1. Atomic Theory. Composition of atom
2. The periodic law. Metals and nonmetals
3. Electronic structure of atoms
   - The quantum mechanical atom
   - Electron configuration of atoms; principal levels and sublevels
4. Macro and microelements and their biological role

Be prepared to answer!
• Characteristics of atoms
• Specification of macro and microelements
• Electronic configuration of atoms
• Biological role of biogenic elements

ATOMIC THEORY. COMPOSITION OF ATOM

Our present model of atom is based on the ideas of John Dalton. Some of Dalton’ ideas have not stood the test of time. The major features of modern atomic theory are summarized in the following statement:

1. An element is composed of extremely small? Invisible particles called atoms
2. Atoms of a particular elements have a unique set of chemical properties which distinguish them from atoms of all other elements
3. In the course of an ordinary chemical reaction, atoms do not disappear or change into atoms of another elements
4. Compounds are formed when atoms of two or more elements combine with each other
5. In a particular compound, the ratio of different kinds of atoms is fixed. Ordinary, this ratio can be expressed in terms of simple whole numbers

COMPOSITION OF THE ATOM

From his experiments Rutherford proposed a model of the atom which we use today. According to the model, atoms have a very small center, called nucleus. The nucleus is positively charged and contains almost all of the mass of an atom. Outside the nucleus is a large region which defines the volume of the atom. Within that region, there are enough electrons to make the atom electrically neutral. Rutherford showed that the diameter of an atomic nucleus is about 10^{-14} m. In contrast, the diameter of an average atom is about 10^{-10} m.

Components of the atom

<table>
<thead>
<tr>
<th>Particle</th>
<th>Charge</th>
<th>Location</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron</td>
<td>-1</td>
<td>Outside nucleus</td>
</tr>
<tr>
<td>Proton</td>
<td>+1</td>
<td>In the nucleus</td>
</tr>
<tr>
<td>Neutron</td>
<td>0</td>
<td>In the nucleus</td>
</tr>
</tbody>
</table>

The number of protons in an atom of the element is referred to as its atomic number

Atomic number = Number of protons in nucleus = Number of electrons
Most elements consist of atoms of different masses, called isotopes. The isotopes of a given element contain the same number of protons (and also the same number of electrons) because they are atoms of the same element. They differ in mass because they contain different numbers of neutrons in their nuclei.

**Isotopes are atoms of the same element with different masses; they are atoms containing the same number of protons but different numbers of neutrons.**

For example, there are three distinct kinds of hydrogen atoms, commonly called hydrogen, deuterium, and tritium. (This is the only element for which we give each isotope a different name.) Each contains one proton in the atomic nucleus. The predominant form of hydrogen contains no neutrons, but each deuterium atom contains one neutron and each tritium atom contains two neutrons in its nucleus.

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Nuclide Symbol</th>
<th>Mass (amu)</th>
<th>Atomic Abundance in Nature</th>
<th>No. of Protons</th>
<th>No. of Neutrons</th>
<th>No. of Electrons (in neutral atoms)</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>H</td>
<td>1H</td>
<td>1.007825</td>
<td>99.985%</td>
<td>1</td>
<td>0</td>
<td>1</td>
</tr>
<tr>
<td>deuterium</td>
<td>D</td>
<td>2H</td>
<td>2.01400</td>
<td>0.015%</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>tritium*</td>
<td>T</td>
<td>3H</td>
<td>3.01635</td>
<td>0.000%</td>
<td>1</td>
<td>2</td>
<td>1</td>
</tr>
</tbody>
</table>

All three forms of hydrogen display very similar chemical properties. The mass number of an atom is the sum of the number of protons and the number of neutrons in its nucleus; that is

**Mass number = number of protons 1 number of neutrons = atomic number 1 neutron number**

A mass number is a count of the number of protons plus neutrons present, so it must be a whole number. Because the masses of the proton and the neutron are both about 1 amu, the mass number is approximately equal to the actual mass of the isotope (which is not a whole number).

The mass number for normal hydrogen atoms is 1; for deuterium, 2; and for tritium, 3. The composition of a nucleus is indicated by its nuclide symbol. This consists of the symbol for the element (E), with the atomic number (Z) written as a subscript at the lower left and the mass number (A) as a superscript at the upper left, \(^A\)\(^Z\)E. By this system, the three isotopes of hydrogen are designated as \(^1\)H, \(^2\)H, \(^3\)H.
**Example. DETERMINATION OF ATOMIC MAKEUP**

Determine the number of protons, neutrons, and electrons in each of the following species. Are the members within each pair isotopes?

(a) \(^{35}\text{Cl}\) and \(^{37}\text{Cl}\)  
(b) \(^{63}\text{Cu}\) and \(^{65}\text{Cu}\)

**Plan**
Knowing that the number at the bottom left of the nuclide symbol is the atomic number or number of protons, we can verify the identity of the element in addition to knowing the number or protons per nuclide. From the mass number at the top left, we know the number of protons plus neutrons. The number of protons (atomic number) minus the number of electrons must equal the charge, if any, shown at the top right. From these data one can determine if two nuclides have the same number of protons and are therefore the same element. If they are the same element, they will be isotopes only if their mass numbers differ.

**Solution**

(a) For \(^{35}\text{Cl}\)

Atomic number = 17. There are therefore 17 protons per nucleus.

Mass number = 35. There are therefore 35 protons plus neutrons or, because we know that there are 17 protons, there are 18 neutrons.

Because no charge is indicated, there must be equal numbers of protons and electrons, or 17 electrons.

For \(^{37}\text{Cl}\): There are 17 protons, 20 neutrons, and 17 electrons per atom.

These are isotopes of the same element. Both have 17 protons, but they differ in their numbers of neutrons: one has 18 neutrons and the other has 20.

(b) For \(^{63}\text{Cu}\)

Atomic number = 29. There are 29 protons per nucleus.

Mass number = 63. There are 29 protons plus 34 neutrons.

Because no charge is indicated, there must be equal numbers of protons and electrons, or 29 electrons.

For \(^{65}\text{Cu}\): There are 29 protons, 36 neutrons, and 29 electrons per atom.

These are isotopes. Both have 29 protons, but they differ in their numbers of neutrons: one isotope has 34 neutrons and the other has 36.

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**THE PERIODIC LAW. METALS AND NONMETALS**

In 1869, the Russian chemist Dmitri Mendeleev (1834–1907) and the German chemist Lothar Meyer (1830–1895) independently published arrangements of known elements that are much like the periodic table in use today. Mendeleev’s classification was based primarily on chemical properties of the elements, whereas Meyer’s classification was based largely on physical properties. The tabulations were surprisingly similar. Both emphasized the periodicity, or regular periodic repetition, of properties with increasing atomic weight.

Mendeleev arranged the known elements in order of increasing atomic weight in successive sequences so that elements with similar chemical properties fell into the same column. He noted that both physical and chemical properties of the elements vary in a periodic fashion with atomic weight. His periodic table of 1872 contained the 62 known elements. Mendeleev placed H, Li, Na, and K in his table as “Gruppe I.” These were known to combine with F, Cl, Br, and I of “Gruppe VII” to produce compounds that have similar formulas such as HF, LiCl, NaCl, and KI. All these compounds dissolve in water to produce solutions that conduct electricity. The “Gruppe II” elements were known to form compounds such as BeCl₂,
MgBr₂, and CaCl₂, as well as compounds with O and S from “Gruppe VI” such as MgO, CaO, MgS, and CaS. These and other chemical properties led him to devise a table in which the elements were arranged by increasing atomic weights and grouped into vertical families. In most areas of human endeavor progress is slow and faltering. Occasionally, however, an individual develops concepts and techniques that clarify confused situations. Mendeleev was such an individual. One of the brilliant successes of his periodic table was that it provided for elements that were unknown at the time. When he encountered “missing” elements, Mendeleev left blank spaces. Some appreciation of his genius in constructing the table as he did can be gained by comparing the predicted (1871) and observed properties of germanium, which was not discovered until 1886. Mendeleev called the undiscovered element eka-silicon because it fell below silicon in his table. He was familiar with the properties of germanium’s neighboring elements. They served as the basis for his predictions of properties of germanium. Some modern values for properties of germanium differ significantly from those reported in 1886. But many of the values on which Mendeleev based his predictions were also inaccurate. Because Mendeleev’s arrangement of the elements was based on increasing atomic weights, several elements would have been out of place in his table. Mendeleev put the controversial elements (Te and I, Co and Ni) in locations consistent with their properties, however. He thought the apparent reversal of atomic weights was due to inaccurate values for those weights. Careful redetermination showed that the values were correct. Explanation of the locations of these “out-of-place” elements had to await the development of the concept of atomic number, approximately 50 years after Mendeleev’s work. The atomic number of an element is the number of protons in the nucleus of its atoms. (It is also the number of electrons in a neutral atom of an element.) This quantity is fundamental to the identity of each element. Elements are now arranged in the periodic table in order of increasing atomic number. With the development of this concept, the periodic law attained its present form:

The properties of the elements are periodic functions of their atomic numbers.

The periodic law tells us that if we arrange the elements in order of increasing atomic number, we periodically encounter elements that have similar chemical and physical properties. The presently used “long form” of the periodic table is such an arrangement. The vertical columns are referred to as groups or families, and the horizontal rows are called periods. Elements in a group have similar chemical and physical properties, and those within a period have properties that change progressively across the table. Several groups of elements have common names that are used so frequently they should be learned. The Group 1A elements, except H, are referred to as alkali metals, and the Group 2A elements are called the alkaline earth metals.
The Periodic Table shows how we classify the known elements as metals (shown in blue), nonmetals (tan), and metalloids (green). The elements to the left of those touching the heavy stairstep line are metals (except hydrogen), and those to the right are nonmetals. Such a classification is somewhat arbitrary, and several elements do not fit neatly into either class. Most elements adjacent to the heavy line are often called metalloids (or semimetals), because they are metallic (or nonmetallic) only to a limited degree.

Cesium, atomic number 55, is the most active naturally occurring metal. Francium and radium are radioactive and do not occur in nature in appreciable amounts. Noble gases seldom bond with other elements. They are unreactive, monatomic gases. The most active nonmetal is fluorine, atomic number 9.

<table>
<thead>
<tr>
<th>Metallic character increases from top to bottom and decreases from left to right with respect to position in the periodic table</th>
</tr>
</thead>
</table>

Some Physical Properties of Metals and Nonmetals

<table>
<thead>
<tr>
<th>Metals</th>
<th>Nonmetals</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. High electrical conductivity that decreases with increasing temperature&lt;br&gt;2. High thermal conductivity&lt;br&gt;3. Metallic gray or silver luster*&lt;br&gt;4. Almost all are solids&lt;br&gt;5. Malleable (can be hammered into sheets)&lt;br&gt;6. Ductile (can be drawn into wires)</td>
<td>1. Poor electrical conductivity (except carbon in the form of graphitic)&lt;br&gt;2. Good heat insulators (except carbon)&lt;br&gt;3. No metallic luster&lt;br&gt;4. Solids, liquids, or gases&lt;br&gt;5. Brittle in solid state&lt;br&gt;6. Nonductile</td>
</tr>
</tbody>
</table>

Some Chemical Properties of Metals and Nonmetals

<table>
<thead>
<tr>
<th>Metals</th>
<th>Nonmetals</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Outer shells contain few electrons—usually three or fewer&lt;br&gt;2. Form cations (positive ions) by losing electrons&lt;br&gt;3. Form ionic compounds with nonmetals&lt;br&gt;4. Solid state characterized by metallic bonding</td>
<td>1. Outer shells contain four or more electrons&lt;br&gt;2. Form anions (negative ions) by gaining electrons&lt;br&gt;3. Form ionic compounds with metals and molecular (covalent) other compounds with nonmetals&lt;br&gt;4. Covalently bonded molecules; noble gases are monatomic</td>
</tr>
</tbody>
</table>

**ELECTRONIC STRUCTURE OF ATOMS**

In 1913 Niels Bohr (1885–1962), a Danish physicist, provided an explanation for Balmer and Rydberg’s observations. He wrote equations that described the electron of a hydrogen atom as revolving around its nucleus in one of a discrete set of circular orbits. He included the assumption that the electronic energy is quantized; that is, only certain values of electronic energy are possible. This led him to suggest that electrons can only be in certain discrete orbits, and that they absorb or emit energy in discrete amounts as they move from one orbit to another. Each orbit thus corresponds to a definite energy level for the electron. When an electron is excited from a lower energy level to a higher one, it absorbs a definite (or quantized) amount of energy. When the electron falls back to the original energy level, it emits exactly the same amount of energy it absorbed in moving from the lower to the higher energy level.

The values of $n_1$ and $n_2$ in the Balmer–Rydberg equation identify the lower and higher levels, respectively, of these electronic transitions.
(a) The radii of the first four Bohr orbits for a hydrogen atom. The dot at the center represents the nuclear position. The radius of each orbit is proportional to $n^2$, so the orbits are more widely spaced as the $n$ value increases. These four radii are in the ratio $1 : 4 : 9 : 16$.

(b) Relative values for the energies associated with some Bohr energy levels in a hydrogen atom. By convention, the potential energy of the electron is defined as zero when it is at an infinite distance from the nucleus. Any more stable arrangement would have a lower potential energy. The energy spacing between orbits gets smaller as the $n$ value increases. For very large values of $n$, the energy levels are so close together that they form a continuum. Some possible electronic transitions corresponding to lines in the hydrogen emission spectrum are indicated by arrows. Transitions in the opposite direction account for lines in the absorption spectrum. The biggest energy change occurs when an electron jumps between $n = 1$ and $n = 2$; a considerably smaller energy change occurs when the electron jumps between $n = 3$ and $n = 4$.

Permissible Values of the Quantum Numbers Through $n = 4$

<table>
<thead>
<tr>
<th>$n$</th>
<th>$\ell$</th>
<th>$m_\ell$</th>
<th>$m_s$</th>
<th>Electron Capacity of Subshell = $4\ell + 2$</th>
<th>Electron Capacity of Shell = $2n^2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0(1s)</td>
<td>0</td>
<td>$\pm \frac{1}{2}$</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>0(2s)</td>
<td>0</td>
<td>$\pm \frac{1}{2}$</td>
<td>2</td>
<td>8</td>
</tr>
<tr>
<td></td>
<td>1(2p)</td>
<td>$-1$, 0, 1</td>
<td>$\pm \frac{1}{2}$</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>0(3s)</td>
<td>0</td>
<td>$\pm \frac{1}{2}$</td>
<td>2</td>
<td>18</td>
</tr>
<tr>
<td></td>
<td>1(3p)</td>
<td>$-1$, 0, 1</td>
<td>$\pm \frac{1}{2}$</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td>2(3d)</td>
<td>$-2$, $-1$, 0, $+1$, $+2$</td>
<td>$\pm \frac{1}{2}$</td>
<td>10</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>0(4s)</td>
<td>0</td>
<td>$\pm \frac{1}{2}$</td>
<td>2</td>
<td>32</td>
</tr>
<tr>
<td></td>
<td>1(4p)</td>
<td>$-1$, 0, 1</td>
<td>$\pm \frac{1}{2}$</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td>2(4d)</td>
<td>$-2$, $-1$, 0, $+1$, $+2$</td>
<td>$\pm \frac{1}{2}$</td>
<td>10</td>
<td></td>
</tr>
<tr>
<td></td>
<td>3(4f)</td>
<td>$-3$, $-2$, $-1$, 0, $+1$, $+2$, $+3$</td>
<td>$\pm \frac{1}{2}$</td>
<td>14</td>
<td></td>
</tr>
</tbody>
</table>

Let us now examine the electronic structures of atoms of different elements. The electronic arrangement that we will describe for each atom is called the ground state electron configuration. This corresponds to the isolated atom in its lowest energy, or unexcited, state. We will consider the elements in order of increasing atomic number, using the periodic table as our guide.

In describing a ground state electron configuration, the guiding idea is that the total energy of the atom is as low as possible. To determine these configurations, we use the **Aufbau Principle** as a guide (The German verb *aufbauen* means “to build up.”):
Each atom is “built up” by (1) adding the appropriate numbers of protons and neutrons in the nucleus as specified by the atomic number and the mass number, and (2) adding the necessary number of electrons into orbitals in the way that gives the lowest total energy for the atom.

As we apply this principle, we will focus on the difference in electronic arrangement between a given element and the element with an atomic number that is one lower. In doing this, we emphasize the particular electron that distinguishes each element from the previous one; however, we should remember that this distinction is artificial because electrons are not really distinguishable. Though we do not always point it out, we must keep in mind that the atomic number (the charge on the nucleus) also differs from one element to the next.

The orbitals increase in energy with increasing value of the quantum number $n$. For a given value of $n$, energy increases with increasing value of $\ell$. In other words, within a particular main shell, the $s$ subshell is lowest in energy, the $p$ subshell is the next lowest, then the $d$, then the $f$, and so on. As a result of changes in the nuclear charge and interactions among the electrons in the atom, the order of energies of the orbitals can vary somewhat from atom to atom. Two general rules help us to predict electron configurations.

1. **Electrons are assigned to orbitals in order of increasing value of $(n + \ell)$**.
2. **For subshells with the same value of $(n + \ell)$, electrons are assigned first to the subshell with lower $n$**.

For example, the $2s$ subshell has $(n + \ell = 2 + 0 = 2)$, and the $2p$ subshell has $(n + \ell = 2 + 1 = 3)$, so we would expect to fill the $2s$ subshell before the $2p$ subshell (Rule 1).

This rule also predicts that the $4s$ subshell $(n + \ell = 4 + 0 = 4)$ will fill before the $3d$ subshell $(n + \ell = 3 + 2 = 5)$. Rule 2 reminds us to fill $2p$ $(n + \ell = 2 + 1 = 3)$ before $3s$ $(n + \ell = 3 + 0 = 3)$ because $2p$ has a lower value of $n$. The usual order of energies of orbitals of an atom and a helpful reminder of this order are shown in Figures:

![Energy Level Diagram](image)

But we should consider these only as a guide to predicting electron arrangements. The observed electron configurations of lowest total energy do not always match those predicted by the
Aufbau guide, and we will see a number of exceptions, especially for elements in the B groups (transition metals) of the periodic table. The electronic structures of atoms are governed by the **Pauli Exclusion Principle**: 

No two electrons in an atom may have identical sets of four quantum numbers.

An orbital is described by a particular allowed set of values for n, ℓ, and m. Two electrons can occupy the same orbital only if they have opposite spins, m. Two such electrons in the same orbital are said to be **spin-paired**, or simply **paired**. A single electron that occupies an orbital by itself is said to be **unpaired**.

**Row 1.** The first shell consists of only one atomic orbital, 1s. This can hold a maximum of two electrons. Hydrogen, as we have already noted, contains just one electron. Helium, a noble gas, has a filled first main shell (two electrons) and is so stable that no chemical reactions of helium are known.

<table>
<thead>
<tr>
<th>Orbital Notation</th>
<th>Simplified Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s</td>
<td></td>
</tr>
<tr>
<td>1H</td>
<td>( \uparrow )</td>
</tr>
<tr>
<td>2He</td>
<td>( \uparrow\downarrow )</td>
</tr>
</tbody>
</table>

**Row 2.** Elements of atomic numbers 3 through 10 occupy the second period, or horizontal row, in the periodic table. In neon atoms the second main shell is filled completely. Neon, a noble gas, is extremely stable. No reactions of it are known.

<table>
<thead>
<tr>
<th>Orbital Notation</th>
<th>Simplified Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s</td>
<td>2s</td>
</tr>
<tr>
<td>Li</td>
<td></td>
</tr>
<tr>
<td>Be</td>
<td></td>
</tr>
<tr>
<td>B</td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
</tr>
<tr>
<td>N</td>
<td></td>
</tr>
<tr>
<td>O</td>
<td></td>
</tr>
<tr>
<td>F</td>
<td></td>
</tr>
<tr>
<td>Ne</td>
<td></td>
</tr>
</tbody>
</table>

---

**Principal Quantum Number** n
**Angular Momentum Quantum Number** ℓ
**Number of Electrons in Orbital or Set of Equivalent Orbitals**

\( 1s \) \( 2s \) \( 2p \)

In the simplified notation, we indicate with superscripts the number of electrons in each subshell.
We see that some atoms have unpaired electrons in the same set of energetically equivalent, or degenerate, orbitals. We have already seen that two electrons can occupy a given atomic orbital (with the same values of n, ℓ, and m_ℓ) only if their spins are paired (have opposite values of m_s). Even with pairing of spins, however, two electrons that are in the same orbital repel each other more strongly than do two electrons in different (but equal-energy) orbitals. Thus, both theory and experimental observations lead to **Hund’s Rule:**

**Electrons occupy all the orbitals of a given subshell singly before pairing begins. These unpaired electrons have parallel spins.**

Thus, carbon has two unpaired electrons in its 2p orbitals, and nitrogen has three.

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**THE PERIODIC TABLE AND ELECTRON CONFIGURATIONS**

In the periodic table, elements are arranged in blocks based on the kinds of atomic orbitals that are being filled. The periodic tables in this text are divided into “A” and “B” groups. The A groups contain elements in which s and p orbitals are being filled. Elements within any particular A group have similar electron configurations and chemical properties, as we shall see in the next chapter. The B groups include the transition metals in which there are one or two electrons in the s orbital of the outermost occupied shell, and the d orbitals, one shell smaller, are being filled. Lithium, sodium, and potassium, elements of the leftmost column of the periodic table (Group 1A), have a single electron in their outermost s orbital (ns^1). Beryllium and magnesium, of Group 2A, have two electrons in their outermost shell, ns^2, while boron and aluminum (Group 3A) have three electrons in their outermost shell, ns^2np^1. Similar observations can be made for other A group elements. The electron configurations of the A group elements and the noble gases can be predicted. However, there are some more pronounced irregularities in the B groups below the fourth period. In the heavier B group elements, the higher energy subshells in different principal shells have energies that are very nearly equal. It is easy for an electron to jump from one orbital to another of nearly the same energy, even in a different set. This is because the orbital energies are perturbed (changed slightly) as the nuclear charge changes, and an extra electron is added in going from one element to the next. This phenomenon gives rise to other irregularities that are analogous to those seen for Cr and Cu, described earlier. We can extend the information in Figure to indicate the electron configurations that are represented by each group (column) of the periodic table.

The periodic table has been described as “the chemist’s best friend.” Chemical reactions involve loss, gain, or sharing of electrons. In this chapter, we have seen that the fundamental basis of the periodic table is that it reflects similarities and trends in electron configurations. It is easy to use the periodic table to determine many important aspects of electron configurations of atoms. Practice until you can use the periodic table with confidence to answer many questions about electron configurations. As we continue our study, we will learn many other useful ways to interpret the periodic table. We should always keep in mind that the many trends in chemical and physical properties that we correlate with the periodic table are ultimately based on the trends in electron configurations.
Example. ELECTRON CONFIGURATION

Use Table to determine the electron configurations of (a) magnesium, Mg; (b) germanium, Ge.

**Plan**

We will use the electron configurations indicated in Table for each group. Each period (row) begins filling a new shell (new value of n). Elements to the right of the d orbital block have the d orbitals in the \( (n-1) \) shell already filled. We often find it convenient to collect all sets of orbitals with the same value of n together, to emphasize the number of electrons in the outermost shell, that is, the shell with the highest value of n.

**Solution**

(a) Magnesium, Mg, is in Group 2A, which has the general configuration \( s^2 \); it is in Period 3 (third row). The last filled noble gas configuration is that of neon, or [Ne]. The electron configuration of Mg is \([\text{Ne}] \) \( 3s^2 \).

(b) Germanium, Ge, is in Group 4A, for which Table shows the general configuration \( s^2p^2 \). It is in Period 4 (the fourth row), so we interpret this as \( 4s^24p^2 \). The last filled noble gas configuration is that of argon, Ar, accounting for 18 electrons. In addition, Ge lies beyond the d orbital block, so we know that the 3d orbitals are completely filled. The electron configuration of Ge is \([\text{Ar}]4s^23d^{10}4p^2 \) or \([\text{Ar}]3d^{10}4s^24p^2 \).
Almost every one of the chemical elements plays some role in Earth’s living systems, however, ~20 elements account for the vast majority of material in living systems. These biogenic elements are divided into:

→ six major biogenic elements (elements found in almost all of Earth’s living systems, often in relatively large quantities),
→ five minor biogenic elements (elements found in many of Earth’s living systems, and/or in relatively small quantities),
→ trace elements (essential elements necessary only in very small quantities to maintain the chemical reactions on which life depends.

The biogenic elements can be classified as:

**MACROELEMENTS** (or macrominerals) – the content in the organism is more than $10^{-2}\%$

**MICROELEMENTS** (or trace elements) – the content in the organism is $10^{-3}$ – $10^{-5}\%$

Macroelements are elements found in almost all of Earth’s living systems. There are 11 of them. Six are called organogens or major biogenic elements. The content of them is 97% in the organism.

### Major Biogenic Elements or Organogens
- Carbon,
- Hydrogen,
- Oxygen,
- Nitrogen,
- Sulfur,
- Phosphorous

### Minor Biogenic Elements
- Sodium,
- Potassium,
- Magnesium,
- Calcium,
- Chlorine

Microelements are essential elements necessary only in very small quantities to maintain the chemical reaction on which life depends. These are:

### Biogenic Trace Elements
- Manganese,
- Iron,
- Cobalt,
- Copper,
- Zinc,
- Boron,
- Aluminum,
- Vanadium,
- Molybdenum,
- Iodine,
- Silicon,
- Nickel,
- Bromine

According to their abundance in the organism, biogenic elements can be classified into macroelements, microelements and contaminating elements.

Macroelements (12 elements in total) form up to 99% of any organism, and can be further subdivided into:

a) a group of stable primary elements (1-60% of total organism weight). These are: O,C, H,N,
b) a group of stable secondary elements (0.05/1% of total organism weight). These are Ca, S, Mg, Cl, Na, K, Fe

Microelements can be divided into three categories:

a) a subgroup of 8 stable elements (less than 0.05%). These are the elements: Cu, Zn, Mn, Co, B, Si, F, I

b) a subgroup of approximately 20 elements that are present at conc. of 0.001% and lower.

c) a subgroup of contaminating elements: Their constant excess in the organism leads to disease: Mn, He, Ar, Hg, Tl, Bi, Al, Cr, Cd.