers of hydrogen, nitrogen, and uranium are one, seven, and 92, respectively.

Electron Energy Levels

The electrons in an atom are grouped into one or more energy levels. Each **energy level** represents a specific amount of energy that an electron can have. (Don't confuse energy levels, which determine the amount of energy in each electron, with orbitals, which are related to electron location.) The first energy level, which corresponds to the lowest energy, can hold only two electrons. Higher energy levels can hold progressively more electrons. The second energy level, for example, can hold eight electrons.

Electrons tend to occupy the lowest available energy levels. As a result, lower energy levels generally must be filled before electrons can move into higher levels. Examine Figure 13-3, which diagrams the electron energy levels for aluminum. The atomic number of aluminum is 13, so each aluminum atom has 13 protons and 13 electrons. Two of the 13 electrons are in the first energy level and eight are in the second. Hence, the first and second energy levels are filled. Aluminum's three remaining electrons are in the third energy level, which is not filled.

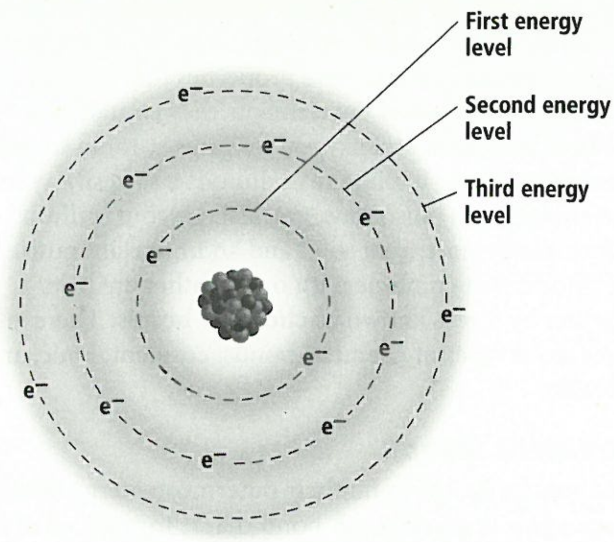


Figure 13-3 An alumnium atom has 13 electrons: two in the first energy level, eight in the second energy level, and three in the third energy level.

When an atom has all of its electrons in the lowest available energy levels, the atom is said to be in its **ground state.** If an atom absorbs energy, however, the added energy may boost one or more electrons into a higher energy level. This condition is referred to as an **excited state.** The electrons that have moved to a higher energy level eventually return to their original, lower energy level. As they do, they release, or emit, energy, sometimes in the form of visible light. For example, the light from an incandescent lightbulb is produced when metal atoms in a thin wire return to their ground state after having been put in an excited state by electricity moving through the wire.

Valence Electrons

Electrons in the outermost energy level of an atom are called **valence electrons.** The number of valence electrons in an element largely determines the chemical properties of the element. For example, sodium (atomic number 11) and potassium (atomic number 19) both have one valence electron, and both are highly reactive metals, which means they combine easily with other elements. Chemical reactivity also depends somewhat on atomic size, however. Potassium, with eight more electrons than sodium, is slightly larger than sodium, so the chemical properties of these two elements are not identical.

There are six elements—helium, neon, argon, krypton, xenon, and radon—that have outermost energy levels that are full. These six elements are referred to as **noble gases.** All six have very low chemical reactivity. In fact, the first three are chemically inert, which means they rarely combine with other elements.

Isotopes

Although all atoms of an element have the same number of protons and electrons, the number of neutrons can vary. Chlorine atoms, for example, always have 17 protons and 17 electrons, but they may have either 18 or 20 neutrons. Atoms of the same element that have different numbers of neutrons are called **isotopes.**

Recall that protons and neutrons have very similar masses and that the mass of an electron is much smaller. As a result, the mass of an isotope depends almost entirely on the number of protons and neutrons in the isotope, which is defined as the **mass number.** The mass number of the chlorine isotope with 18 neutrons is 35 (17 protons + 18 neutrons). The mass number of the chlorine isotope with 20 neutrons is 37 (17 protons + 20 neutrons).

Because isotopes have different masses, they can be distinguished from one another by physical properties that depend on mass, such as density. However, isotopes have the same chemical properties because they have the same number of electrons.

Scientists use several notations to represent isotopes. Sometimes the name or chemical symbol of the element is followed by a hyphen and the mass number of the isotope, as in chlorine-35 or Cl-35. Alternatively, the chemical symbol may be preceded by the mass number written as a superscript (^{35}Cl); in a variation of this notation, the atomic number is also written before the chemical symbol as a subscript ($^{35}_{17}Cl$).

Atomic Mass

Each element has an **atomic mass,** which is the mass in atomic mass units of one atom of the element. If an element has more than one isotope, its atomic mass is the average of the isotopes' mass numbers, adjusted for the relative abundance of the isotopes. Consider the isotopes of chlorine, for example. Chlorine-35 has a mass number of 35 and accounts for 77 percent of all chlorine atoms in nature; chlorine-37 has a mass number of 37 and accounts for 23 percent of all chlorine atoms. Therefore, the atomic mass of chlorine equals $(35 \times 0.77) + (37 \times 0.23)$, or 35.45.

Radioactivity

A small fraction of the isotopes on Earth have nuclei that are unstable. Unstable nuclei emit radiation as they spontaneously change, or decay. This process is called **radioactive decay.** During radioactive decay, a nucleus can lose protons and neutrons, a proton can change into a neutron, or a neutron can change into a proton. Note that all of these changes affect the number of protons in the nucleus. Hence, radioactive decay transforms the nucleus of one element into the nucleus of another element. For example, carbon-14 decays through a process in which a neutron changes into a proton. After decay, the nucleus has seven protons instead of six, so it has become the nucleus of a nitrogen atom. Since protons and neutrons have virtually the same mass, the mass number of the new nucleus is still 14, and the isotope that is formed is nitrogen-14.

Each radioactive isotope decays at a constant rate that is specific for that isotope. Thus, by measuring the amount of carbon-14 and other radioactive isotopes in objects such as fossils and rocks, scientists can determine the age of those objects.

SUBTOPIC B
HOW ATOMS COMBINE

Covers National Science Content Standards UCP.1, UCP.2, UCP.3, UCP.5; B.1, B.2, B.3, B.5, B.6

Unifying Concepts and Processes

UCP.1	Systems, order, and organization
UCP.2	Evidence, models, and explanation
UCP.3	Change, constancy, and measurement
UCP.5	Form and function

Physical Science

B.1	Structure of atoms
B.2	Structure and properties of matter
B.3	Chemical reactions
B.5	Conservation of energy and increase in disorder
B.6	Interactions of energy and matter

VOCABULARY

compound	chemical reaction
chemical bond	solution
covalent bond	concentration
molecule	solubility
ion	acid
ionic bond	base
metallic bond	

An atom is chemically most stable when its outermost energy level is full. Only atoms of the noble gases have a full outermost energy level in the uncombined state. Atoms of other elements fill their outermost energy level by combining with other atoms to form pure substances that are more stable. A pure substance composed of the atoms of two or more elements is called a **compound.** Water, which is composed of two hydrogen atoms and one oxygen atom, is a compound. So is table salt, or sodium chloride, which is composed of one sodium atom and one chlorine atom.

Compounds usually have properties that are very different from those of the elements they contain. For example, hydrogen and oxygen are gases at room temperature, but water is a liquid. Sodium is a soft, silvery metal and chlorine is a poisonous, greenish gas, but sodium chloride is a transparent solid and an important nutrient.

Atoms that combine with one another are held together by forces known as **chemical bonds.** There are three main types of chemical bonds: covalent, ionic, and metallic.

Covalent Bonds

One way for atoms to fill their outermost energy level is to share a pair of electrons. A bond that is formed when two atoms share a pair of electrons is called a **covalent bond.** Figure 13-4 shows how hydrogen and oxygen form covalent bonds in water. A hydrogen atom has only one electron and needs another electron to fill its first energy level. An oxygen atom has two electrons in its first energy level and six valence electrons in its second energy level. Thus, oxygen needs two electrons to fill its second energy level. When two hydrogen atoms each share their electron with an oxygen atom, the oxygen atom acquires a full second energy level. The oxygen atom, in turn, shares one of its valence electrons with each hydrogen atom, giving each hydrogen atom a full first energy level. Two covalent bonds are formed, one between each hydrogen atom and the oxygen atom.

A pure substance that consists of atoms held together by covalent bonds is defined as a **molecule.** Molecules may contain as few as two atoms or more than a thousand. Most molecules are also compounds; that is, they usually contain atoms of two or more elements. Water is an example of such a molecule, called a molecular compound. However, some molecules contain atoms of only one kind of element. For example, hydrogen gas consists of molecules containing two hydrogen atoms. Because it contains only one kind of element, a hydrogen molecule is not a compound.

Molecules are represented by chemical formulas that include the symbol of each element in the molecule followed by a subscript number that indicates how many atoms of that element are in the molecule. If an element has only one atom in the molecule, no subscript follows the symbol. For example, the chemical formula for water is H_2O, indicating that there are two hydrogen atoms and one oxygen atom in each water molecule. A molecule of hydrogen contains two hydrogen atoms, so its chemical formula is H_2.

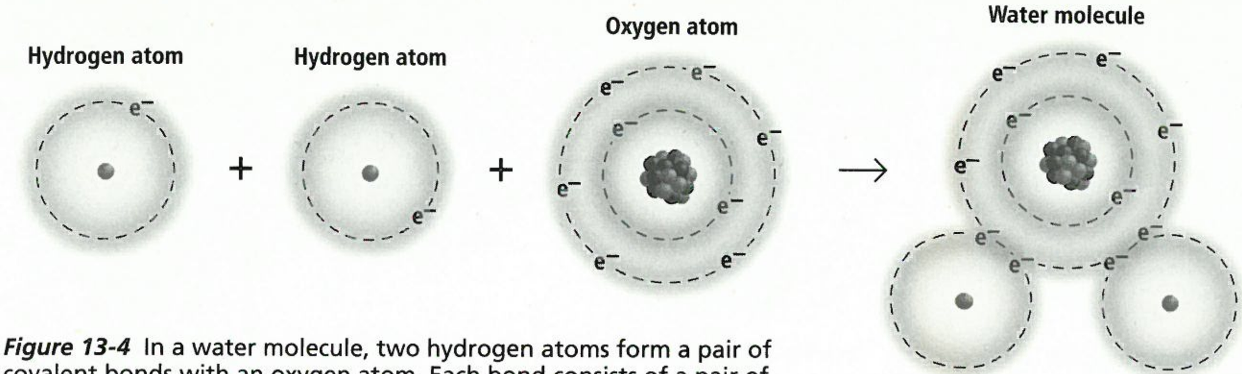


Figure 13-4 In a water molecule, two hydrogen atoms form a pair of covalent bonds with an oxygen atom. Each bond consists of a pair of valence electrons that are shared by one of the hydrogen atoms and the oxygen atom.

Polar Covalent Bonds

In all molecules, the positively charged nuclei of the atoms are attracted to the negatively charged electrons. Some nuclei have a greater attraction for electrons than do other nuclei, however. In a water molecule, for example, the oxygen nucleus attracts electrons more strongly than do the hydrogen nuclei. As a result, the two electrons that form each covalent bond are not shared equally; they spend more time near the oxygen nucleus than near the hydrogen nuclei. Covalent bonds in which electrons are not shared equally are called polar covalent bonds. As Figure 13-5 shows, this unequal sharing of electrons causes the water molecule to have a negatively charged end and two positively charged ends. The molecule as a whole has no net charge, though, because it has an equal number of protons and electrons.

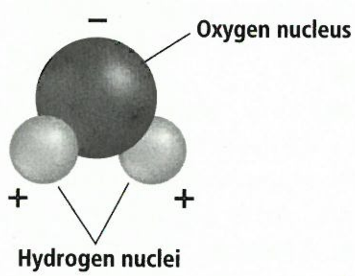


Figure 13-5 A water molecule has a negatively charged end where the oxygen nucleus is and two positively charged ends where the hydrogen nuclei are.

Ionic Bonds

Another way for atoms to fill their outermost energy level is to lose or gain one or more electrons. An atom that loses or gains electrons becomes a charged particle called an **ion.** Atoms whose outermost energy level is less than half full tend to form ions by losing electrons. Such ions have more protons than electrons and therefore are positively charged. For example, recall that a sodium atom has one valence electron; it can lose that electron, forming a sodium ion with a filled second energy level and a charge of +1. A magnesium atom can lose its two valence electrons to form a magnesium ion with a charge of +2.

Conversely, atoms whose outermost energy level is more than half full tend to form ions by gaining electrons. Ions formed in this way have more electrons than protons and are negatively charged. A chlorine atom, for example, has seven valence electrons; it can fill its outermost energy level by gaining one electron, forming a chloride ion with a charge of −1. An oxygen atom has six valence electrons and usually gains two electrons, forming an oxide ion with a charge of −2.

Atoms whose outermost energy level is exactly half full may form either positive or negative ions. Because negative ions include additional electrons, they tend to be larger than positive ions.

Ions are represented by a chemical symbol followed by a superscript number and a superscript plus or minus sign that indicate the charge of the ion. If the charge is +1 or −1, the number is omitted. Thus, the sodium, magnesium, chloride, and oxide ions are represented by Na^+, Mg^{2+}, Cl^-, and O^{2-}, respectively. Some ions consist of two or more atoms that function as a single ion. Two such ions that are important components of the materials at Earth's surface are the silicate ion (SiO_4^{4-}) and the carbonate ion (CO_3^{2-}).

Ions with opposite charge attract each other and can form stable compounds, which are known as ionic compounds. The bonds that hold ions together in an ionic compound are called **ionic bonds.** Figure 13-6 on the next page illustrates the ionic bond between an Na^+ ion and a Cl^- ion in sodium chloride, NaCl. Note that the positive ion is always written first in the chemical formula of an ionic compound. In NaCl, there are as many Na^+ ions as Cl^- ions, so the compound as a whole has no net charge. The same holds for all ionic compounds. In sodium monoxide (Na_2O), for example, there are two Na^+ ions for every O^{2-} ion, so Na_2O is electrically uncharged.

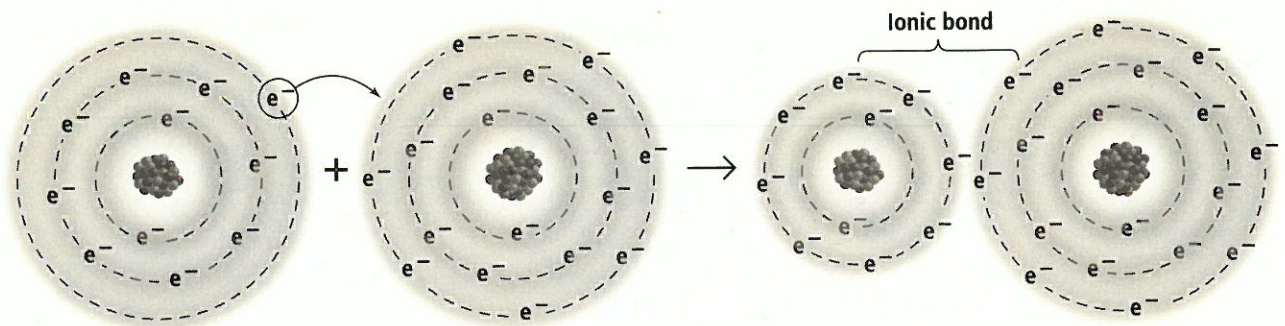


Figure 13-6 A sodium atom loses one electron to form a positively charged sodium ion. A chlorine atom gains one electron to form a negatively charged chloride ion. The oppositely charged ions are held together by an ionic bond in sodium chloride, which is an ionic compound.

Metallic Bonds

In metals, valence electrons are not shared by adjacent atoms, as in a molecule, nor are they transferred from one atom to another, as in an ionic compound. Instead, valence electrons in a metal are shared by all the atoms and are free to move throughout the metal. The positive nuclei in a metal are held together by the negative electrons between them. This type of bond is called a **metallic bond.** Because electrons can move freely in a metal, metals make good conductors of electricity, which is the flow of electrons from one place to another.

Chemical Reactions

The bonds that hold atoms together are not permanent. They can be broken as atoms recombine to form other substances. The change of one or more pure substances into other substances by means of breaking and reforming chemical bonds is known as a **chemical reaction.** Substances produced in a chemical reaction, called products, have properties that are different from those of the reacting substances, or reactants.

Scientists write equations to describe chemical reactions. For example, the chemical reaction in which hydrogen gas (H_2) and oxygen gas (O_2) combine to form water is described by the following equation.

$$2H_2 + O_2 \rightarrow 2H_2O$$

The number 2 to the left of H_2 and H_2O in the equation means that two molecules of hydrogen react with one molecule of oxygen to produce two molecules of water. Note that the equation is balanced; that is, the number of hydrogen atoms (4) and oxygen atoms (2) on the left side of the arrow equals the number of those atoms on the right side. This means that the same amount of matter is present before and after the reaction. Thus, the chemical equation reflects the fundamental law that matter can change form but cannot be created or destroyed. All chemical reactions obey this law, which is known as the law of conservation of matter.

Chemical reactions also obey the law of conservation of energy (the first law of thermodynamics), which states that energy can change form but cannot be created or destroyed. Energy is released when bonds are formed and absorbed when bonds are broken. Reactants and products rarely contain the same amount of energy, however. A reaction may produce products that contain more energy than the reactants if energy is absorbed from the surroundings. If the products contain less energy than the reactants, energy is released into the environment.

Mixtures

Molecules and ionic compounds are pure substances because they are made of only one kind of component, which can be represented by a chemical formula. Their properties are usually different from those of the elements they contain. A mixture, in contrast, is a collection of two or more pure substances whose proportions can vary. Therefore, no single formula can represent the composition of a mixture. Each pure substance in a mixture retains its original properties and can be separated from the others by physical means, such as filtering.

There are two main types of mixtures—heterogeneous and homogeneous. The substances in a heterogeneous mixture are not evenly distributed. As a result, different regions of the mixture have different properties. Soil, which may contain various-sized pieces of numerous minerals, rocks, and organic material, is an example of a heterogeneous mixture. A homogeneous mixture is one in which the substances are uniformly distributed, so all regions of the mixture have identical properties. Steel, air, and brewed coffee are examples of homogeneous mixtures.

Solutions

A **solution** is a homogeneous mixture in which the substances mix as individual particles. The simplest solutions consist of two substances. The substance that is present in the larger amount is usually called the *solvent*, while the other substance is called the *solute*. The solute is said to dissolve in the solvent. An example of a solution is salt water, which consists of NaCl (the solute) dissolved in water (the solvent).

The **concentration** of a solution is the ratio of the amount of solute to the amount of solvent or solution. Concentration may be expressed in several ways, such as percent by volume, percent by mass, and parts per million (ppm). For example, a solution that contains 10 g of NaCl in 100 g of water has a concentration of 10 percent by mass. The maximum amount of a particular solute that will dissolve in a given amount of solvent is called the **solubility** of that solute. Solubility depends on a variety of factors, including temperature, pressure, and the chemical natures of both the solute and solvent.

Most solutions are liquids, like salt water. Magma is another liquid solution; it contains ions that were liberated from the crystals of rocks by melting. Solutions can also be gaseous or solid, however. Air is a gaseous solution composed mostly of nitrogen and oxygen. At room temperature, most alloys are solid solutions of two or more metals. For example, bronze is a solution of copper and tin, and brass is a solution of copper and zinc.

Acids and Bases

Many chemical reactions involve compounds called acids and bases. An **acid** is a compound that produces hydrogen (H^+) ions when it is dissolved in water. Recall that a hydrogen atom consists of one proton and one electron. Therefore, an H^+ ion is simply a proton. A **base** is a compound that produces hydroxide (OH^-) ions when it is dissolved in water. A base can neutralize an acid by providing OH^- ions, which combine with H^+ ions from the acid to form water in the following chemical reaction.

$$H^+ + OH^- \rightarrow H_2O$$

The most common acid in our environment is carbonic acid (H_2CO_3), which is produced when carbon dioxide (CO_2) reacts with water.

$$CO_2 + H_2O \rightarrow H_2CO_3$$

Some of the carbonic acid molecules that are produced in this reaction break apart in water to form H^+ ions and hydrogen carbonate (HCO_3^-) ions.

$$H_2CO_3 \rightarrow H^+ + HCO_3^-$$

These two reactions involving carbonic acid play a major role in the decomposition of limestone and the formation of caves. However, the reaction between carbonic acid and limestone is exceedingly slow. It may take thousands of years for the carbonic acid in groundwater to break down limestone to form a cave.

The strength of acidic and basic solutions is measured by the pH scale. As Figure 13-7 shows, the pH scale ranges from 0 to 14. A solution with a pH less than 7 at 25°C is acidic; the lower the pH is, the more acidic the solution is. In contrast, a solution with a pH greater than 7 at 25°C is basic; the higher the pH is, the more basic the solution is. A pH of 7 at 25°C indicates that a solution is neutral—neither acidic nor basic. Distilled water has a pH of 7, but rainwater is usually slightly acidic, having a pH of 5.0 to 5.6. The pH values of several other materials are indicated in Figure 13-7.

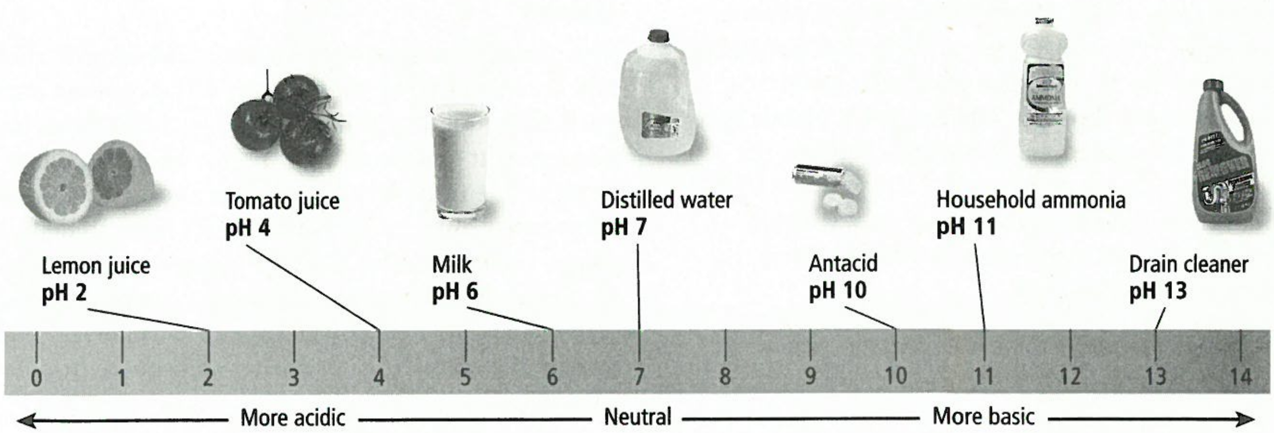


Figure 13-7 The pH scale measures the strength of acidic and basic solutions. At 25°C, a pH less than 7 indicates an acidic solution, and a pH greater than 7 indicates a basic solution. The farther from 7 the pH is, the more acidic or basic the solution is.

SUBTOPIC C
STATES OF MATTER

Covers National Science Content Standards UCP.1, UCP.2, UCP.3, UCP.5; B.2, B.4, B.5, B.6

Unifying Concepts and Processes

UCP.1	Systems, order, and organization
UCP.2	Evidence, models, and explanation
UCP.3	Change, constancy, and measurement
UCP.5	Form and function

Physical Science

B.2	Structure and properties of matter
B.4	Motions and forces
B.5	Conservation of energy and increase in disorder
B.6	Interactions of energy and matter

VOCABULARY

crystal	evaporation
glass	sublimation
plasma	condensation

Matter on Earth normally exists in one of three states: solid, liquid, or gas. Matter in a fourth state, plasma, is found less frequently on Earth but is more common elsewhere in the universe.

Solids

The solid state of matter is one in which particles—atoms, molecules, or ions—are densely packed. The particles are held tightly in position by their strong attraction for one another, which gives solids a fixed volume and a definite shape. Most solids exist as **crystals,** symmetrical structures in which the particles are arranged in regular geometric patterns, such as those shown in Figure 13-8. Crystals have flat surfaces and straight edges. The angles between the surfaces depend on the arrangement of the particles in the crystal. Sodium chloride forms cubic crystals, whereas magnesium, quartz, and many other substances form hexagonal, or six-sided, crystals.

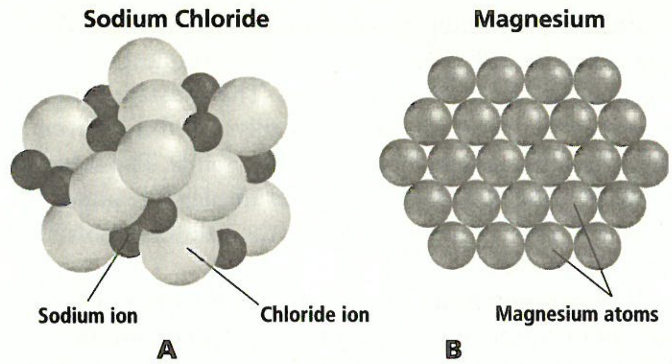


Figure 13-8 The particles in a crystal are arranged in a regular geometric pattern. The pattern is cubic in sodium chloride (A) and hexagonal in magnesium (B).

Large, well-formed crystals are rare. Instead, crystals are usually small and found in aggregations of other crystals, called polycrystalline solids. Most rocks, including granite, are polycrystalline solids.

Solids that have no crystalline structure are called **glasses.** They form when molten substances cool too rapidly for the particles to organize themselves into a regular pattern. As a result, the particles become frozen in a random arrangement. Window glass consists mostly of disordered molecules of silicon dioxide, commonly called silica. Glass also occurs in nature as the volcanic rock obsidian.

Liquids

The particles in a liquid do not attract one another as strongly as do the particles in a solid. As a result, particles in a liquid can easily slide past one another. This property allows liquids to flow. It also means that liquids do not have a definite shape; they take the shape of the container they are placed in. The distances between particles in a liquid are relatively constant, however, so liquids do have a fixed volume.

Gases

The particles in a gas have no attraction for one another and are separated by much larger distances than are the particles in solids and liquids. Gas particles move independently at extremely high speeds. They travel in a straight line until they collide with another gas particle or some other object. Like liquids, gases have no definite shape. Gases also have no fixed volume and can expand to fill any available space, unless they are restrained by the walls of a container or by a force such as gravity. Earth's gravity prevents the gases in the atmosphere from escaping into space.

Plasmas

Most of the matter in the universe is in the plasma state. **Plasmas** are gases composed of positive ions and free electrons that form when matter is heated to temperatures greater than 5000°C. Such extremely high temperatures exist in stars, and therefore the gases of stars consist entirely of plasmas. On Earth, plasmas occur naturally in the form of lightning and in parts of some flames. Plasmas can also be created by sending electricity through certain gases, as happens inside neon lights and fluorescent tubes.

Changes of State

Under the right conditions of temperature and pressure, matter can change state between solid, liquid, and gas. The particles in a solid vibrate at any temperature above absolute zero (−273°C). As the temperature rises, the particles absorb thermal energy and the intensity of their vibrations increases. At the melting point of the solid, the vibrations become vigorous enough to overcome the forces that hold the particles together, and the solid becomes a liquid.

The particles in a liquid can also absorb thermal energy. If they absorb enough energy to overcome the attractive forces in the liquid, they escape from the surface of the liquid as a gas. The process of change from a liquid to a gas is called **evaporation,** or vaporization. Particles can also enter the gaseous state from solids without entering the liquid state first. This process, known as **sublimation,** is what causes ice cubes to slowly shrink if they are left in a freezer for a long time.

A change of state from a gas to a liquid, called **condensation,** occurs when gas particles release thermal energy. A similar release of energy happens when a gas or a liquid changes into a solid.

Like chemical reactions, changes of state obey the laws of conservation of matter and conservation of energy. However, changes of state are physical changes rather than chemical changes because they are alterations in the arrangement of particles but do not involve the breaking of chemical bonds.

QUESTIONS FOR SUBTOPIC A

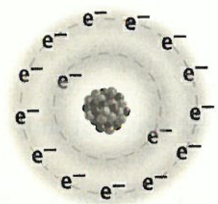
Type A

1. The most abundant element in the universe is
   - a. oxygen.
   - b. silicon.
   - c. helium.
   - d. hydrogen.
2. The major portion of an atom's mass consists of
   - a. electrons and protons.
   - b. electrons and neutrons.
   - c. neutrons and protons.
   - d. protons and orbitals.
3. Which statement is true about electrons that move between energy levels?
   - a. An electron absorbs energy when it moves to a higher energy level.
   - b. An electron releases energy when it moves to a higher energy level.
   - c. An electron absorbs energy when it moves to a lower energy level.
   - d. An electron neither absorbs nor releases energy when it changes energy levels.
4. Two elements are most likely to have similar chemical properties if they have the same
   - a. number of energy levels.
   - b. number of valence electrons.
   - c. number of neutrons.
   - d. number.
5. An atom of $^{40}_{18}Ar$ contains a total of
   - a. 18 protons.
   - b. 18 neutrons.
   - c. 40 protons.
   - d. 40 neutrons.

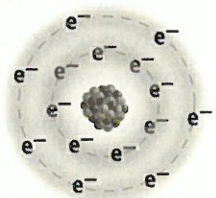
Type B

6. What is the net electrical charge on an atom with 5 protons, 6 neutrons, and 5 electrons?
   - **a.** +5
   - **b.** +1
   - **c.** 0
   - **d.** −1

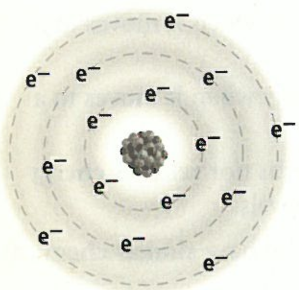
7. The atomic number of phosphorus is 15. Which diagram best represents a phosphorus atom in its ground state?



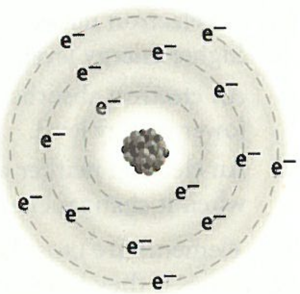
a.



c.



b.



d.

8. Which two elements would be least likely to react with each other?
   - **a.** sodium and chlorine
   - **b.** argon and aluminum
   - **c.** carbon and hydrogen
   - **d.** iron and oxygen

9. Cobalt-60 undergoes the type of radioactive decay in which a neutron changes into a proton. The atomic number of cobalt is 27. What are the atomic number and mass number of the isotope that is formed when cobalt-60 decays?

Type C

Base your answers to questions 10–12 on the partially complete table below.

Isotopes of Element Q

	Isotope #1	Isotope #2
Relative abundance	72.0%	28.0%
Mass number	85	87
Number of protons	37	
Number of neutrons		

10. How many neutrons does isotope #1 have?

11. How many protons and neutrons does isotope #2 have?

12. What is the atomic mass of element Q?

QUESTIONS FOR SUBTOPIC B

Type A

13. Which material is a pure substance?
    - **a.** soil
    - **b.** steel
    - **c.** air
    - **d.** water

14. Which substance contains atoms joined by covalent bonds?
    - **a.** aluminum
    - **b.** hydrogen
    - **c.** sodium chloride
    - **d.** sodium monoxide

15. What happens to an aluminum atom when it becomes an Al^{3+} ion?
    - **a.** It loses three electrons.
    - **b.** It gains three electrons.
    - **c.** It loses three protons.
    - **d.** It gains three protons.

16. Which is a correctly balanced equation for the reaction between hydrogen gas and oxygen gas?
    - **a.** $H_2 + O_2 \rightarrow H_2O$
    - **b.** $H_2 + O_2 \rightarrow 2H_2O$
    - **c.** $2H_2 + 2O_2 \rightarrow H_2O$
    - **d.** $2H_2 + O_2 \rightarrow 2H_2O$

17. A homogeneous mixture differs from a heterogeneous mixture in that a homogeneous mixture
   a. is a pure substance.
   b. can be represented by a single chemical formula.
   c. is composed of substances that are uniformly distributed.
   d. contains different regions that have different properties.

18. When NaCl dissolves in water, NaCl is the
   a. solvent.
   b. solute.
   c. solution.
   d. solubility.

19. A base is a compound that produces
   a. OH^- ions when it is dissolved in water.
   b. H^+ ions when it is dissolved in water.
   c. protons when it is dissolved in water.
   d. electrons when it is dissolved in water.

Type B

20. Which molecule is most likely to contain a polar covalent bond?
   a. HCl
   b. H_2
   c. Cl_2
   d. N_2

21. Carbonate (CO_3^{2-}) ions react with potassium (K^+) ions to form the ionic compound potassium carbonate. What is the correct chemical formula for potassium carbonate?
   a. CO_3K_2
   b. CO_3K
   c. K_2CO_3
   d. KCO_3

22. Which solution has the greatest concentration?
   a. 0.1 g of LiCl in 1 g of water
   b. 1 g of LiCl in 1 g of water
   c. 1 g of LiCl in 2 g of water
   d. 1 g of LiCl in 10 g of water

23. A student measured the pH of four materials and obtained the following results.

Material	pH
Baking soda	8.9
Black tea	5.0
Lye	13.8
Vinegar	3.0

Which material is the most acidic?
   a. baking soda
   b. black tea
   c. lye
   d. vinegar

Base your answers to questions 24 and 25 on the information below.

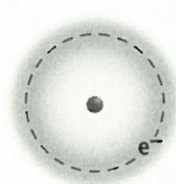
> A heat pack can be used to warm the hands in cold weather. When the pack is opened, oxygen from the air reacts with iron in the pack, producing iron oxide and releasing heat in the following reaction.
>
> $$3O_2 + 4Fe \rightarrow 2Fe_2O_3$$

24. Is there more energy in the reactants or the product of this reaction?

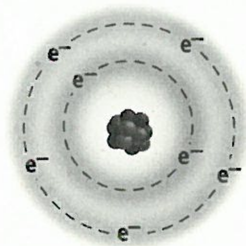
25. Explain the energy changes in this reaction in terms of the law of conservation of energy.

Type C

Base your answers to questions 26 and 27 on the diagrams of a hydrogen atom and a nitrogen atom below. The atoms are shown in their ground states.



Hydrogen atom



Nitrogen atom

Hydrogen and nitrogen combine to form a molecule, ammonia, in which the outermost energy level of each atom is full.

26. How many atoms of hydrogen and nitrogen are in a molecule of ammonia?

27. Draw a diagram of ammonia, showing all of the electrons in the molecule.

28. Glucose ($C_6H_{12}O_6$) combines with oxygen (O_2) to produce carbon dioxide and water. Write a balanced equation to describe this chemical reaction.

29. A chemist analyzes a city's drinking water supply and finds 0.015 g of lead in a 1000-g sample of the water. What is the concentration of lead in the water in parts per million?

30. The presence of certain pollutants in the atmosphere leads to the production of acid precipitation, which can have a pH as low as 1.5. Why does acid precipitation pose a threat to buildings and statues made of limestone?

QUESTIONS FOR SUBTOPIC C

Type A

31. Solids that lack a crystalline structure are known as

a. polycrystalline solids.

b. glasses.

c. hexagonal solids.

d. plasmas.

32. Which statement about gases is correct?

a. Gases have a definite shape.

b. Gases have a fixed volume.

c. Gas particles travel along curved paths.

d. Gas particles are separated by larger distances than are the particles in a liquid.

Type B

33. The particles in a sample of matter are arranged in a regular geometric pattern. As they absorb thermal energy, they enter the state of matter in which they have no attraction for one another and move independently at high speeds. The change of state the particles are undergoing is called

a. melting. **c.** sublimation.

b. evaporation. **d.** condensation.

34. State two examples of plasmas that occur on Earth.

Type C

35. The change of state from a liquid to a solid is called freezing. Explain freezing in terms of the transfer of thermal energy and the attractive forces between particles of matter.