



This is “Elements, Atoms, and the Periodic Table”, chapter 2 from the book [Introduction to Chemistry: General, Organic, and Biological \(index.html\)](#) (v. 1.0).

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Chapter 2

Elements, Atoms, and the Periodic Table

Opening Essay

The hardest material in the human body is tooth enamel. It has to be hard so that our teeth can serve us for a lifetime of biting and chewing; however, tough as it is, tooth enamel is susceptible to chemical attack. Acids found in some foods or made by bacteria that feed on food residues on our teeth are capable of dissolving enamel. Unprotected by enamel, a tooth will start to decay, thus developing cavities and other dental problems.

In the early 1900s, a dentist in Colorado Springs, Colorado, noted that many people who lived in the area had brown-stained teeth that, while unsightly, were surprisingly resistant to decay. After years of study, excess fluorine compounds in the drinking water were discovered to be the cause of both these effects. Research continued, and in the 1930s, the US Public Health Service found that low levels of fluorine in water would provide the benefit of resisting decay without discoloring teeth.

The protective effects of fluorine have a simple chemical explanation. Tooth enamel consists mostly of a mineral called hydroxyapatite, which is composed of calcium, phosphorus, oxygen, and hydrogen. We know now that fluorine combines with hydroxyapatite to make *fluorapatite*, which is more resistant to acidic decay than hydroxyapatite is. Currently about 50% of the US population drinks water that has some fluorine added (in the form of sodium fluoride, NaF) to reduce tooth decay. This intentional *fluoridation*, coupled with the use of fluoride-containing toothpastes and improved oral hygiene, has reduced tooth decay by as much as 60% in children. The nationwide reduction of tooth decay has been cited as an important public health advance in history. (Another important advance was the eradication of polio.)

Just as a language has an alphabet from which words are built, chemistry has an alphabet from which matter is described. However, the chemical alphabet is larger than the one we use for spelling. You may have already figured out that the chemical alphabet consists of the chemical elements. Their role is central to chemistry, for they combine to form the millions and millions of known compounds.

2.1 The Elements

LEARNING OBJECTIVES

1. Define a chemical element and give examples of the abundance of different elements.
2. Represent a chemical element with a chemical symbol.

An element, as defined in [Chapter 1 "Chemistry, Matter, and Measurement"](#), is a substance that cannot be broken down into simpler chemical substances. There are about 90 naturally occurring elements known on Earth. Using technology, scientists have been able to create nearly 30 additional elements that do not occur in nature. Today, chemistry recognizes 118 elements—some of which were created an atom at a time. [Figure 2.1 "Samples of Elements"](#) shows some of the chemical elements.

Figure 2.1 *Samples of Elements*



Gold is a yellowish solid, iron is a silvery solid, while mercury is a silvery liquid at room temperature.

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Abundance

The elements vary widely in abundance. In the universe as a whole, the most common element is hydrogen (about 90% of atoms), followed by helium (most of the remaining 10%). All other elements are present in relatively minuscule amounts, as far as we can detect.

On the planet Earth, however, the situation is rather different. Oxygen makes up 46.1% of the mass of Earth's crust (the relatively thin layer of rock forming Earth's surface), mostly in combination with other elements, while silicon makes up 28.5%. Hydrogen, the most abundant element in the universe, makes up only 0.14% of Earth's crust. [Table 2.1 "Elemental Composition of Earth"](#) lists the relative abundances of elements on Earth as a whole and in Earth's crust. [Table 2.2 "Elemental Composition of a Human Body"](#) lists the relative abundances of elements in the human body. If you compare [Table 2.1 "Elemental Composition of Earth"](#) and [Table 2.2 "Elemental Composition of a Human Body"](#), you will find disparities between the percentage of each element in the human body and on Earth. Oxygen has the highest percentage in both cases, but carbon, the element with the second highest percentage in the body, is relatively rare on Earth and does not even appear as a separate entry in [Table 2.1 "Elemental Composition of Earth"](#); carbon is part of the 0.174% representing "other" elements. How does the human body concentrate so many apparently rare elements?

Table 2.1 Elemental Composition of Earth

Earth's Crust		Earth (overall)	
Element	Percentage	Element	Percentage
oxygen	46.1	iron	34.6
silicon	28.2	oxygen	29.5
aluminum	8.23	silicon	15.2
iron	5.53	magnesium	12.7
calcium	4.15	nickel	2.4
sodium	2.36	sulfur	1.9
magnesium	2.33	all others	3.7
potassium	2.09		
titanium	0.565		
hydrogen	0.14		

Earth's Crust		Earth (overall)	
Element	Percentage	Element	Percentage
phosphorus	0.105		
all others	0.174		

Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 14–17.

Table 2.2 Elemental Composition of a Human Body

Element	Percentage by Mass
oxygen	61
carbon	23
hydrogen	10
nitrogen	2.6
calcium	1.4
phosphorus	1.1
sulfur	0.20
potassium	0.20
sodium	0.14
chlorine	0.12
magnesium	0.027
silicon	0.026
iron	0.006
fluorine	0.0037
zinc	0.0033
all others	0.174

Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 7–24.

The relative amounts of elements in the body have less to do with their abundances on Earth than with their availability in a form we can assimilate. We obtain oxygen from the air we breathe and the water we drink. We also obtain hydrogen from water. On the other hand, although carbon is present in the atmosphere as carbon dioxide, and about 80% of the atmosphere is nitrogen, we obtain those two elements from the food we eat, not the air we breathe.

Looking Closer: Phosphorus, the Chemical Bottleneck

There is an element that we need more of in our bodies than is proportionately present in Earth's crust, and *this* element is not easily accessible. Phosphorus makes up 1.1% of the human body but only 0.105% of Earth's crust. We need phosphorus for our bones and teeth, and it is a crucial component of all living cells. Unlike carbon, which can be obtained from carbon dioxide, there is no phosphorus compound present in our surroundings that can serve as a convenient source. Phosphorus, then, is nature's bottleneck. Its availability limits the amount of life our planet can sustain.

Higher forms of life, such as humans, can obtain phosphorus by selecting a proper diet (plenty of protein); but lower forms of life, such as algae, must absorb it from the environment. When phosphorus-containing detergents were introduced in the 1950s, wastewater from normal household activities greatly increased the amount of phosphorus available to algae and other plant life. Lakes receiving this wastewater experienced sudden increases in growth of algae. When the algae died, concentrations of bacteria that ate the dead algae increased. Because of the large bacterial concentrations, the oxygen content of the water dropped, causing fish to die in large numbers. This process, called *eutrophication*, is considered a negative environmental impact.

Today, many detergents are made without phosphorus so the detrimental effects of eutrophication are minimized. You may even see statements to that effect on detergent boxes. It can be sobering to realize how much impact a single element can have on life—or the ease with which human activity can affect the environment.

Names and Symbols

Each element has a name. Some of these names date from antiquity, while others are quite new. Today, the names for new elements are proposed by their

discoverers but must be approved by the International Union of Pure and Applied Chemistry, an international organization that makes recommendations concerning all kinds of chemical terminology.

Note

Today, new elements are usually named after famous scientists.

The names of the elements can be cumbersome to write in full, especially when combined to form the names of compounds. Therefore, each element name is abbreviated as a one- or two-letter **chemical symbol**¹. By convention, the first letter of a chemical symbol is a capital letter, while the second letter (if there is one) is a lowercase letter. The first letter of the symbol is usually the first letter of the element's name, while the second letter is some other letter from the name. Some elements have symbols that derive from earlier, mostly Latin names, so the symbols may not contain any letters from the English name. [Table 2.3 "Element Names and Symbols"](#) lists the names and symbols of some of the most familiar elements.

Table 2.3 Element Names and Symbols

aluminum	Al	magnesium	Mg
argon	Ar	manganese	Mn
arsenic	As	mercury	Hg*
barium	Ba	neon	Ne
bismuth	Bi	nickel	Ni
boron	B	nitrogen	N
bromine	Br	oxygen	O
calcium	Ca	phosphorus	P
carbon	C	platinum	Pt
chlorine	Cl	potassium	K*
*The symbol comes from the Latin name of element.			
†The symbol for tungsten comes from its German name—<i>wolfram</i>.			

1. A one- or two-letter abbreviation for an element.

chromium	Cr	silicon	Si
copper	Cu*	silver	Ag*
fluorine	F	sodium	Na*
gold	Au*	strontium	Sr
helium	He	sulfur	S
hydrogen	H	tin	Sn*
iron	Fe	tungsten	W [†]
iodine	I	uranium	U
lead	Pb*	zinc	Zn
lithium	Li	zirconium	Zr
*The symbol comes from the Latin name of element.			
†The symbol for tungsten comes from its German name—<i>wolfram</i>.			

Note

Element names in languages other than English are often close to their Latin names. For example, gold is *oro* in Spanish and *or* in French (close to the Latin *aurum*), tin is *estaño* in Spanish (compare to *stannum*), lead is *plomo* in Spanish and *plomb* in French (compare to *plumbum*), silver is *argent* in French (compare to *argentum*), and iron is *fer* in French and *hierro* in Spanish (compare to *ferrum*). The closeness is even more apparent in pronunciation than in spelling.

EXAMPLE 1

Write the chemical symbol for each element without consulting Table 2.3 "Element Names and Symbols".

1. bromine
2. boron
3. carbon
4. calcium
5. gold

Solution

1. Br
2. B
3. C
4. Ca
5. Au

SKILL-BUILDING EXERCISE

Write the chemical symbol for each element without consulting Table 2.3 "Element Names and Symbols".

1. manganese
2. magnesium
3. neon
4. nitrogen
5. silver

EXAMPLE 2

What element is represented by each chemical symbol?

1. Na
2. Hg
3. P
4. K
5. I

Solution

1. sodium
2. mercury
3. phosphorus
4. potassium
5. iodine

SKILL-BUILDING EXERCISE

What element is represented by each chemical symbol?

1. Pb
2. Sn
3. U
4. O
5. F

CONCEPT REVIEW EXERCISES

1. What is an element?
2. Give some examples of how the abundance of elements varies.
3. Why are chemical symbols so useful? What is the source of the letter(s) for a chemical symbol?

ANSWERS

1. An element is the basic chemical building block of matter; it is the simplest chemical substance.
2. Elements vary from being a small percentage to more than 30% of the atoms around us.
3. Chemical symbols are useful to concisely represent the elements present in a substance. The letters usually come from the name of the element.

KEY TAKEAWAYS

- All matter is composed of elements.
- Chemical elements are represented by a one- or two-letter symbol.

EXERCISES

- Which of the following substances are elements?
 - sodium
 - milk
 - gold
 - water
 - air
 - liquefied nitrogen
- Which of the following substances are elements?
 - paper
 - electricity
 - neon
 - carbon
 - wood
 - concrete
- Write the chemical symbol for each element.
 - silver
 - sulfur
 - nitrogen
 - neon
- Write the chemical symbol for each element.
 - bromine
 - oxygen
 - lithium
 - boron
- Explain why it is improper to write CO as the chemical symbol for cobalt.
- Explain why it is improper to write NO as the chemical symbol for nobelium.
- Complete the following table.

Element Symbol	Element Name
F	
Fe	
I	

Element Symbol	Element Name
Cr	
C	
P	

8. Complete the following table.

Element Symbol	Element Name
Mg	
Mn	
Ca	
Cl	
K	
Pt	

ANSWERS

1.
 - a. element
 - b. not an element
 - c. element
 - d. not an element
 - e. not an element
 - f. element

3.
 - a. Ag
 - b. S
 - c. N
 - d. Ne

5. By convention, the second letter in an element's symbol is always lowercase.

7.

Element Symbol	Element Name
F	fluorine
Fe	iron
I	iodine
Cr	chromium
C	carbon
P	phosphorus

2.2 Atomic Theory

LEARNING OBJECTIVES

1. Explain how all matter is composed of atoms.
2. Describe the modern atomic theory.

Take some aluminum foil. Cut it in half. Now you have two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil.

It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can you take this exercise, at least in theory? Can you continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? (Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century BC.)

The **modern atomic theory**², proposed about 1803 by the English chemist John Dalton ([Figure 2.2 "John Dalton"](#)), is a fundamental concept that states that all elements are composed of atoms. In [Chapter 1 "Chemistry, Matter, and Measurement"](#), we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

2. The fundamental concept that all elements are composed of atoms.

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules, as discussed in Chapter 1 "Chemistry, Matter, and Measurement". Several important elements exist as two-atom combinations and are called **diatomic molecules**³. In representing a diatomic molecule, we use the symbol of the element and include the subscript 2 to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H₂), oxygen (O₂), nitrogen (N₂), fluorine (F₂), chlorine (Cl₂), bromine (Br₂), and iodine (I₂).

Figure 2.2 John Dalton



John Dalton was an English scientist who enunciated the modern atomic theory.

Looking Closer: Atomic Theory

Dalton's ideas are called the *modern* atomic theory because the concept of atoms is very old. The Greek philosophers Leucippus and Democritus originally introduced atomic concepts in the fifth century BC. (The word *atom* comes from the Greek word *atomos*, which means "indivisible" or "uncuttable.") Dalton had something that the ancient Greek philosophers didn't have, however; he had experimental evidence, such as the formulas of simple chemicals and the behavior of gases. In the 150 years or so before Dalton, natural philosophy had been maturing into modern science, and the scientific method was being used to study nature. So when Dalton announced a modern atomic theory, he was proposing a fundamental theory to describe many previous observations of the natural world; he was not just participating in a philosophical discussion.

3. A two-atom grouping that behaves as a single chemical entity.

CONCEPT REVIEW EXERCISES

1. What is the modern atomic theory?
2. What are atoms?

ANSWERS

1. The modern atomic theory states that all matter is composed of atoms.
2. Atoms are the smallest parts of an element that maintain the identity of that element.

KEY TAKEAWAYS

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.

EXERCISES

1. Which of the following elements exist as diatomic molecules?
 - a. helium
 - b. hydrogen
 - c. iodine
 - d. gold
2. Which of the following elements exist as diatomic molecules?
 - a. chlorine
 - b. potassium
 - c. silver
 - d. oxygen
3. Why is it proper to represent the elemental form of helium as He but improper to represent the elemental form of hydrogen as H?
4. Why is it proper to represent the elemental form of chlorine as Cl₂ but improper to represent the elemental form of calcium as Ca₂?

ANSWERS

1.
 - a. no
 - b. yes
 - c. yes
 - d. no
3. Hydrogen exists as a diatomic molecule in its elemental form; helium does not exist as a diatomic molecule.

2.3 The Structure of Atoms

LEARNING OBJECTIVES

1. Describe the three main subatomic particles.
2. State how the subatomic particles are arranged in atoms.

There have been several minor but important modifications to Dalton's atomic theory. For one thing, Dalton considered atoms to be indivisible. We know now that atoms not only can be divided but also are composed of three different kinds of particles with their own properties that are different from the chemical properties of atoms.

Subatomic Particles

The first subatomic particle was identified in 1897 and called the **electron**⁴. It is an extremely tiny particle, with a mass of about 9.109×10^{-31} kg. Experiments with magnetic fields showed that the electron has a negative electrical charge.

By 1920, experimental evidence indicated the existence of a second particle. A **proton**⁵ has the same amount of charge as an electron, but its charge is positive, not negative. Another major difference between a proton and an electron is mass. Although still incredibly small, the mass of a proton is 1.673×10^{-27} kg, which is almost 2,000 times greater than the mass of an electron. Because opposite charges attract each other (while like charges repel each other), protons attract electrons (and vice versa).

Finally, additional experiments pointed to the existence of a third particle. Evidence produced in 1932 established the existence of the **neutron**⁶, a particle with about the same mass as a proton but with no electrical charge.

We understand now that all atoms can be broken down into subatomic particles: protons, neutrons, and electrons. [Table 2.4 "Properties of the Subatomic Particles"](#) lists some of their important characteristics and the symbols used to represent each particle.

4. A subatomic particle with a negative electric charge.
5. A subatomic particle with a positive charge.
6. A subatomic particle with no electric charge.

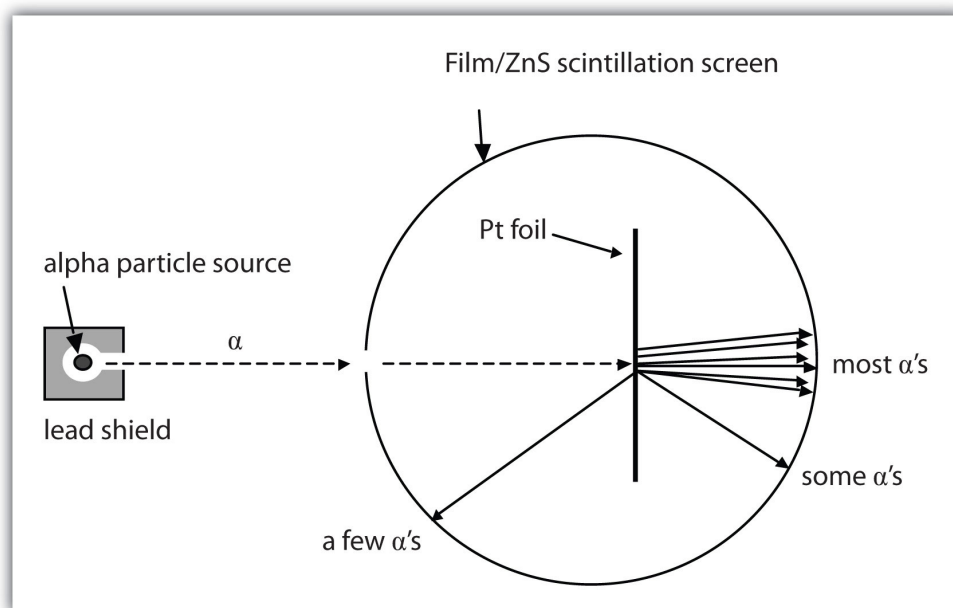
Table 2.4 Properties of the Subatomic Particles

Particle	Symbol	Mass (kg)	Relative Mass (proton = 1)	Relative Charge
proton	p^+	1.673×10^{-27}	1	+1
neutron	n^0	1.675×10^{-27}	1	0
electron	e^-	9.109×10^{-31}	0.00055	-1

The Nucleus

How are these subatomic particles arranged? Between 1909 and 1911, Ernest Rutherford, a Cambridge physicist, and his associates Hans Geiger and Ernest Marsden performed experiments that provided strong evidence concerning the internal structure of an atom. They took a very thin metal foil, such as gold or platinum, and aimed a beam of positively charged particles (called alpha particles, which are combinations of two protons and two neutrons) from a radioactive source toward the foil. Surrounding the foil was a detector—either a scintillator (a material that glows when hit by such particles) or some unexposed film (which is exposed where the particles hit it). The detector allowed the scientists to determine the distribution of the alpha particles after they interacted with the foil. [Figure 2.3 "The Geiger-Marsden Experimental Setup"](#) shows a diagram of the experimental setup.

Figure 2.3 The Geiger-Marsden Experimental Setup

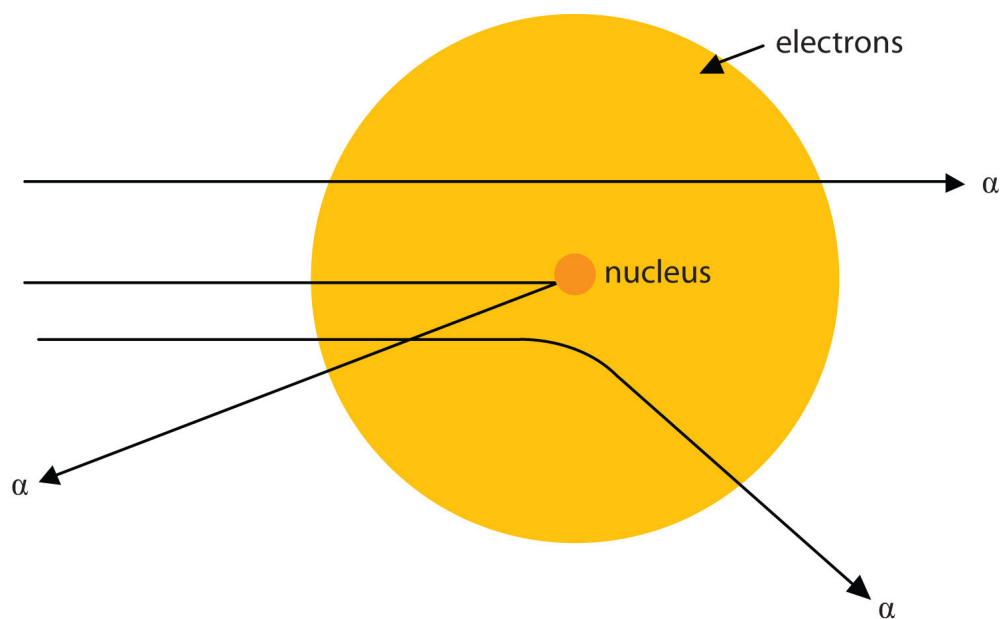


Experiments using this setup were used to investigate the structure of atoms.

Most of the particles traveled straight through the foil, but some alpha particles were deflected off to one side. Some were even deflected back toward the source. This was unexpected. Rutherford once said, “It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.”

Rutherford proposed the following model to explain these experimental results. Protons and neutrons are concentrated in a central region he called the **nucleus**⁷ (plural, *nuclei*) of the atom. Electrons are outside the nucleus and orbit about it because they are attracted to the positive charge in the nucleus. Most of the mass of an atom is in the nucleus, while the orbiting electrons account for an atom’s size. As a result, an atom consists largely of empty space. Rutherford called his description the “planetary model” of the atom. [Figure 2.4 "Rutherford's Metal-Foil Experiments"](#) shows how this model explains the experimental results.

Figure 2.4 Rutherford's Metal-Foil Experiments



Rutherford explained the results of the metal-foil experiments by proposing that most of the mass and the positive charge of an atom are located in its nucleus, while the relatively low-mass electrons orbit about the nucleus. Most alpha particles go straight through the empty space, a few particles are deflected, and fewer still ricochet back toward the source. The nucleus is much smaller proportionately than depicted here.

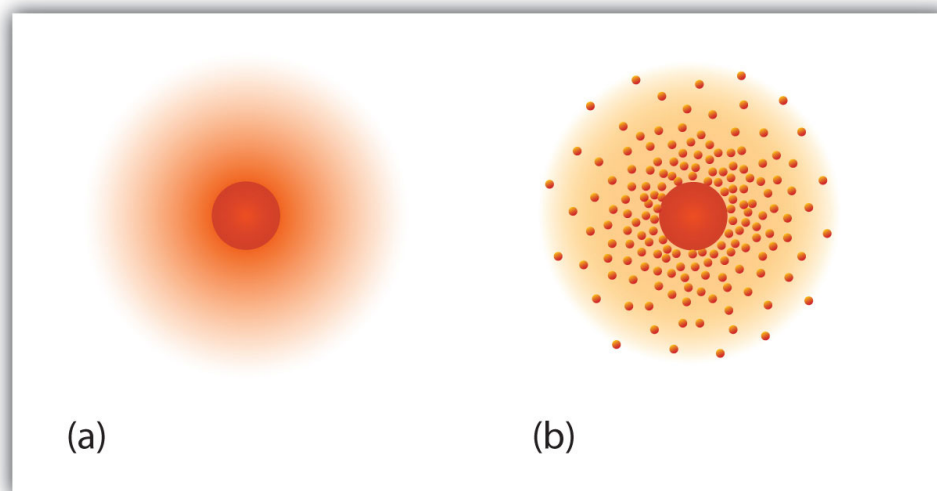
7. The central part of an atom that contains protons and neutrons.

Note

The planetary model of the atom replaced the plum pudding model, which had electrons floating around aimlessly like plums in a “pudding” of positive charge.

Rutherford’s model is essentially the same model that we use today to describe atoms but with one important modification. The planetary model suggests that electrons occupy certain specific, circular orbits about the nucleus. We know now that this model is overly simplistic. A better description is that electrons form fuzzy clouds around nuclei. [Figure 2.5 "A Modern Depiction of Atomic Structure"](#) shows a more modern version of our understanding of atomic structure.

Figure 2.5 A Modern Depiction of Atomic Structure



A more modern understanding of atoms, reflected in these representations of the electron in a hydrogen atom, is that electrons occupy regions of space about the nucleus; they are not in discrete orbits like planets around the sun. (a) The darker the color, the higher the probability that an electron will be at that point. (b) In a two-dimensional cross section of the electron in a hydrogen atom, the more crowded the dots, the higher the probability that an electron will be at that point. In both (a) and (b), the nucleus is in the center of the diagram.

CONCEPT REVIEW EXERCISES

1. What are the charges and the relative masses of the three subatomic particles?
2. Describe the structure of an atom in terms of its protons, neutrons, and electrons.

ANSWERS

1. proton: +1, large; neutron: 0, large; electron: -1, small
2. Protons and neutrons are located in a central nucleus, while electrons orbit about the nucleus.

KEY TAKEAWAYS

- Atoms are composed of three main subatomic particles: protons, neutrons, and electrons.
- Protons and neutrons are grouped together in the nucleus of an atom, while electrons orbit about the nucleus.

EXERCISES

1. Which is smaller—an electron or a helium atom?
2. Which is larger—a proton or an atom of lead?
3. Which subatomic particle has a positive charge? Which subatomic particle has a negative charge?
4. Which subatomic particle is electrically neutral? Does it exist inside or outside the nucleus?
5. Protons are among the (most, least) massive subatomic particles, and they are found (inside, outside) the nucleus.
6. Electrons are among the (most, least) massive subatomic particles, and they are found (inside, outside) the nucleus.
7. Describe why Rutherford used the term *planetary model* to describe his model of atomic structure.
8. Why is the planetary model not an appropriate way to describe the structure of an atom?
9. What happened to most of the alpha particles in Rutherford's experiment? Explain why that happened.
10. Electrons account for the (majority, minority) of the (mass, volume) of an atom.

ANSWERS

1. An electron is smaller.
3. proton; electron
5. most; inside
7. Electrons are in orbit about the nucleus.
9. Most of the alpha particles went through the metal sheet because atoms are mostly empty space.

2.4 Nuclei of Atoms

LEARNING OBJECTIVES

1. Define and differentiate between the atomic number and the mass number of an element.
2. Explain how isotopes differ from one another.

Now that we know how atoms are generally constructed, what do atoms of any particular element look like? How many protons, neutrons, and electrons are in a specific kind of atom?

First, if an atom is electrically neutral overall, then the number of protons equals the number of electrons. Because these particles have the same but opposite charges, equal numbers cancel out, producing a neutral atom.

Atomic Number

In the 1910s, experiments with X rays led to this useful conclusion: the magnitude of the positive charge in the nucleus of every atom of a particular element is the same. In other words, all atoms of the same element have the same number of protons. Furthermore, different elements have a different number of protons in their nuclei, so the number of protons in the nucleus of an atom is characteristic of a particular element. This discovery was so important to our understanding of atoms that the number of protons in the nucleus of an atom is called the **atomic number**⁸.

For example, hydrogen has the atomic number 1; all hydrogen atoms have 1 proton in their nuclei. Helium has the atomic number 2; all helium atoms have 2 protons in their nuclei. There is no such thing as a hydrogen atom with 2 protons in its nucleus; a nucleus with 2 protons would be a helium atom. The atomic number *defines* an element. [Chapter 21 "Appendix: Periodic Table of the Elements"](#) lists the elements and their atomic numbers. From this table, you can determine the number of protons in the nucleus of any element. The largest atoms have over 100 protons in their nuclei.

8. The number of protons in the nucleus of an atom.

EXAMPLE 3

What is the number of protons in the nucleus of each element? (Use the table in [Chapter 21 "Appendix: Periodic Table of the Elements"](#).)

1. aluminum
2. iron
3. carbon

Solution

1. According to the table, aluminum has an atomic number of 13. Therefore, every aluminum atom has 13 protons in its nucleus.
2. Iron has an atomic number of 26. Therefore, every iron atom has 26 protons in its nucleus.
3. Carbon has an atomic number of 6. Therefore, every carbon atom has 6 protons in its nucleus.

SKILL-BUILDING EXERCISE

What is the number of protons in the nucleus of each element? (Use the table in [Chapter 21 "Appendix: Periodic Table of the Elements"](#).)

1. sodium
2. oxygen
3. chlorine

How many electrons are in an atom? Previously we said that for an electrically neutral atom, the number of electrons equals the number of protons, so the total opposite charges cancel. Thus, the atomic number of an element also gives the number of electrons in an atom of that element. (Later we will find that some elements may gain or lose electrons from their atoms, so those atoms will no longer be electrically neutral. Thus we will need a way to differentiate the number of electrons for those elements.)

EXAMPLE 4

How many electrons are present in the atoms of each element?

1. sulfur
2. tungsten
3. argon

Solution

1. The atomic number of sulfur is 16. Therefore, in a neutral atom of sulfur, there are 16 electrons.
2. The atomic number of tungsten is 74. Therefore, in a neutral atom of tungsten, there are 74 electrons.
3. The atomic number of argon is 18. Therefore, in a neutral atom of argon, there are 18 electrons.

SKILL-BUILDING EXERCISE

How many electrons are present in the atoms of each element?

1. magnesium
2. potassium
3. iodine

Isotopes

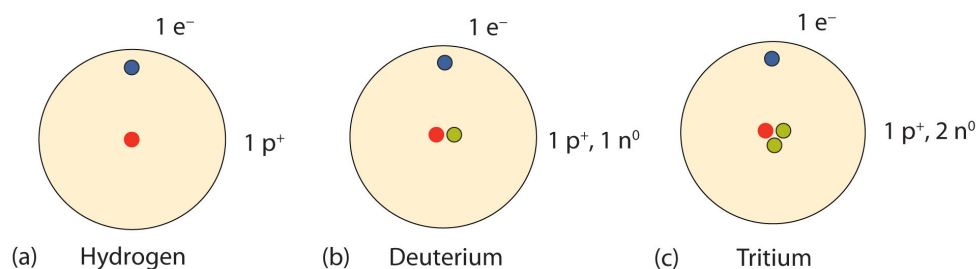
How many neutrons are in atoms of a particular element? At first it was thought that the number of neutrons in a nucleus was also characteristic of an element. However, it was found that atoms of the same element can have *different* numbers of neutrons. Atoms of the same element that have different numbers of neutrons are called **isotopes**⁹. For example, 99% of the carbon atoms on Earth have 6 neutrons and 6 protons in their nuclei; about 1% of the carbon atoms have 7 neutrons in their nuclei. Naturally occurring carbon on Earth, therefore, is actually a mixture of isotopes, albeit a mixture that is 99% carbon with 6 neutrons in each nucleus.

9. Atoms of the same element that have different numbers of neutrons.

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei,

however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. [Figure 2.6 "Isotopes of Hydrogen"](#) compares the three isotopes of hydrogen.

Figure 2.6 *Isotopes of Hydrogen*



Most hydrogen atoms have only a proton in the nucleus (a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

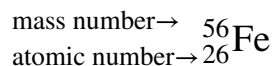
Note

The discovery of isotopes required a minor change in Dalton's atomic theory. Dalton thought that all atoms of the same element were exactly the same.

Most elements exist as mixtures of isotopes. In fact, there are currently over 3,500 isotopes known for all the elements. When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. The **mass number**¹⁰ of an atom is the sum of the numbers of protons and neutrons in the nucleus. Given the mass number for a nucleus (and knowing the atomic number of that particular atom), you can determine the number of neutrons by subtracting the atomic number from the mass number.

A simple way of indicating the mass number of a particular isotope is to list it as a superscript on the left side of an element's symbol. Atomic numbers are often listed as a subscript on the left side of an element's symbol. Thus, we might see

10. The sum of the numbers of protons and neutrons in a nucleus of an atom.



which indicates a particular isotope of iron. The 26 is the atomic number (which is the same for all iron atoms), while the 56 is the mass number of the isotope. To determine the number of neutrons in this isotope, we subtract 26 from 56: $56 - 26 = 30$, so there are 30 neutrons in this atom.

EXAMPLE 5

How many protons and neutrons are in each atom?

1. ${}_{17}^{35}\text{Cl}$
2. ${}_{53}^{127}\text{I}$

Solution

1. In ${}_{17}^{35}\text{Cl}$, there are 17 protons, and $35 - 17 = 18$ neutrons in each nucleus.
2. In ${}_{53}^{127}\text{I}$, there are 53 protons, and $127 - 53 = 74$ neutrons in each nucleus.

SKILL-BUILDING EXERCISE

How many protons and neutrons are in each atom?

1. ${}_{79}^{197}\text{Au}$
2. ${}_{11}^{23}\text{Na}$

It is not absolutely necessary to indicate the atomic number as a subscript because each element has its own unique atomic number. Many isotopes are indicated with a superscript only, such as ${}^{13}\text{C}$ or ${}^{235}\text{U}$. You may also see isotopes represented in print as, for example, carbon-13 or uranium-235.

CONCEPT REVIEW EXERCISES

1. Why is the atomic number so important to the identity of an atom?
2. What is the relationship between the number of protons and the number of electrons in an atom?
3. How do isotopes of an element differ from each other?
4. What is the mass number of an element?

ANSWERS

1. The atomic number defines the identity of an element.
2. In an electrically neutral atom, the number of protons equals the number of electrons.
3. Isotopes have different numbers of neutrons in their nuclei.
4. The mass number is the sum of the numbers of protons and neutrons in the nucleus of an atom.

KEY TAKEAWAYS

- Elements can be identified by their atomic number and mass number.
- Isotopes are atoms of the same element that have different masses.

EXERCISES

- How many protons are in the nucleus of each element?
 - radon
 - tungsten
 - chromium
 - beryllium
- How many protons are in the nucleus of each element?
 - sulfur
 - uranium
 - calcium
 - lithium
- What are the atomic numbers of the elements in Exercise 1?
- What are the atomic numbers of the elements in Exercise 2?
- How many electrons are in neutral atoms of the elements in Exercise 1?
- How many electrons are in neutral atoms of the elements in Exercise 2?
- Complete the following table.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
80	120		
			${}_{26}^{55}\text{Fe}$
	2	hydrogen	

- Complete the following table.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
			${}_{2}^3\text{He}$
95	153		
	21	potassium	

- State the number of protons, neutrons, and electrons in neutral atoms of each isotope.

- a. ^{131}I
 - b. ^{40}K
 - c. ^{201}Hg
 - d. ^{19}F
10. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.
- a. ^3H
 - b. ^{133}Cs
 - c. ^{56}Fe
 - d. ^{207}Pb
11. What is the mass number of a gallium atom that has 38 neutrons in it?
12. What is the mass number of a uranium atom that has 143 neutrons in it?
13. Complete each sentence.
- a. ^{48}Ti has _____ neutrons.
 - b. ^{40}Ar has _____ neutrons.
 - c. ^3H has _____ neutrons.
14. Complete each sentence.
- a. ^{18}O has _____ neutrons.
 - b. ^{60}Ni has _____ neutrons.
 - c. ^{127}I has _____ neutrons.

ANSWERS

1. a. 86
b. 74
c. 24
d. 4
3. 86, 74, 24, and 4
5. 86, 74, 24, and 4

7.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
80	120	mercury	${}^{200}_{80}\text{Hg}$
26	29	iron	${}^{55}_{26}\text{Fe}$
1	2	hydrogen	${}^3_1\text{H}$

9. a. protons: 53; neutrons: 78; electrons: 53
b. protons: 19; neutrons: 21; electrons: 19
c. protons: 80; neutrons: 121; electrons: 80
d. protons: 9; neutrons: 10; electrons: 9
11. 69
13. a. 26
b. 22
c. 2

2.5 Atomic Masses

LEARNING OBJECTIVE

1. Define atomic mass and atomic mass unit.

Even though atoms are very tiny pieces of matter, they have mass. Their masses are so small, however, that chemists often use a unit other than grams to express them—the atomic mass unit.

The **atomic mass unit**¹¹ (abbreviated u, although amu is also used) is defined as 1/12 of the mass of a ¹²C atom:

$$1 \text{ u} = \frac{1}{12} \text{ the mass of } ^{12}\text{C atom}$$

It is equal to 1.661×10^{-24} g.

Masses of other atoms are expressed with respect to the atomic mass unit. For example, the mass of an atom of ¹H is 1.008 u, the mass of an atom of ¹⁶O is 15.995 u, and the mass of an atom of ³²S is 31.97 u. Note, however, that these masses are for particular isotopes of each element. Because most elements exist in nature as a mixture of isotopes, any sample of an element will actually be a mixture of atoms having slightly different masses (because neutrons have a significant effect on an atom's mass). How, then, do we describe the mass of a given element? By calculating an average of an element's atomic masses, weighted by the natural abundance of each isotope, we obtain a weighted average mass called the **atomic mass**¹² (also commonly referred to as the *atomic weight*) of an element.

For example, boron exists as a mixture that is 19.9% ¹⁰B and 80.1% ¹¹B. The atomic mass of boron would be calculated as $(0.199 \times 10.0 \text{ u}) + (0.801 \times 11.0 \text{ u}) = 10.8 \text{ u}$. Similar average atomic masses can be calculated for other elements. Carbon exists on Earth as about 99% ¹²C and about 1% ¹³C, so the weighted average mass of carbon atoms is 12.01 u.

The table in [Chapter 21 "Appendix: Periodic Table of the Elements"](#) also lists the atomic masses of the elements.

11. One-twelfth the mass of a ¹²C atom.
12. A weighted average of the masses of all the element's naturally occurring isotopes.

EXAMPLE 6

What is the average mass of a carbon atom in grams?

Solution

This is a simple one-step conversion, similar to conversions we did in Chapter 1 "Chemistry, Matter, and Measurement". We use the fact that $1 \text{ u} = 1.661 \times 10^{-24} \text{ g}$:

$$12.01 \cancel{\text{ u}} \times \frac{1.661 \times 10^{-24} \text{ g}}{1 \cancel{\text{ u}}} = 1.995 \times 10^{-23} \text{ g}$$

This is an extremely small mass, which illustrates just how small individual atoms are.

SKILL-BUILDING EXERCISE

1. What is the average mass of a tin atom in grams? The atomic mass of tin is 118.71 u.

CONCEPT REVIEW EXERCISES

1. Define atomic mass. Why is it considered a weighted average?
2. What is an atomic mass unit?

ANSWERS

1. The atomic mass is an average of an element's atomic masses, weighted by the natural abundance of each isotope of that element. It is a weighted average because different isotopes have different masses.
2. An atomic mass unit is 1/12th of the mass of a ^{12}C atom.

KEY TAKEAWAY

- Atoms have a mass that is based largely on the number of protons and neutrons in their nucleus.

EXERCISES

1. What is the atomic mass of zinc in atomic mass units?
2. What is the atomic mass of barium in atomic mass units?
3. What is the average mass of a single magnesium atom in grams?
4. What is the average mass of a single calcium atom in grams?
5. What is the mass of 1.00×10^{24} aluminum atoms in grams?
6. What is the mass of 5.00×10^{23} carbon atoms in grams?
7. Which has more mass—1 tungsten atom or 11 oxygen atoms?
8. Which has more mass—1 magnesium atom or 6 helium atoms?

ANSWERS

1. 65.4 u
3. 4.04×10^{-23} g
5. 44.8 g
7. 1 tungsten atom

2.6 Arrangements of Electrons

LEARNING OBJECTIVE

1. Describe how electrons are grouped within atoms.

Although we have discussed the general arrangement of subatomic particles in atoms, we have said little about how electrons occupy the space about the nucleus. Do they move around the nucleus at random, or do they exist in some ordered arrangement?

The modern theory of electron behavior is called **quantum mechanics**¹³. It makes the following statements about electrons in atoms:

- Electrons in atoms can have only certain specific energies. We say that the energies of the electrons are **quantized**¹⁴.
- Electrons are organized according to their energies into sets called **shells**¹⁵. Generally the higher the energy of a shell, the farther it is (on average) from the nucleus. Shells do not have specific, fixed distances from the nucleus, but an electron in a higher-energy shell will spend more time farther from the nucleus than does an electron in a lower-energy shell.
- Shells are further divided into subsets of electrons called **subshells**¹⁶. The first shell has only one subshell, the second shell has two subshells, the third shell has three subshells, and so on. The subshells of each shell are labeled, in order, with the letters *s*, *p*, *d*, and *f*. Thus, the first shell has only an *s* subshell, the second shell has an *s* and a *p* subshell, the third shell has *s*, *p*, and *d* subshells, and so forth.
- Different subshells hold a different maximum number of electrons. Any *s* subshell can hold up to 2 electrons; *p*, 6; *d*, 10; and *f*, 14.

13. The modern theory of electron behavior.

14. Having a fixed value.

15. A grouping of electrons within an atom.

16. A grouping of electrons within a shell.

It is the arrangement of electrons into shells and subshells that most concerns us here, so we will focus on that.

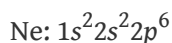
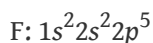
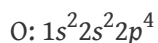
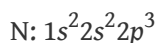
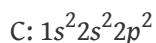
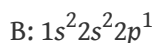
We use numbers to indicate which shell an electron is in. The first shell, closest to the nucleus and with the lowest-energy electrons, is shell 1. This first shell has only one subshell, which is labeled *s* and can hold a maximum of 2 electrons. We combine the shell and subshell labels when referring to the organization of electrons about a

nucleus and use a superscript to indicate how many electrons are in a subshell. Thus, because a hydrogen atom has its single electron in the s subshell of the first shell, we use $1s^1$ to describe the electronic structure of hydrogen. This structure is called an **electron configuration**¹⁷. Electron configurations are shorthand descriptions of the arrangements of electrons in atoms. The electron configuration of a hydrogen atom is spoken out loud as “one-ess-one.”

Helium atoms have 2 electrons. Both electrons fit into the $1s$ subshell because s subshells can hold up to 2 electrons; therefore, the electron configuration for helium atoms is $1s^2$ (spoken as “one-ess-two”).

The $1s$ subshell cannot hold 3 electrons (because an s subshell can hold a maximum of 2 electrons), so the electron configuration for a lithium atom cannot be $1s^3$. Two of the lithium electrons can fit into the $1s$ subshell, but the third electron must go into the second shell. The second shell has two subshells, s and p , which fill with electrons in that order. The $2s$ subshell holds a maximum of 2 electrons, and the $2p$ subshell holds a maximum of 6 electrons. Because lithium’s final electron goes into the $2s$ subshell, we write the electron configuration of a lithium atom as $1s^2 2s^1$.

The next largest atom, beryllium, has 4 electrons, so its electron configuration is $1s^2 2s^2$. Now that the $2s$ subshell is filled, electrons in larger atoms start filling the $2p$ subshell. Thus, the electron configurations for the next six atoms are as follows:



With neon, the $2p$ subshell is completely filled. Because the second shell has only two subshells, atoms with more electrons now must begin the third shell. The third shell has three subshells, labeled s , p , and d . The d subshell can hold a maximum of 10 electrons. The first two subshells of the third shell are filled in order—for

17. A shorthand description of the arrangement of electrons in an atom.

example, the electron configuration of aluminum, with 13 electrons, is $1s^2 2s^2 2p^6 3s^2 3p^1$. However, a curious thing happens after the $3p$ subshell is filled: the $4s$ subshell begins to fill before the $3d$ subshell does. In fact, the exact ordering of subshells becomes more complicated at this point (after argon, with its 18 electrons), so we will not consider the electron configurations of larger atoms.

A fourth subshell, the f subshell, is needed to complete the electron configurations for all elements. An f subshell can hold up to 14 electrons.

EXAMPLE 7

What is the electron configuration of a neutral phosphorus atom?

Solution

A neutral phosphorus atom has 15 electrons. Two electrons can go into the $1s$ subshell, 2 can go into the $2s$ subshell, and 6 can go into the $2p$ subshell. That leaves 5 electrons. Of those 5 electrons, 2 can go into the $3s$ subshell, and the remaining 3 electrons can go into the $3p$ subshell. Thus, the electron configuration of neutral phosphorus atoms is $1s^2 2s^2 2p^6 3s^2 3p^3$.

SKILL-BUILDING EXERCISE

1. What is the electron configuration of a neutral chlorine atom?

Chemistry results from interactions between the outermost shells of electrons on different atoms. Thus, it is convenient to separate electrons into two groups. **Valence shell electrons**¹⁸ (or, more simply, the *valence electrons*) are the electrons in the highest-numbered shell, or **valence shell**¹⁹, while **core electrons**²⁰ are the electrons in lower-numbered shells. We can see from the electron configuration of a carbon atom— $1s^2 2s^2 2p^2$ —that it has 4 valence electrons ($2s^2 2p^2$) and 2 core electrons ($1s^2$).

18. An electron in the highest-numbered shell of an atom.

19. The highest-numbered shell of an atom that contains electrons.

20. An electron in a lower-numbered shell of an atom.

EXAMPLE 8

From the electron configuration of neutral phosphorus atoms in Example 7, how many valence electrons and how many core electrons does a neutral phosphorus atom have?

Solution

The highest-numbered shell is the third shell, which has 2 electrons in the 3s subshell and 3 electrons in the 3p subshell. That gives a total of 5 electrons, so neutral phosphorus atoms have 5 valence electrons. The 10 remaining electrons, from the first and second shells, are core electrons.

SKILL-BUILDING EXERCISE

1. From the electron configuration of neutral chlorine atoms (see the Skill-Building Exercise following Example 7), how many valence electrons and how many core electrons does a neutral chlorine atom have?

CONCEPT REVIEW EXERCISES

1. How are electrons organized in atoms?
2. What information does an electron configuration convey?
3. What is the difference between core electrons and valence electrons?

ANSWERS

1. Electrons are organized into shells and subshells around nuclei.
2. The electron configuration states the arrangement of electrons in shells and subshells.
3. Valence electrons are in the highest-numbered shell; all other electrons are core electrons.

KEY TAKEAWAY

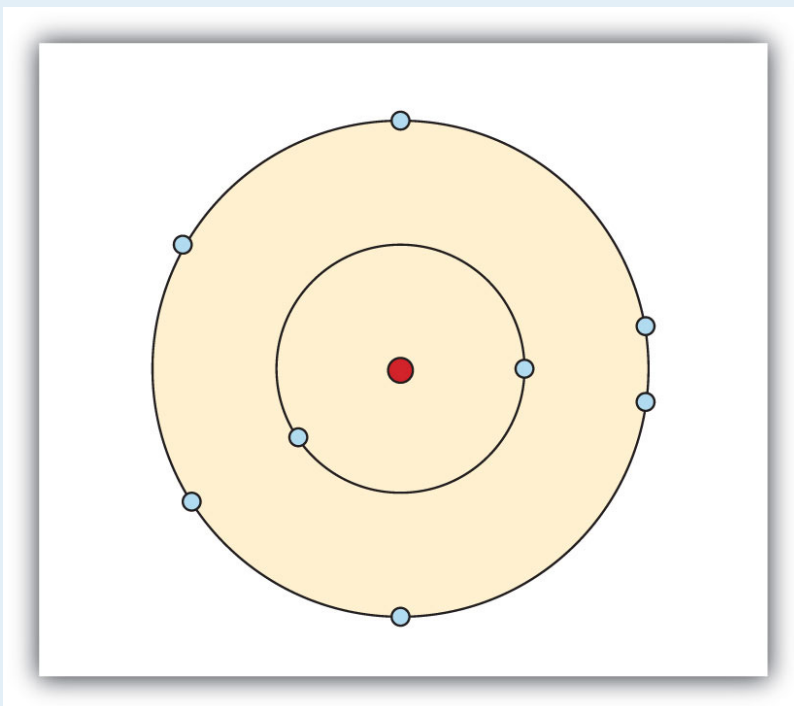
- Electrons are organized into shells and subshells about the nucleus of an atom.

EXERCISES

1. What is the maximum number of electrons that can fit in an s subshell? Does it matter what shell the s subshell is in?
2. What is the maximum number of electrons that can fit in a p subshell? Does it matter what shell the p subshell is in?
3. What is the maximum number of electrons that can fit in a d subshell? Does it matter what shell the d subshell is in?
4. What is the maximum number of electrons that can fit in an f subshell? Does it matter what shell the f subshell is in?
5. What is the electron configuration of a carbon atom?
6. What is the electron configuration of a sulfur atom?
7. What is the valence shell electron configuration of a calcium atom?
8. What is the valence shell electron configuration of a selenium atom?
9. What atom has the electron configuration $1s^2 2s^2 2p^5$?
10. What atom has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^3$?
11. Draw a representation of the electronic structure of an oxygen atom.
12. Draw a representation of the electronic structure of a phosphorus atom.
13. A potassium atom has ____ core electrons and ____ valence electrons.
14. A silicon atom has ____ core electrons and ____ valence electrons.

ANSWERS

1. 2; no
3. 10; no
5. $1s^2 2s^2 2p^2$
7. $4s^2$
9. fluorine



- 11.
13. 18; 1

2.7 The Periodic Table

LEARNING OBJECTIVES

1. Explain how elements are organized into the periodic table.
2. Describe how some characteristics of elements relate to their positions on the periodic table.

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. For example, chlorine, bromine, and iodine react with other elements (such as sodium) to make similar compounds. Likewise, lithium, sodium, and potassium react with other elements (such as oxygen) to make similar compounds. Why is this so?

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the **periodic table**²¹.

Note

Mendeleev had to list some elements out of the order of their atomic masses to group them with other elements that had similar properties.

The periodic table is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern version is shown in [Figure 2.7 "A Modern Periodic Table"](#). Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.

21. A chart of elements that groups the elements by some of their properties.

Figure 2.7 A Modern Periodic Table

B = Solids Hg = Liquids Kr = Gases Pm = Not found in nature

1 H 1.00794																	18 He 4.002602
3 Li 6.941	4 Be 9.012182											5 B 10.811	6 C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797
11 Na 22.989770	12 Mg 24.3050	3	4	5	6	7	8	9	10	11	12	13 Al 26.981538	14 Si 28.0855	15 P 30.973761	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6534	29 Cu 63.545	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.504	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29
55 Cs 132.90545	56 Ba 137.327	71 Lu 174.967	72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.96655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.98038	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	103 Lr (262)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Ds (269)	111 Rg (272)	112 Cn (277)	113 Uut (277)	114 Uuq (277)	115 Uup (277)	116 Uuh (277)		118 Uuo (277)
57 La 138.9055	58 Ce 140.116	59 Pr 140.50765	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.92534	66 Dy 162.50	67 Ho 164.93032	68 Er 167.26	69 Tm 168.93421	70 Yb 173.04				
89 Ac 232.0381	90 Th 232.0381	91 Pa 231.03688	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)				

A modern periodic table lists elements left to right by atomic number.

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called **groups (or families)**²². As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements).

Note

The word *halogen* comes from the Greek for “salt maker” because these elements combine with other elements to form a group of compounds called salts.

22. A column of elements on the periodic table.

To Your Health: Radon

Radon is an invisible, odorless noble gas that is slowly released from the ground, particularly from rocks and soils whose uranium content is high. Because it is a noble gas, radon is not chemically reactive. Unfortunately, it is radioactive, and increased exposure to it has been correlated with an increased lung cancer risk.

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building's inhabitants. Fortunately, specialized ventilation minimizes the amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

After smoking, radon is thought to be the second-biggest *preventable* cause of lung cancer in the United States. The American Cancer Society estimates that 10% of all lung cancers are related to radon exposure. There is uncertainty regarding what levels of exposure cause cancer, as well as what the exact causal agent might be (either radon or one of its breakdown products, many of which are also radioactive and, unlike radon, not gases). The US Environmental Protection Agency recommends testing every floor below the third floor for radon levels to guard against long-term health effects.

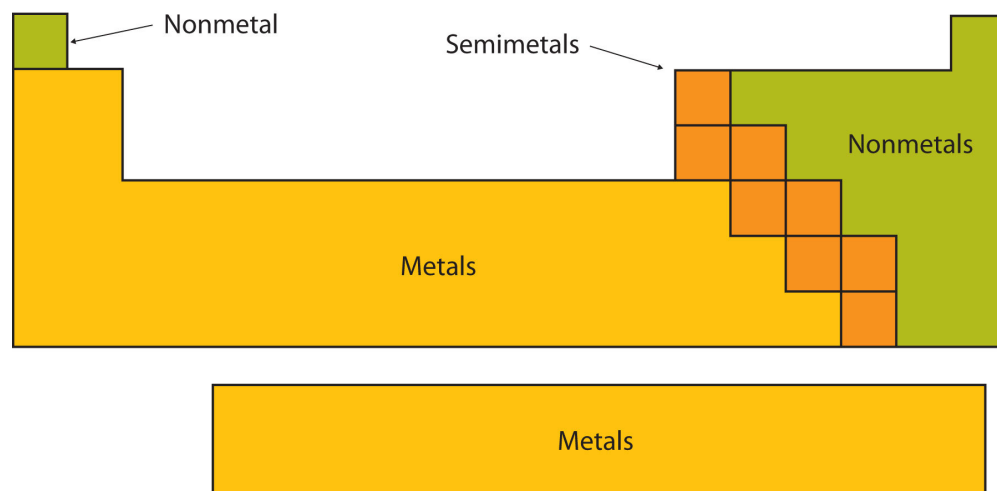
Each row of elements on the periodic table is called a **period**²³. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a semimetal, as shown in [Figure 2.8 "Types of Elements"](#). A **metal**²⁴ is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal**²⁵ is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in [Figure 2.8 "Types of Elements"](#), metals occupy the left three-fourths of the

23. A row of elements on the periodic table.
24. An element that is shiny, typically silvery in color, an excellent conductor of heat and electricity, malleable, and ductile.
25. An element that is typically dull, not a good conductor of heat and electricity, and brittle.

periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **semimetals (or metalloids)**²⁶. Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

Figure 2.8 Types of Elements



Elements are either metals, nonmetals, or semimetals. Each group is located in a different part of the periodic table.

Another way to categorize the elements of the periodic table is shown in [Figure 2.9 "Special Names for Sections of the Periodic Table"](#). The first two columns on the left and the last six columns on the right are called the **main group elements**²⁷. The ten-column block between these columns contains the **transition metals**²⁸. The two rows beneath the main body of the periodic table contain the **inner transition metals**²⁹. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

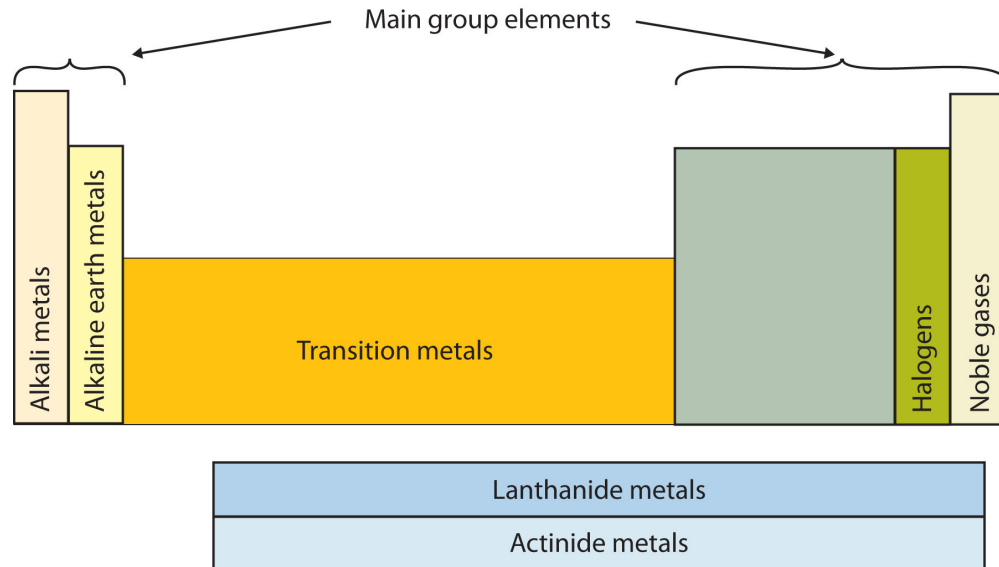
26. An element whose properties are intermediate between metals and nonmetals.

27. An element in the first two or the last six columns on the periodic table.

28. An element between the main group elements on the periodic table.

29. An element in the two rows beneath the main body on the periodic table. Such metals are also called the lanthanide and actinide elements.

Figure 2.9 Special Names for Sections of the Periodic Table



Some sections of the periodic table have special names. The elements lithium, sodium, potassium, rubidium, cesium, and francium are collectively known as alkali metals.

To Your Health: Transition Metals in the Body

According to [Table 2.2 "Elemental Composition of a Human Body"](#), most of the elemental composition of the human body consists of main group elements. The first element appearing on the list that is not a main group element is iron, at 0.006 percentage by mass. Because iron has relatively massive atoms, it would appear even lower on a list organized in terms of percent by *atoms* rather than percent by mass.

Iron is a transition metal. Transition metals have interesting chemical properties, partially because some of their electrons are in *d* subshells. (For more information about electron shells, see [Section 2.6 "Arrangements of Electrons"](#).) The chemistry of iron makes it a key component in the proper functioning of red blood cells.

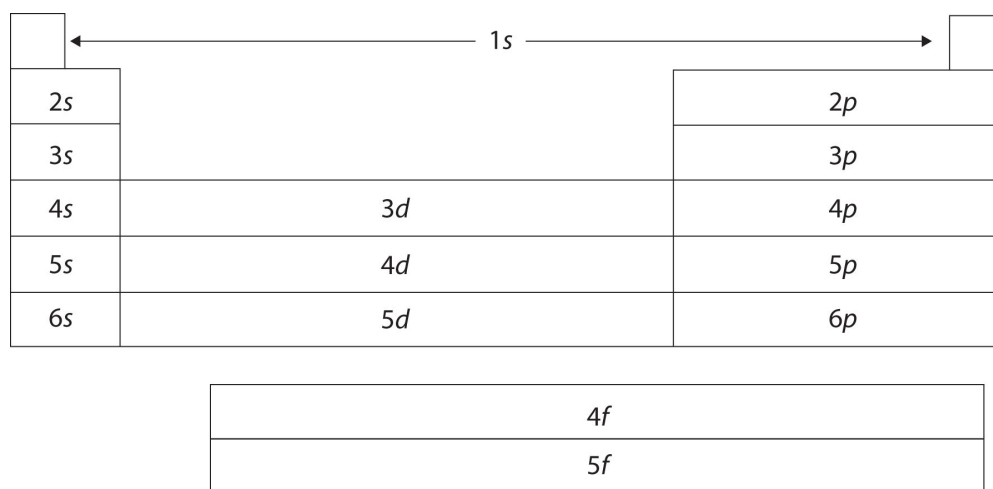
Red blood cells are cells that transport oxygen from the lungs to cells of the body and then transport carbon dioxide from the cells to the lungs. Without red blood cells, animal respiration as we know it would not exist. The critical part of the red blood cell is a protein called *hemoglobin*. Hemoglobin combines with oxygen and carbon dioxide, transporting these gases from one location to another in the body. Hemoglobin is a relatively large molecule, with a mass of about 65,000 u.

The crucial atom in the hemoglobin protein is iron. Each hemoglobin molecule has four iron atoms, which act as binding sites for oxygen. It is the presence of this particular transition metal in your red blood cells that allows you to use the oxygen you inhale.

Other transition metals have important functions in the body, despite being present in low amounts. Zinc is needed for the body's immune system to function properly, as well as for protein synthesis and tissue and cell growth. Copper is also needed for several proteins to function properly in the body. Manganese is needed for the body to metabolize oxygen properly. Cobalt is a necessary component of vitamin B-12, a vital nutrient. (For more information about proteins and vitamins, see [Chapter 18 "Amino Acids, Proteins, and Enzymes"](#).) These last three metals are not listed explicitly in [Table 2.2 "Elemental Composition of a Human Body"](#), so they are present in the body in very small quantities. However, even these small quantities are required for the body to function properly.

The periodic table is organized on the basis of similarities in elemental properties, but what explains these similarities? It turns out that the shape of the periodic table reflects the filling of subshells with electrons, as shown in [Figure 2.10 "The Shape of the Periodic Table"](#). Starting with the first period and going from left to right, the table reproduces the order of filling of the electron subshells in atoms. Furthermore, elements in the same column share the same valence shell electron configuration. For example, all elements in the first column have a single s electron in their valence shells, so their electron configurations can be described as ns^1 (where n represents the shell number). This last observation is crucial. Chemistry is largely the result of interactions between the valence electrons of different atoms. Thus, atoms that have the same valence shell electron configuration will have similar chemistry.

Figure 2.10 *The Shape of the Periodic Table*



The shape of the periodic table reflects the order in which electron shells and subshells fill with electrons.

EXAMPLE 9

Using the variable n to represent the number of the valence electron shell, write the valence shell electron configuration for each group.

1. the alkaline earth metals
2. the column of elements headed by carbon

Solution

1. The alkaline earth metals are in the second column of the periodic table. This column corresponds to the s subshell being filled with 2 electrons. Therefore, the valence shell electron configuration is ns^2 .
2. The electron configuration of carbon is $1s^2 2s^2 2p^2$. Its valence shell electron configuration is $2s^2 2p^2$. Every element in the same column should have a similar valence shell electron configuration, which we can represent as $ns^2 np^2$.

SKILL-BUILDING EXERCISE

Using the variable n to represent the number of the valence electron shell, write the valence shell electron configuration for each group.

1. the halogens
2. the column of elements headed by oxygen

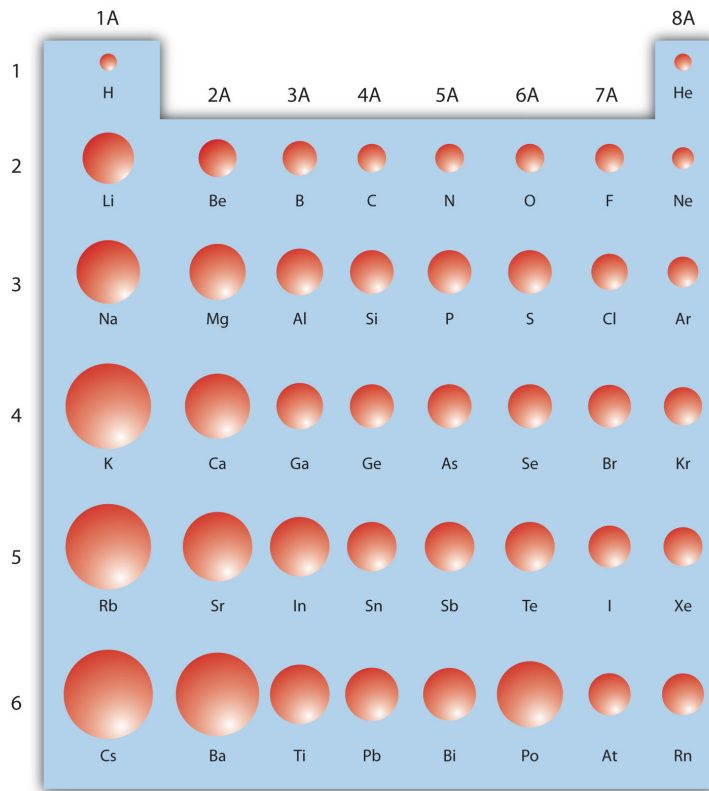
Atomic Radius

The periodic table is useful for understanding atomic properties that show periodic trends. One such property is the **atomic radius**³⁰ (Figure 2.11 "Trends on the Periodic Table"). As mentioned earlier, the higher the shell number, the farther from the nucleus the electrons in that shell are likely to be. In other words, the size of an atom is generally determined by the number of the valence electron shell. Therefore, as we go down a column on the periodic table, the atomic radius increases. As we go *across* a period on the periodic table, however, electrons are being added to the *same* valence shell; meanwhile, more protons are being added to the nucleus, so the positive charge of the nucleus is increasing. The increasing positive charge attracts the electrons more strongly, pulling them closer to the

30. The approximate size of an atom.

nucleus. Consequently, as we go across a period, the atomic radius decreases. These trends are seen clearly in [Figure 2.11 "Trends on the Periodic Table"](#).

Figure 2.11 Trends on the Periodic Table



The relative sizes of the atoms show several trends with regard to the structure of the periodic table. Atoms become larger going down a column and smaller going across a period.

EXAMPLE 10

Using the periodic table (rather than [Figure 2.11 "Trends on the Periodic Table"](#)), which atom is larger?

1. N or Bi
2. Mg or Cl

Solution

1. Because Bi is below N on the periodic table and has electrons in higher-numbered shells, we expect that Bi atoms are larger than N atoms.
2. Both Mg and Cl are in period 3 of the periodic table, but Cl lies farther to the right. Therefore we expect Mg atoms to be larger than Cl atoms.

SKILL-BUILDING EXERCISE

Using the periodic table (rather than [Figure 2.11 "Trends on the Periodic Table"](#)), which atom is larger?

1. Li or F
2. Na or K

Career Focus: Clinical Chemist

Clinical chemistry is the area of chemistry concerned with the analysis of body fluids to determine the health status of the human body. Clinical chemists measure a variety of substances, ranging from simple elements such as sodium and potassium to complex molecules such as proteins and enzymes, in blood, urine, and other body fluids. The absence or presence, or abnormally low or high amounts, of a substance can be a sign of some disease or an indication of health. Many clinical chemists use sophisticated equipment and complex chemical reactions in their work, so they not only need to understand basic chemistry, but also be familiar with special instrumentation and how to interpret test results.

CONCEPT REVIEW EXERCISES

1. How are the elements organized into the periodic table?
2. Looking at the periodic table, where do the following elements appear?
 - a. the metals
 - b. the nonmetals
 - c. the halogens
 - d. the transition metals
3. Describe the trends in atomic radii as related to an element's position on the periodic table.

ANSWERS

1. Elements are organized by atomic number.
2.
 - a. the left three-quarters of the periodic table
 - b. the right quarter of the periodic table
 - c. the next-to-last column of the periodic table
 - d. the middle section of the periodic table
3. As you go across the periodic table, atomic radii decrease; as you go down the periodic table, atomic radii increase.

KEY TAKEAWAYS

- The chemical elements are arranged in a chart called the periodic table.
- Some characteristics of the elements are related to their position on the periodic table.

EXERCISES

1. Which elements have chemical properties similar to those of magnesium?
 - a. sodium
 - b. fluorine
 - c. calcium
 - d. barium
 - e. selenium
2. Which elements have chemical properties similar to those of lithium?
 - a. sodium
 - b. calcium
 - c. beryllium
 - d. barium
 - e. potassium
3. Which elements have chemical properties similar to those of chlorine?
 - a. sodium
 - b. fluorine
 - c. calcium
 - d. iodine
 - e. sulfur
4. Which elements have chemical properties similar to those of carbon?
 - a. silicon
 - b. oxygen
 - c. germanium
 - d. barium
 - e. argon
5. Which elements are alkali metals?
 - a. sodium
 - b. magnesium
 - c. aluminum
 - d. potassium
 - e. calcium
6. Which elements are alkaline earth metals?
 - a. sodium
 - b. magnesium
 - c. aluminum

- d. potassium
 - e. calcium
7. Which elements are halogens?
- a. oxygen
 - b. fluorine
 - c. chlorine
 - d. sulfur
 - e. carbon
8. Which elements are noble gases?
- a. helium
 - b. hydrogen
 - c. oxygen
 - d. neon
 - e. chlorine
9. Which pairs of elements are located in the same period?
- a. H and Li
 - b. H and He
 - c. Na and S
 - d. Na and Rb
10. Which pairs of elements are located in the same period?
- a. V and Nb
 - b. K and Br
 - c. Na and P
 - d. Li and Mg
11. In each pair of atoms, which atom has the greater atomic radius?
- a. H and Li
 - b. N and P
 - c. Cl and Ar
 - d. Al and Cl
12. In each pair of atoms, which atom has the greater atomic radius?
- a. H and He
 - b. N and F
 - c. Cl and Br
 - d. Al and B

13. Scandium is a (metal, nonmetal, semimetal) and is a member of the (main group elements, transition metals).
14. Silicon is a (metal, nonmetal, semimetal) and is a member of the (main group elements, transition metals).

ANSWERS

1.
 - a. no
 - b. no
 - c. yes
 - d. yes
 - e. no

3.
 - a. no
 - b. yes
 - c. no
 - d. yes
 - e. no

5.
 - a. yes
 - b. no
 - c. no
 - d. yes
 - e. no

7.
 - a. no
 - b. yes
 - c. yes
 - d. no
 - e. no

9.
 - a. no
 - b. yes
 - c. yes
 - d. no

11.
 - a. Li
 - b. P
 - c. Cl
 - d. Al

13. metal; transition metals

2.8 End-of-Chapter Material

Chapter Summary

To ensure that you understand the material in this chapter, you should review the meanings of the following bold terms and ask yourself how they relate to the topics in the chapter.

An element is a substance that cannot be broken down into simpler chemical substances. Only about 90 naturally occurring elements are known. They have varying abundances on Earth and in the body. Each element has a one- or two-letter **chemical symbol**.

The **modern atomic theory** states that the smallest piece of an element is an **atom**. Individual atoms are extremely small, on the order of 10^{-10} m across. Most elements exist in pure form as individual atoms, but some exist as **diatomic molecules**. Atoms themselves are composed of subatomic particles. The **electron** is a tiny subatomic particle with a negative charge. The **proton** has a positive charge and, while small, is much larger than the electron. The **neutron** is also much larger than an electron but has no electrical charge.

Protons, neutrons, and electrons have a specific arrangement in an atom. The protons and neutrons are found in the center of the atom, grouped together into a **nucleus**. The electrons are found in fuzzy clouds around the nucleus.

Each element has a characteristic number of protons in its nucleus. This number of protons is the **atomic number** of the element. An element may have different numbers of neutrons in the nuclei of its atoms; such atoms are referred to as **isotopes**. Two isotopes of hydrogen are deuterium, with a proton and a neutron in its nucleus, and tritium, with a proton and two neutrons in its nucleus. The sum of the numbers of protons and neutrons in a nucleus is called the **mass number** and is used to distinguish isotopes from each other.

Masses of individual atoms are measured in **atomic mass units**. An atomic mass unit is equal to 1/12th of the mass of a single carbon-12 atom. Because different isotopes of an element have different masses, the **atomic mass** of an element is a weighted average of the masses of all the element's naturally occurring isotopes.

The modern theory of electron behavior is called **quantum mechanics**. According to this theory, electrons in atoms can only have specific, or **quantized**, energies. Electrons are grouped into general regions called **shells**, and within these into more specific regions called **subshells**. There are four types of subshells, and each type can hold up to a maximum number of electrons. The distribution of electrons into shells and subshells is the **electron configuration** of an atom. Chemistry typically occurs because of interactions between the electrons of the outermost shell of different atoms, called the valence shell electrons. Electrons in inner shells are called core electrons.

Elements are grouped together by similar chemical properties into a chart called the **periodic table**. Vertical columns of elements are called **groups** or **families**. Some of the groups of elements have names, like the alkali metals, the alkaline earth metals, the halogens, and the noble gases. A horizontal row of elements is called a **period**. Periods and groups have differing numbers of elements in them. The periodic table separates elements into **metals**, **nonmetals**, and **semimetals**. The periodic table is also separated into **main group elements**, **transition metals**, **lanthanide elements**, and **actinide elements**. The lanthanide and actinide elements are also referred to as **inner transition metal elements**. The shape of the periodic table reflects the sequential filling of shells and subshells in atoms.

The periodic table helps us understand trends in some of the properties of atoms. One such property is the **atomic radius** of atoms. From top to bottom of the periodic table, atoms get bigger because electrons are occupying larger and bigger shells. From left to right across the periodic table, electrons are filling the same shell but are being attracted by an increasing positive charge from the nucleus, and thus the atoms get smaller.

ADDITIONAL EXERCISES

1. If the atomic radius of sodium atoms is 1.86×10^{-10} m, how many sodium atoms are needed to make a line that is 1.00 cm in length?
2. If the atomic radius of osmium atoms is 1.34×10^{-10} m, how many osmium atoms are needed to make a line that is 5.85 m in length?
3. What might be the electron configuration of K^+ , an atom that has lost an electron?
4. What might be the electron configuration of Cl^- , an atom that has gained an additional electron?
5. The electron configuration of the Ti atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$. What is the valence shell electron configuration of Ti?
6. The electron configuration of the Ge atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$. What is the valence shell electron configuration of Ge?
7. What is the mass of an electron in atomic mass units?
8. In a footnote in this chapter, an alpha particle was defined as a particle with 2 protons and 2 neutrons. What is the mass, in grams, of an alpha particle? (Hint: what element does an alpha particle resemble?)
9. A sample of the nonexistent element mythium consists of 25.59% of an isotope with mass number 580, 32.74% of an isotope with mass number 582, and 41.67% of an isotope with mass number 581. What is the atomic mass of mythium?
10. Because the distribution of isotopes is different on different planets in the solar system, the average atomic mass of any element differs from planet to planet. Assume that on Mercury, a rather hot planet, there is more deuterium left in the atmosphere than on Earth, so that 92.55% of the hydrogen on Mercury is 1H , while the remainder is 2H . What is the atomic mass of hydrogen on Mercury?
11. The compound that sodium makes with chlorine has sodium and chlorine atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with sodium.
12. The compound that magnesium makes with oxygen has magnesium to oxygen atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with magnesium.

ANSWERS

1. 5.38×10^7 atoms
3. $1s^2 2s^2 2p^6 3s^2 3p^6$
5. $4s^2$
7. 0.000545 u
9. 581.1 u
11. potassium and bromine (answers will vary)