UNIT 1

Chemistry and Cells

Chapter 2

The Chemical Context of Life

Chapter Focus

This chapter considers the basic principles of chemistry that explain the behavior of atoms and molecules. You will learn how the subatomic particles—protons, neutrons, and electrons—are organized in atoms, how atoms are connected by covalent bonds, and how ions are attracted to each other in ionic bonds. The chapter also focuses on the properties of water, which emerge from the polarity and hydrogen bonding capacity of this small, essential molecule.

Chapter Review

2.1 Matter consists of chemical elements in pure form and in combinations called compounds

Elements and Compounds **Matter** is anything that takes up space and has mass. **Elements** are substances that cannot be chemically broken down to other types of matter. A **compound** is made up of two or more elements combined in a fixed ratio. The characteristics of a compound differ from those of its constituent elements, an example of *emergent properties* arising in higher levels of organization.

The Elements of Life Your body is composed of 25 different elements. The elements needed for an organism to live and reproduce are called **essential elements** (the list varies somewhat for different organisms). Carbon (C), oxygen (O), hydrogen (H), and nitrogen (N) make up 96% of living matter. Some elements, like iron (Fe) and iodine (I), may be required in very minute quantities and are called **trace elements**.

FOCUS QUESTION 2.1

Fill in the names beside the symbols of the following elements that, along with a few others, make up about 4% of an organism's mass.

Ca	K
P	S

Evolution of Tolerance to Toxic Elements Some plants exhibit evolutionary adaptations that allow them to grow in soils containing toxic elements.

2.2 An element's properties depend on the structure of its atoms

Each element has its own type of **atom**, the smallest unit of matter retaining the properties of that element.

Subatomic Particles Three *subatomic particles* are important to your understanding of atoms. Uncharged **neutrons** and positively charged **protons** are packed tightly together to form the **atomic nucleus** of an atom. Negatively charged **electrons** form a cloud around the nucleus.

Protons and neutrons have a similar mass of about 1.7×10^{-24} g, or close to 1 dalton each. A **dalton** is the measurement unit for atomic mass. Electrons have negligible mass.

Atomic Number and Atomic Mass So what makes the atoms of different elements different? Each element has a characteristic **atomic number**, or number of protons in the nucleus of each of its atoms. Unless otherwise indicated, an atom has a neutral electrical charge, and thus the number of protons is equal to the number of electrons. A subscript to the left of the symbol for an element indicates its atomic number; a superscript indicates its mass number. The **mass number** is equal to the number of protons and neutrons in the nucleus and approximates the mass of an atom of that element in daltons. The term **atomic mass** refers to the total mass of an atom.

FOCUS QUESTION 2.2

The difference between the mass number and the atomic number of an atom is equal to the number of __________. An atom of phosphorus, $^{31}_{15}P$, contains __________ protons, __________ electrons, and __________ neutrons. The atomic mass of phosphorus is approximately __________.

Isotopes Although the number of protons is constant, the number of neutrons can vary among the atoms of an element, creating different **isotopes** that have slightly different masses. Some isotopes are unstable; the nuclei of **radioactive isotopes** spontaneously decay, giving off particles and energy. Radioactive isotopes are important tools in biological research and medicine. Too great an exposure to radiation from decaying isotopes poses a significant health hazard.

The Energy Levels of Electrons **Energy** is defined as the capacity to cause change—to do work. **Potential energy** is energy stored in matter as a consequence of its position or structure. The potential energy of electrons increases as their distance from the positively charged nucleus increases. Electrons can be located in different **electron shells**, each with a characteristic energy level and distance from the nucleus.

FOCUS QUESTION 2.3

To move to a shell farther from the nucleus, an electron must __________ energy; an electron __________ energy when it moves to a closer shell.

Electron Distribution and Chemical Properties What determines the chemical behavior of an atom? It is a function of the distribution of its electrons—in particular, the number of **valence electrons** in its outermost electron shell, or **valence shell**. A valence shell of eight electrons is complete, resulting in an unreactive or inert atom. (Remember that the first shell can hold only two electrons.) Atoms with incomplete valence shells are chemically reactive. The elements in each row, or period, of the *periodic table of the elements* have the same number of electron shells and are arranged in order of increasing number of electrons in the outer shell.

FOCUS QUESTION 2.4

Draw an electron distribution diagram for the following atoms. *(Note that electrons do not pair up in the second and third shells until after four electrons are present.)*

a. $_{1}H$

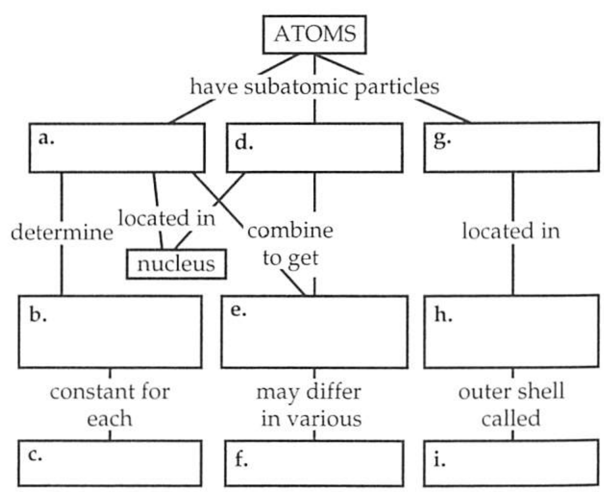
b. $_{6}C$

c. $_{8}O$

d. $_{11}Na$

FOCUS QUESTION 2.5

Fill in the blanks in the following concept map to help you review the atomic structure of atoms.



2.3 The formation and function of molecules depend on chemical bonding between atoms

Atoms with incomplete valence shells can either share electrons with or completely transfer electrons to or from other atoms such that each atom is able to complete its valence shell. These interactions usually result in attractions called **chemical bonds**, which hold the atoms close together.

Covalent Bonds When two atoms share a pair of valence electrons, a **covalent bond** is formed. A **molecule** consists of two or more atoms held together by covalent bonds. An electron distribution diagram shows the shared electrons in a molecule. In a *structural formula*, such as H—H, the line indicates a **single bond**. A *molecular formula*, such as O_2, indicates only the kinds and numbers of atoms. In an oxygen molecule, two pairs of valence electrons are shared between oxygen atoms, forming a double covalent bond, or simply a **double bond** (O=O). The **valence**, or bonding capacity, of an atom usually equals the number of electrons required to complete its valence shell.

FOCUS QUESTION 2.6

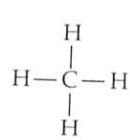
What are the valences of the four most common elements of living matter?

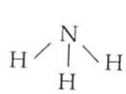
Electronegativity is the attraction of a particular atom for shared electrons. If the atoms in a molecule have similar electronegativities, the electrons remain equally shared, and the bond is said to be a **nonpolar covalent bond**. If one element is more electronegative, it pulls the shared electrons closer to itself, creating a **polar covalent bond**. This unequal sharing of electrons results in a "polarity" or separation of charges, with a slight negative charge ($\delta-$) associated with the more electronegative atom and a slight positive charge ($\delta+$) associated with the atom from which the electrons are pulled.

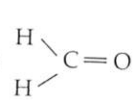
FOCUS QUESTION 2.7

Explain whether the following molecules contain nonpolar or polar covalent bonds. (*Hint:* N and O both have high electronegativities. C and H have lower, and similar, electronegativities.)

a. nitrogen molecule N≡N

c. methane 

b. ammonia 

d. formaldehyde 

Ionic Bonds What happens when two atoms are very unequal in their attraction for valence electrons? The more electronegative atom may completely transfer an electron from the other atom, resulting in the formation of charged atoms called **ions**. The atom that lost the electron is a positively charged **cation**. The negatively charged atom that gained the electron is called an **anion**. An **ionic bond** may hold these ions together because of the attraction of their opposite charges.

Ionic compounds, or **salts**, often exist as three-dimensional crystalline lattice arrangements held together by electrical attractions. The number of ions present in a salt crystal is not fixed, but the atoms are present in specific ratios. Salts have strong ionic bonds when dry, but the crystal dissolves in water.

Covalent molecules that are electrically charged are also referred to as *ions*.

FOCUS QUESTION 2.8

Calcium ($_{20}Ca$) and chlorine ($_{17}Cl$) can combine to form the salt calcium chloride. Based on the number of electrons in their valence shells and their bonding capacities, what would the formula for this salt be? **a.** ___________ Which atom becomes the cation? **b.** ___________

Weak Chemical Bonds Ionic bonds and other weak bonds may form temporary interactions between molecules. Weak bonds within many large molecules help to create those molecules' three-dimensional functional shapes.

A hydrogen atom that is covalently bonded to an electronegative atom has a partial positive charge and can be attracted to another nearby electronegative atom. This attraction is called a **hydrogen bond**.

All atoms and molecules are attracted to each other when in close contact by **van der Waals interactions**. Momentary uneven electron distributions produce changing positive and negative regions that create these weak attractions.

Molecular Shape and Function A molecule's characteristic size and shape affect how it interacts with other molecules. A carbon atom bonded to four other atoms has a tetrahedral shape.

FOCUS QUESTION 2.9

Draw the electron distribution diagram of a water molecule, showing its V shape and covalently shared electrons. Indicate the areas with slight negative and positive charges that enable a water molecule to form hydrogen bonds with other polar molecules. Then draw a second water molecule and indicate a hydrogen bond between the two.

2.4 Chemical reactions make and break chemical bonds

Chemical reactions involve the making or breaking of chemical bonds. Matter is conserved in chemical reactions; the same number and kinds of atoms are present in both **reactants** and **products**, although the rearrangement of atoms causes the properties of these molecules to be different.

Chemical reactions are reversible—the products of the forward reaction can become reactants in the reverse reaction. Increasing the concentrations of reactants can speed up the rate of a reaction. **Chemical equilibrium** is reached when the forward and reverse reactions proceed at the same rate, and the relative concentrations of reactants and products no longer change.

FOCUS QUESTION 2.10

Fill in the missing coefficients for respiration, the conversion of glucose and oxygen to carbon dioxide and water, with the release of energy. Make sure that all atoms are conserved in the chemical reaction.

$C_6H_{12}O_6$ + _____ O_2 → _____ CO_2 + _____ H_2O + Energy

2.5 Hydrogen bonding gives water properties that help make life possible on Earth

The V-shaped water molecule is a **polar molecule** with a slight positive charge on each hydrogen atom ($\delta+$) and a slight negative charge ($\delta-$) associated with the oxygen. Hydrogen bonds between water molecules create a structural organization that leads to the emergent properties of water.

Cohesion of Water Molecules Liquid water is unusually cohesive due to the constant forming and reforming of hydrogen bonds that hold the molecules close together. This **cohesion** creates a more structurally organized liquid and helps water to be pulled upward in plants. The **adhesion** of water molecules to the walls of plant vessels also contributes to water transport. Hydrogen bonding between water molecules produces a high **surface tension** at the interface between water and air, making the surface unusually difficult to break.

Moderation of Temperature by Water **Thermal energy** is a measure of the **kinetic energy** associated with the random movement of atoms and molecules. **Temperature** measures the *average* kinetic energy of the molecules in a body of matter; thermal energy reflects the *total* kinetic energy in that matter, which relates to the volume of the body of matter. The thermal energy that transfers from a warmer to a cooler body of matter is defined as **heat**.

A **calorie (cal)** is the amount of heat it takes to raise 1 g of water 1°C. A **kilocalorie (kcal)** is 1,000 calories, the amount of heat required or released to change the temperature of 1 kg of water by 1°C. A **joule (J)** equals 0.239 cal; a calorie equals 4.184 J.

Specific heat is the amount of heat absorbed or lost when 1 g of a substance changes its temperature by 1°C. Water's specific heat of 1 cal/g·°C is unusually high compared with other common substances. Why does water absorb or release a relatively large quantity of heat as its temperature changes? Heat must be absorbed to break hydrogen bonds before water molecules can move faster and the temperature can rise; conversely, heat is released when hydrogen bonds form as the temperature of water drops. The high proportion of water in the environment and within organisms keeps temperature fluctuations within limits that permit life.

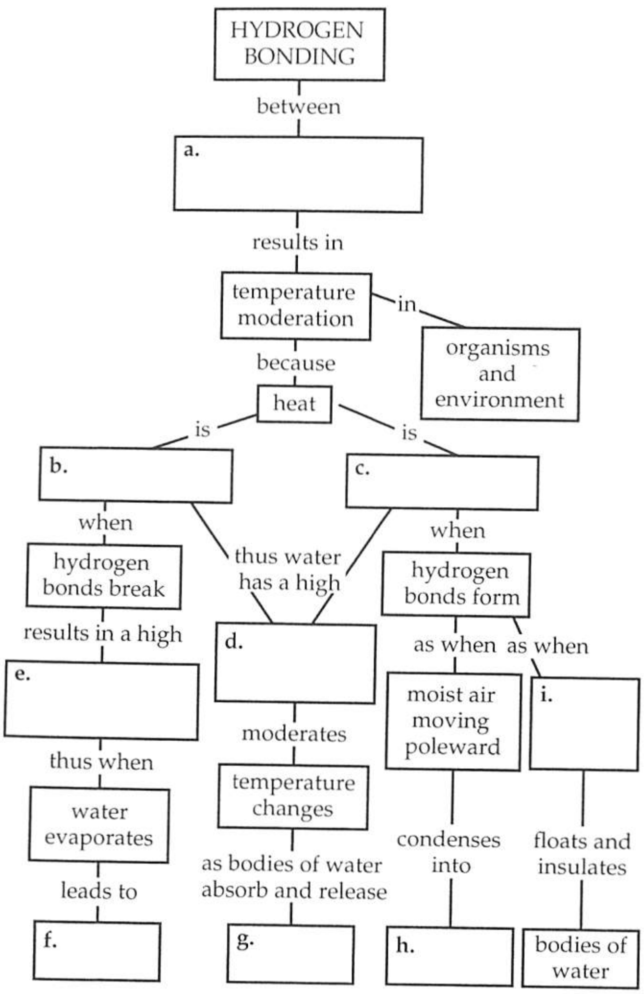
Vaporization or *evaporation* occurs when molecules with sufficient kinetic energy overcome their attraction to other molecules in a liquid and escape into the air as a gas. The **heat of vaporization** is the quantity of heat that must be absorbed for 1 g of a liquid to be converted to a gas. Water's high heat of vaporization is again related to the large amount of heat needed to break the hydrogen bonds holding water molecules together. Water helps moderate Earth's climate as solar heat absorbed by tropical seas is dissipated during evaporation, and heat is released as moist tropical air moving poleward condenses to form rain.

As a liquid vaporizes, the surface left behind loses the kinetic energy of the escaping molecules and cools down. **Evaporative cooling** helps to protect terrestrial organisms from overheating and contributes to the stability of temperatures in lakes and ponds.

Floating of Ice on Liquid Water As water cools below 4°C, it expands. By 0°C, each water molecule is hydrogen-bonded to four other molecules, creating a crystalline lattice that spaces the molecules apart. Ice is thus less dense than liquid water and so it floats.

FOCUS QUESTION 2.11

The following concept map is one way to show how the breaking and forming of hydrogen bonds are related to temperature moderation. Fill in the blanks and compare your choice of concepts to those given in the answer section. Or, even better, create your own map to help you understand how water stabilizes temperature.



Water: The Solvent of Life A **solution** is a liquid homogeneous mixture of two or more substances; the dissolving agent is called the **solvent**, and the substance that is dissolved is the **solute**. Water is the solvent in an **aqueous solution**. The positive and negative regions of water molecules are attracted to oppositely charged ions or partially charged regions of polar molecules. Thus, solute molecules become surrounded by water molecules (a **hydration shell**) and dissolve into solution.

Ionic and polar substances are **hydrophilic**; they have an affinity for water due to electrical attractions and hydrogen bonding. Nonpolar and nonionic substances are **hydrophobic**; they will not easily mix with or dissolve in water.

Most of the chemical reactions of life take place in water. A **mole (mol)** is the amount of a substance that has a mass in *grams* numerically equivalent to its **molecular mass** (the sum of the mass of all atoms in the molecule). A mole of any substance has exactly the same number of molecules—6.02×10^{23}, called Avogadro's number. The **molarity** of a solution (abbreviated M) refers to the number of moles of a solute dissolved in 1 liter of solution.

FOCUS QUESTION 2.12

a. How many grams of lactic acid ($C_3H_6O_3$) are in 1 liter of a 0.5 M solution of lactic acid (^{12}C, ^{1}H, ^{16}O)?

b. How many molecules of lactic acid are in the solution in a?

Acids and Bases A water molecule can dissociate into a **hydrogen ion**, H^+ (which binds to another water molecule to form a **hydronium ion**, H_3O^+), and a **hydroxide ion**, OH^-. In pure water at 25°C, the concentrations of H^+ and OH^- are the same; both are equal to 10^{-7} M.

When acids or bases dissolve in water, the H^+ and OH^- balance shifts. An **acid** adds H^+ to a solution, whereas a **base** reduces H^+ in a solution by accepting hydrogen ions or by adding hydroxide ions (which then combine with H^+ and thus remove hydrogen ions). A strong acid or strong base dissociates completely when mixed with water. A weak acid or base reversibly dissociates, either releasing or binding H^+.

In an aqueous solution, the *product* of the $[H^+]$ and $[OH^-]$ is constant at 10^{-14}. Brackets, [ ], indicate molar concentration. If the $[H^+]$ is higher, then the $[OH^-]$ is lower, because the excess hydrogen ions combine with the hydroxide ions in solution and form water. Likewise, an increase in $[OH^-]$ causes an equivalent decrease in $[H^+]$.

The **pH** of a solution is defined as the negative log (base 10) of the $[H^+]$: $pH = -\log [H^+]$. For a neutral aqueous solution, $[H^+]$ is 10^{-7} M, and the $pH = 7$. As the $[H^+]$ increases in an acidic solution, the pH value decreases. (This inverse relationship makes sense because the exponent becomes smaller: 10^{-4} indicates a higher $[H^+]$ than 10^{-7}.) The difference between each unit of the pH scale represents a tenfold difference in the concentration of $[H^+]$ and $[OH^-]$.

FOCUS QUESTION 2.13

Complete the following table to review your understanding of pH.

$[H^+]$	$[OH^-]$	pH	Acidic, Basic, or Neutral?
10^{-8}			
	$[10^{-7}]$		
		1	

Buffers within a cell maintain a stable pH (usually close to 7) by accepting excess H^+ or donating H^+ when H^+ concentration decreases. Weak acid-base pairs that reversibly bind hydrogen ions are typical of most buffering systems.

FOCUS QUESTION 2.14

The carbonic acid/bicarbonate system is an important biological buffer. Label the molecules and ions in this equation, and indicate which is the H^+ donor and which is the acceptor.

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

In which direction will this reaction proceed

a. when the pH of a solution begins to fall?

b. when the pH rises above normal level?

The increasing release of CO_2 to the atmosphere is linked to fossil fuel combustion. The oceans absorb about 25% of this CO_2, which lowers the pH of seawater. The resulting ocean acidification decreases the concentration of carbonate (CO_3^{2-}), an important ion needed for coral reef calcification.

FOCUS QUESTION 2.15

a. Add to the formula in Focus Question 2.14 to show why increasing $[CO_2]$ dissolving in water leads to a lower pH.

b. Use this formula to explain how a lower pH would affect the $[CO_3^{2-}]$ in the ocean.

$$HCO_3^- \rightleftharpoons CO_3^{2-} + H^+$$

c. Assuming a fairly constant $[Ca^{2+}]$ in the ocean, how would a change in $[CO_3^{2-}]$ affect the calcification rate—the production of calcium carbonate ($CaCO_3$)—by the coral in a reef ecosystem?

Word Roots

an- = not (*anion:* a negatively charged ion)

co- = together; **-valent** = strength (*covalent bond:* a strong bond in which two atoms share one or more pairs of valence electrons)

electro- = electricity (*electronegativity:* the attraction of a given atom for the electrons of a covalent bond)

hydro- = water; **-philos** = loving; **-phobos** = fearing (*hydrophilic:* having an affinity for water; *hydrophobic:* having no affinity for water)

iso- = equal (*isotope:* one of several forms of an element, each with the same number of protons but a different number of neutrons, thus differing in atomic mass)

kilo- = a thousand (*kilocalorie:* a thousand calories; the amount of heat required to raise the temperature of 1 kg of water by 1°C)

neutr- = neither (*neutron:* a subatomic particle having no electrical charge, found in the nucleus of an atom)

pro- = before (*proton:* a subatomic particle with a single positive electrical charge, found in the nucleus of an atom)

Structure Your Knowledge

Take the time to write out or discuss your answers to the following questions. Then refer to the suggested answers at the end of the book.

1. Fill in the following chart concerning the major subatomic particles of an atom.

Particle	Charge	Mass	Location

2. Atoms can have various numbers associated with them.
   a. Define the following and show where each of them is placed relative to the symbol of an element such as C (use the most common isotope, C-12): atomic number, mass number, atomic mass.
   b. Define valence.
   c. Which of these four numbers is most related to the chemical behavior of an atom? Explain.
3. Fill in the following table, which summarizes the emergent properties of water that contribute to the fitness of the environment for life.

Property	Explanation of Property	Example of Benefit to Life
a.	Hydrogen bonds hold water molecules together and adhere them to a hydrophilic surface.	b.
High specific heat	c.	Temperature changes in environment and organisms are moderated.
d.	Hydrogen bonds must be broken for water to evaporate.	e.
f.	Water molecules with high kinetic energy evaporate; remaining molecules are cooler.	g.
Less dense as a solid	h.	i.
j.	k.	Most chemical reactions in life involve solutes dissolved in water.

Test Your Knowledge

MULTIPLE CHOICE: *Choose the one best answer.*

1. Each element has its own characteristic atom in which
   a. the atomic mass is constant.
   b. the atomic number is constant.
   c. the mass number is constant.
   d. Two of the above are correct.
   e. All of the above are correct.
2. Which of the following is *not* a trace element in the human body?
   a. iodine
   b. zinc
   c. iron
   d. calcium
   e. fluorine
3. A sodium ion (Na^+) contains 10 electrons, 11 protons, and 12 neutrons. What is the atomic number of sodium?
   a. 10
   b. 11
   c. 12
   d. 23
   e. 33
4. Radioactive isotopes can be used in studies of metabolic pathways because
   a. their half-life allows a researcher to time an experiment.
   b. they are more reactive.
   c. the cell does not recognize the extra protons in the nucleus, so isotopes are readily used in metabolism.
   d. their location or quantity can be experimentally determined because of their radioactivity.
   e. their extra neutrons produce different colors that can be traced through the body.
5. Which of the following atomic numbers would describe the element that is least reactive?
   a. 1
   b. 8
   c. 12
   d. 16
   e. 18
6. An atom of argon has three electron shells, all of which are full. Its atomic mass is 40. How many neutrons does it have?
   a. 8
   b. 16
   c. 20
   d. 22
   e. 24
7. Which of the following describes what happens as a chlorophyll pigment absorbs energy from sunlight?
   a. An electron moves to a higher electron shell and the electron's potential energy increases.
   b. An electron moves to a higher electron shell and its potential energy decreases.
   c. An electron drops to a lower electron shell and releases its energy as heat.
   d. An electron drops to a lower electron shell and its potential energy increases.
   e. An electron of sunlight is transferred to chlorophyll, producing a chlorophyll ion with higher potential energy.

Use this information to answer questions 8 through 13.

The six elements most common in living organisms are

$^{12}_{6}C$ $^{16}_{8}O$ $^{1}_{1}H$ $^{14}_{7}N$ $^{32}_{16}S$ $^{31}_{15}P$

8. How many electrons does phosphorus have in its valence shell?
   a. 3
   b. 5
   c. 7
   d. 15
   e. 16

9. What is the atomic mass of phosphorus?
   a. 15
   b. 16
   c. 31
   d. 46
   e. 62

10. A radioactive isotope of carbon has the mass number 14. How many neutrons does this isotope have?
   a. 2
   b. 6
   c. 8
   d. 12
   e. 14

11. How many covalent bonds is a sulfur atom most likely to form?
   a. 1
   b. 2
   c. 3
   d. 4
   e. 5

12. Based on electron configuration, which of the following elements would have chemical behavior most like that of oxygen?
   a. C
   b. H
   c. N
   d. P
   e. S

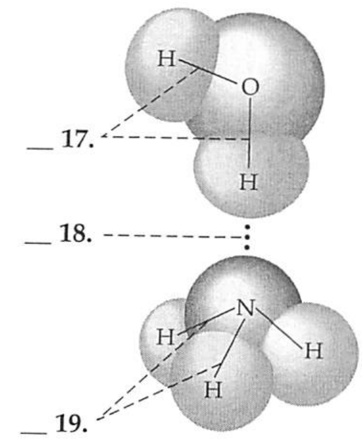
13. How many of the elements listed above are found next to each other (side by side) on the periodic table?
   a. one group of two
   b. two groups of two
   c. one group of two and one group of three
   d. one group of three
   e. all of them

14. A covalent bond between two atoms is likely to be nonpolar if
   a. one of the atoms is much more electronegative than the other.
   b. the two atoms are about equally electronegative.
   c. the two atoms are of the same element.
   d. one atom is an anion and the other is a cation.
   e. Both **b** and **c** are correct.

15. A triple covalent bond would
   a. not be possible.
   b. involve the bonding of three atoms.
   c. involve the bonding of six atoms.
   d. produce a triangularly shaped molecule.
   e. involve the sharing of six electrons.

16. A cation
   a. has gained an electron.
   b. can easily form hydrogen bonds.
   c. is more likely to form in an atom with seven electrons in its valence shell.
   d. has a positive charge.
   e. Both **c** and **d** are correct.

For questions 17 through 19, choose from the following answers to identify the types of bonds in this diagram of a water molecule interacting with an ammonia molecule.



a. nonpolar covalent bond
b. polar covalent bond
c. ionic bond
d. hydrogen bond
e. cannot determine without more information

20. In what type of bond would you expect potassium ($^{39}_{19}K$) to participate?
   a. ionic; it would lose one electron and carry a positive charge
   b. ionic; it would gain one electron and carry a negative charge
   c. covalent; it would share one electron and make one covalent bond
   d. covalent; it would share two electrons and form two bonds
   e. none; potassium is an inert element

21. Which of the following may form between any closely aligned molecules?
   a. nonpolar covalent bonds
   b. polar covalent bonds
   c. ionic bonds
   d. hydrogen bonds
   e. van der Waals interactions

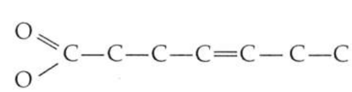
22. What is the molecular shape of methane (CH_4)?
    a. planar or flat, with the four H extending out from the carbon
    b. pentagonal, or a flat five-sided arrangement
    c. tetrahedral, with carbon in the center and H at each corner
    d. circular, with the four H attached in a ring around the carbon
    e. linear, since all the bonds are nonpolar covalent bonds

23. The ability of morphine to mimic the effects of the body's endorphins is due to
    a. a chemical equilibrium developing between morphine and endorphins.
    b. the one-way conversion of morphine into endorphin.
    c. molecular shape similarities that allow morphine to bind to endorphin receptors.
    d. the similarities between morphine and heroin.
    e. hydrogen bonding and other weak bonds forming between morphine and endorphins.

24. Which of the following molecules or compounds would you predict is capable of forming hydrogen bonds?
    a. CH_4
    b. CH_4O
    c. NaCl
    d. H_2
    e. **a**, **b**, and **d** can form hydrogen bonds.

25. Chlorine has an atomic number of 17 and a mass number of 35. How many electrons would a chloride ion have?
    a. 16
    b. 17
    c. 18
    d. 33
    e. 34

26. Taking into account the bonding capacities or valences of carbon (C) and oxygen (O), how many hydrogen (H) must be added to complete the following structural diagram of this molecule?



    a. 9
    b. 10
    c. 11
    d. 12
    e. 13

27. What is the difference between a molecule and a compound?
    a. There is no difference; the terms are interchangeable.
    b. Molecules contain atoms of a single element, whereas compounds contain two or more elements.
    c. A molecule consists of two or more covalently bonded atoms; a compound contains two or more atoms held by ionic bonds.
    d. A compound consists of two or more elements in a fixed ratio; a molecule has two or more covalently bonded atoms of the same or different elements.
    e. Compounds always consist of molecules, but molecules are not always compounds.

28. In a reaction in chemical equilibrium,
    a. the forward and reverse reactions are occurring at the same rate.
    b. the reactants and products are in equal concentration.
    c. the forward reaction has gone further than the reverse reaction.
    d. there are equal numbers of atoms on both sides of the equation.
    e. **a**, **b**, and **d** are correct.

29. What would be the probable effect of adding more product to a reaction that is in equilibrium?
    a. There would be no change because the reaction is in equilibrium.
    b. The reaction would stop because excess product is present.
    c. The reaction would slow down but still continue.
    d. The forward reaction would increase and more product would be formed.
    e. The reverse reaction would increase and more reactants would be formed.

30. What coefficients must be placed in the blanks to balance the following chemical reaction?

$$C_5H_{12} + \_\_\_\_ O_2 \rightarrow \_\_\_\_ CO_2 + \_\_\_\_ H_2O$$

    a. 5; 5; 5
    b. 6; 5; 6
    c. 6; 6; 6
    d. 8; 4; 6
    e. 8; 5; 6

31. The polar covalent bonds of water molecules
    a. promote the formation of hydrogen bonds.
    b. help water to dissolve nonpolar solutes.
    c. lower the heat of vaporization and lead to evaporative cooling.
    d. create a crystalline structure in liquid water.
    e. do all of the above.

32. What contributes to the movement of water up the vessels of a tall tree?
   a. cohesion
   b. hydrogen bonding
   c. adhesion
   d. hydrophilic cell walls
   e. all of the above

33. You have three flasks containing 100 mL of different liquids. Each is warmed with 100 calories of heat. The temperature of the liquid in flask 1 rises 1°C; in flask 2 it rises 1.5°C; and in flask 3 it rises 2°C. Which of these liquids has the highest specific heat?
   a. the liquid in flask 1
   b. the liquid in flask 2
   c. the liquid in flask 3
   d. You cannot tell unless you know what liquid is in each flask.
   e. This type of experiment does not relate to the specific heat of a substance.

34. Climates tend to be moderate near large bodies of water because
   a. a large amount of solar heat is absorbed during the gradual rise in temperature of the water.
   b. water releases heat to the environment as it cools.
   c. the high specific heat of water helps to moderate air temperatures.
   d. a great deal of heat is absorbed and released as hydrogen bonds break or form.
   e. all of the above are true.

35. A burn from steam at 100°C is more severe than a burn from boiling water because
   a. the steam is hotter than boiling water.
   b. steam releases a great deal of heat as it condenses on the skin.
   c. steam has a higher heat of vaporization than does water.
   d. a person is more likely to come into contact with steam than with boiling water.
   e. steam stays on the skin longer than does boiling water.

36. Ice floats because
   a. air is trapped in the crystalline lattice.
   b. the formation of hydrogen bonds releases heat; warmer objects float.
   c. it has a smaller surface area than liquid water.
   d. it insulates bodies of water so they do not freeze from the bottom up.
   e. hydrogen bonding spaces the molecules farther apart, creating a less dense structure.

37. Why is water such an excellent solvent?
   a. As a polar molecule, it can surround and dissolve ionic and polar molecules.
   b. It forms ionic bonds with ions, hydrogen bonds with polar molecules, and hydrophobic interactions with nonpolar molecules.
   c. It forms hydrogen bonds with itself.
   d. It has a high specific heat and a high heat of vaporization.
   e. It is liquid and has a high surface tension.

38. The molarity of a solution is equal to
   a. Avogadro's number of molecules in 1 liter of solvent.
   b. the number of moles of a solute in 1 liter of solution.
   c. the molecular mass of a solute in 1 liter of solution.
   d. the number of solute molecules in 1 liter of solvent.
   e. 342 g if the solute is sucrose.

39. Which of the following substances would you add to enough water to yield 1 liter of solution in order to make a 0.1 *M* solution of glucose ($C_6H_{12}O_6$)? The mass numbers for these elements are approximately C = 12, O = 16, and H = 1.
   a. 6 g C, 12 g H, and 6 g O
   b. 72 g C, 12 g H, and 96 g O
   c. 18 g of glucose
   d. 29 g of glucose
   e. 180 g of glucose

40. How many molecules of glucose would be in 1 liter of the 0.1 *M* solution made in question 39?
   a. 0.1
   b. 6
   c. 60
   d. 6×10^{23}
   e. 6×10^{22}

41. Adding a base to a solution would
   a. raise the pH.
   b. lower the pH.
   c. decrease $[H^+]$.
   d. do both **a** and **c**.
   e. do both **b** and **c**.

42. Some archaea are able to live in lakes with pH values of 11. How does pH 11 compare with the pH 7 typical of your body cells?
   a. It is four times more acidic than pH 7.
   b. It is four times more basic than pH 7.
   c. It is a thousand times more acidic than pH 7.
   d. It is a thousand times more basic than pH 7.
   e. It is ten thousand times more basic than pH 7.

43. A buffer
    a. releases excess OH^-.
    b. releases excess H^+.
    c. is often a weak acid-base pair.
    d. always maintains a neutral pH.
    e. Both **c** and **d** are correct.

44. In the past century, the average temperature of the oceans has increased by 0.74°C. Would you consider this evidence of global warming?
    a. No, the rise in temperature is too small to be significant.
    b. No, global warming affects air temperature, not water temperature.
    c. No, the change of average temperature does not reflect the quantity of thermal energy in the oceans.
    d. Yes, because of the high specific heat of water and the huge volume of water in the oceans, a small rise in temperature would reflect a large amount of heat absorbed by the oceans.
    e. Yes, the decreased rate of calcification of reef-building organisms is directly related to this temperature increase.